Chemistry through Inquiry

Teacher Guide

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High School

Teacher Guide

21st Century Science

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Introduction

PASCO scientific's probeware and laboratory investigations move students from the low-level task of memorization of science facts to higher-level tasks of data analysis, concept construction, and application. For science to be learned at a deep level, it is essential to combine the teaching of abstract science concepts with "real-world" science investigations. Hands-on, technology-based, laboratory experiences serve to bridge the gap between the theoretical and the concrete, driving students toward a greater understanding of natural phenomenon. Students also gain important science process skills that include: developing and using models, carrying out investigations, interpreting data, and using mathematics.

At the foundation of teaching science are a set of science standards that clearly define the science content and concepts, the instructional approach, and connections among the science disciplines. The Next Generation Science Standards (2012)© are a good example of a robust set of science standards.

The Next Generation Science Standards (NGSS) position student inquiry at the forefront. The standards integrate and enhance science, technology, engineering, and math (STEM) concepts and teaching practices. Three components comprise these standards: Science and Engineering Practices, Disciplinary Core Ideas, and Crosscutting Concepts. The lab activities in PASCO's 21st Century Science Guides are all correlated to the NGSS (see http://pasco.com).

- The *Science and Engineering Practices* help students to develop a systematic approach to problem solving that builds in complexity from kindergarten to their final year in high school. The practices integrate organization, mathematics and interpretive skills so that students can make data-based arguments and decisions.

- *Disciplinary Core Ideas* are for the physical sciences, life sciences, and earth and space sciences. The standards are focused on a limited set of core ideas to allow for deep exploration of important concepts. The core ideas are an organizing structure to support acquiring new knowledge over time and to help students build capacity to develop a more flexible and coherent understanding of science.

- *Crosscutting Concepts* are the themes that connect all of the sciences, mathematics and engineering. As students advance through school, rather than experiencing science as discrete, disconnected topics, they are challenged to identify and practice concepts that cut across disciplines, such as "cause and effect". Practice with these concepts that have broad application helps enrich students' understanding of discipline-specific concepts.

PASCO's lab activities are designed so that students complete guided investigations that help them learn the scientific process and explore a core topic of science, and then are able to design and conduct extended inquiry investigations. The use of electronic sensors reduces the time for data collection, and increases the accuracy of results, providing more time in the classroom for independent investigations.

In addition to supporting the scientific inquiry process, the lab activities fulfill STEM education requirements by bringing together science, technology, engineering, and math. An integration of these areas promotes student understanding of each of these fields and develops their abilities to become self-reliant researchers and innovators. When faced with an idea or problem, students learn to develop, analyze, and evaluate possible solutions. Then collaborate with others to construct and test a procedure or product.

Information and computer tools are essential to modern lab activities and meeting the challenge of rigorous science standard, such as NGSS. The use of sensors, data analysis and graphing tools, models and simulations, and work with instruments, all support the science and engineering practices as implemented in a STEM-focused curriculum, and are explicitly cited in NGSS. PASCO's lab activities provide students with hands-on and minds-on learning experiences, making it possible for them to master the scientific process and the tools to conduct extended scientific investigations.

About the PASCO 21st Century Science Guides

This manual presents teacher-developed laboratory activities using current technologies to help you and your students explore topics, develop scientific inquiry skills, and prepare for state level standardized exams. Using electronic-sensor data collection, display and analysis devices in your classroom fulfills STEM requirements and provides several benefits. Sensor data collection allows students to:

- observe phenomena that occur too quickly or are too small, occur over too long a time span, or are beyond the range of observation by unaided human senses
- perform measurements with equipment that can be used repeatedly over the years
- collect accurate data with time and/or location stamps
- rapidly collect, graphically display, and analyze data so classroom time is used effectively
- practice using equipment and interpreting data produced by equipment that is similar to what they might use in their college courses and adult careers

The Data Collection System

"Data collection system" refers to PASCO's DataStudio®, the Xplorer GLX™, SPARKvue™, and SPARK Science Learning System™ and PASCO Capstone™. Each of these can be used to collect, display, and analyze data in the various lab activities.

Activities are designed so that any PASCO data collection system can be used to carry out the procedure. The DataStudio, Xplorer GLX, SPARKvue, or SPARK Science Learning System Tech Tips provide the steps on how to use the data collection system and are available on the storage device that came with your manual. For assistance in using PASCO Capstone, refer to its help system.

Getting Started with Your Data Collection System

To help you and your students become familiar with the many features of your data collection system, start with the tutorials and instructional videos that are available on PASCO's website (www.pasco.com).

Included on the storage device accompanying your manual is a Scientific Inquiry activity that acts as a tutorial for your data collection system. Each data collection system (except for PASCO Capstone) has its own custom Scientific Inquiry activity. The activity introduces students to the process of conducting science investigations, the scientific method, and introduces teachers and students to the commonly used features of their data collection system. Start with this activity to become familiar with the data collection system.

Teacher and Student Guide Contents

All the teacher and student materials are included on the storage device accompanying the Teacher Guide.

Lab Activity Components

Each activity has two components: Teacher Information and Student Inquiry Worksheets.

Teacher Information is in the Teacher Guide. It contains information on selecting, planning, and implementing a lab, as well as the complete student version with answer keys. Teacher Information includes all sections of a lab activity, including objectives, procedural overview, time requirements, and materials and equipment at-a-glance.

Student Inquiry Worksheets begin with a driving question, providing students with a consistent scientific format that starts with formulating a question to be answered in the process of conducting a scientific investigation.

This table identifies the sections in each of these two activity components.

TEACHER INFORMATION	STUDENT INQUIRY WORKSHEET
Objectives	Driving Questions
Procedural Overview	Background
Time Requirement	Pre-Lab Activity
Materials and Equipment	Materials and Equipment
Concepts Students Should Already Know	
Related Labs in This Guide	
Using Your Data Collection System	
Background	
Pre-Lab Activity	
Lab Preparation	
Safety	Safety
Sequencing Challenge	Sequencing Challenge
Procedure With Inquiry	Procedure (+ conceptual questions)
Data Analysis	Data Analysis
Analysis Questions	Analysis Questions
Synthesis Questions	Synthesis Questions
Multiple Choice Questions	Multiple Choice Questions
Extended Inquiry Suggestions	

Electronic Materials

- ♦ The storage device with PASCO materials and the storage device with ODYSSEY® materials accompany this manual. See the “Using ODYSSEY Molecular Labs” section for details on ODYSSEY software.

The storage device accompanying this manual contains the following:

- ♦ Complete Teacher Guide and Student Guide with Student Inquiry Worksheets in PDF format.
- ♦ The Scientific Inquiry activity for SPARK™, SPARKvue™, Xplorer GLX®, and DataStudio® and the Student Inquiry Worksheets for the laboratory activities are in an editable Microsoft™ Word format. PASCO provides editable files of the student lab activities so that teachers can customize activities to their needs.
- ♦ Tech Tips for the SPARK, SPARKvue, Xplorer GLX, DataStudio, and individual sensor technologies in PDF format.
- ♦ User guides for SPARKvue and GLX.
- ♦ DataStudio and PASCO Capstone® Help is available in the software application itself.

Using ODYSSEY Molecular Labs

Wavefunction's ODYSSEY is a unique software program for use in chemistry classes. With ODYSSEY students can use scientifically based simulations to experiment with core chemistry topics from a molecular perspective. The software enhances and complements the hands-on, experiential PASCO activities in this manual.

ODYSSEY includes a collection of ready-to-use chemistry experiments called “Molecular Labs” and student worksheets. A number of the Molecular Labs applicable to the PASCO activities are identified under selected topic areas in the table of contents. The student worksheets for these labs are provided in the ODYSSEY Molecular Labs section of this manual and the answer key can be found on the accompanying storage device. A fully functional 60-day licensed version of the ODYSSEY Instructor Edition, containing the complete set of ODYSSEY’s Molecular Labs, is included with this manual.

In addition to the Molecular Labs, ODYSSEY provides:

Prelabs which serve as tutorials – ideal learning about how to use the program

Applied Chemistry – a collection of chemistry samples commonly encountered in modern society

Molecular Stockroom – the electronic equivalent of your chemistry stockroom with more than a thousand pre-constructed samples spanning the periodic table

To successfully get started with ODYSSEY, check the system requirements and install the software that is on the accompanying ODYSSEY storage device; use the activation code provided to access the software for 60 days. Contact PASCO (www.pasco.com) for information on instructor and student licensing.

International Baccalaureate Organization (IBO*) Support

IBO Diploma Program

The International Baccalaureate Organization (IBO) uses a specific science curriculum model that includes both theory and practical investigative work. While this lab guide was not produced by the IBO and does not include references to the internal assessment rubrics, it does provide a wealth of information that can be adapted easily to the IB classroom.

By the end of the IB Diploma Program students are expected to have completed a specified number of practical investigative hours and are assessed using the specified internal assessment criteria. Students should be able to design a lab based on an original idea, carry out the procedure, draw conclusions, and evaluate their own results. These scientific processes require an understanding of laboratory techniques and equipment as well as a high level of thinking.

Using these Labs with the IBO Programs

The student versions of the labs are provided in Microsoft Word and are fully editable. Teachers can modify the labs easily to fit a problem-based format.

For IB students, pick one part of the internal assessments rubrics to go over with the students. For example, review the design of the experiment and have students explain what the independent, dependent, and controlled variables are in the experiment. Ask students to design a similar experiment, but change the independent variable.

Delete certain sections. As students become familiar with the skills and processes needed to design their own labs, start deleting certain sections of the labs and have students complete those parts on their own. For example, when teaching students to write their own procedures, have the students complete one lab as it is in the lab guide. In the next lab, keep the Sequencing Challenge, but have students write a more elaborate procedure. Finally, remove both the Sequencing Challenge and the Procedure sections and have students write the entire procedure.

Encourage students to make their own data tables. Leave the procedure, but remove the data tables and require the students to create them on their own. In another lab, leave the driving question and procedure, but remove the analysis questions and have students write their own analysis, conclusion, and evaluation.

Use only the driving question. As students' progress through their understanding of the structure of an experiment, provide them with just the driving question and let them do the rest. Some of the driving questions are too specific (they give the students the independent variable), so revise them appropriately.

Extended inquiry. After students complete an activity in the lab guide, use the extended inquiry suggestions to have the students design their own procedure, or the data collection and processing, or both.

About Correlations to Science Standards

The lab activities in this manual are correlated to a number of standards, including United States National Science Education Standards, the Next Generation Science Standards, and all State Science Standards. See http://pasco.com for the correlations.

Global Number Formats and Standard Units

Throughout this guide, the International System of Units (SI) or metric units is used unless specific measurements, such as air pressure, are conventionally expressed otherwise. In some instances, such as weather parameters, it may be necessary to alter the units used to adapt the material to conventions typically used and widely understood by the students.

Reference

NGSS Lead States. 2013. Next Generation Science Standards: For States, By States. Washington, DC: The National Academies Press.

Normal Laboratory Safety Procedures

Overview

PASCO is concerned with your safety and because of that, we are providing a few guidelines and precautions to use when exploring the labs in our Chemistry guide. This is a list of general guidelines only; it is by no means all-inclusive or exhaustive. Of course, common sense and standard laboratory safety practices should be followed.

Regarding chemical safety, some of the substances and chemicals referred to in this manual are regulated under various safety laws (local, state, national, or international). Always read and comply with the safety information available for each substance or chemical to determine its proper storage, use and disposal.

Since handling and disposal procedures vary, our safety precautions and disposal comments are generic. Depending on your lab, instruct students on proper disposal methods. Each of the lab activities also has a Safety section for procedures necessary for that activity.

General Lab Safety Procedures and Precautions

- Follow all standard laboratory procedures
- Absolutely no food, drink, or chewing gum is allowed in the lab.
- Keep water away from electrical outlets.
- Wear eye protection (splash-proof goggles), lab apron, and protective gloves.
- Do not touch your face with gloved hands. If you need to sneeze or scratch, take off your gloves, wash your hands, and then take care of the situation. Do not leave the lab with gloves on.
- Wash your hands after handling chemicals, glassware, and equipment.
- Know the safety features of your lab such as eye-wash stations, fire extinguisher, first-aid equipment or emergency phone use.
- Insure that loose hair and clothing is secure when in the lab.
- Handle glassware with care.
- Insure you have adequate clear space around your lab equipment before starting an activity.
- Do not wear open toe shoes or short pants in the laboratory.
- Allow heated objects and liquids to return to room temperature before moving.
- Never run or joke around in the laboratory.
- Do not perform unauthorized experiments.
- Students should use a buddy system in case of trouble.
- Keep the work area neat and free from any unnecessary objects.

Water Related Safety Precautions and Procedures

- Keep water away from electrical outlets.
- Keep water away from all electronic equipment.

Chemical Related Safety Precautions and Procedures

- Consult the manufacturer's Material Safety Data Sheets (MSDS) for instructions on handling, storage, and disposing of chemicals. Your teacher should provide the MSDS sheets of the chemicals that you are using. Keep these instructions available in case of accidents.
- Many chemicals are hazardous to the environment and should not be disposed of down the drain. Always follow your teacher’s instructions for disposing of chemicals.
- Sodium hydroxide, hydrochloric acid, and acetic acid are corrosive irritants. Avoid contact with the eyes and wash your hands after handling. In case of skin exposure, wash it off with plenty of water.
- Always add acids and bases to water, not the other way around, as the solutions may boil vigorously.
- Diluting acids and bases creates heat; be extra careful when handling freshly prepared solutions and glassware, as they may be very hot.
- Handle concentrated acids and bases in a fume hood; the fumes are caustic and toxic.
- Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable. Use nitrile gloves if you have latex allergy.
- Read labels on all chemicals and pay particular attention to hazard icons and safety warnings.
- When handling any bacterial species, follow aseptic techniques.
- Wash your hands before and after a laboratory session.
- If any solution comes in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- Follow the teacher’s instructions for disposing of chemicals.
- Check the label to verify it is the correct substance before using it.
- Never point the open end of a test tube containing a substance at yourself or others.
- Use a wafting motion when smelling chemicals
- Do not return unused chemicals to their original container.
- Keep flammable chemicals from open flame.

Dangerous or Harmful Substance Related Lab Safety Precautions

- When handling any bacterial species, follow aseptic techniques.
- Always flame inoculating loops and spreaders before setting them down on the lab bench.
- Pipetting suspension cultures can create an aerosol. Keep your nose and mouth away from the tip of the pipet to avoid inhaling any aerosol
- Use caution when working with acids.
- Use appropriate caution with the matches, burning splint and foods, and other hot materials.
- Be careful using a knife or scalpel.

Other Safety Precautions

- If water is boiled for an experiment involving heat, make sure it is never left unattended. Remember, too, that the hot plate will stay hot well after it is unplugged or turned off.
- Any injury must be reported immediately to the instructor, an accident report has to be completed by the student or a witness.
- If you are suffering from any allergy, illness, or are taking any medication, you must inform the instructor. This information could be very important in an emergency.
- Try to avoid wearing contact lenses. If a solution spills in your eye, the presence of a contact lens makes first aid difficult and can result in permanent damage. Also, organic solvents tend to dissolve in soft contact lenses, causing eye irritation.

Additional Resources

- Flinn Scientific
- The Laboratory Safety Institute (LSI)
- National Science Education Leadership Association (NSELA)/Safe Science Series

Master Materials and Equipment List

Italicized entries indicate items not available from PASCO. The quantity indicated is per student or group. Note: The activities also require protective gear for each student (for example, safety goggles, gloves, apron, or lab coat).

Teachers can conduct some lab activities with sensors other than those listed here. For assistance with substituting compatible sensors for a lab activity, contact PASCO Teacher Support (800-772-8700 inside the United States or http://www.pasco.com/support).

Act	Title	Materials and Equipment	Qty
0	**Scientific Inquiry** This lab is designed to help student familiarize themselves with their data collection system while engaging in scientific investigations.	Data Collection System PASPORT® Temperature Sensor[1] *Cup, 270-mL (9-oz)* *Hot water* *Insulating materials readily available in the laboratory (polystyrene, foil, plastic wrap, cloth, wool, packing peanuts)*	1 1 1 500 mL A variety
1	**Significant Figures** Determine the correct number of significant figures to include when reporting a measurement or a calculated value based upon measurements.	From the PASCO Significant Figure Single, Four-scale meter stick *Graduated cylinder, 10-mL,* *Graduated cylinder, 100-mL,* *Beaker, 100-mL,* *Irregular-shaped object* *Regular-shaped object*	1 1 1 1 1 1
2	**Density** Determine that density is an intensive property of a substance independent of the shape or size of an object.	PASCO Density Set *Beaker, 150-mL* *Graduated cylinder, 50- or 100-mL* Balance Overflow can *Metric ruler (or* calipers*)* *Water* String	1 1 1 2 or 3 per class 1 1 500 mL 1
3	**Graphing Mass versus Volume to Determine Density** Use multiple mass and volume data to graphically determine the density of a substance.	From the PASCO Discover Density Set: four different-sized rectangular aluminum pieces, four different-sized rectangular plastic pieces of the same composition Balance *Metric ruler (or* calipers*)*	1 set 2 or 3 per class 1

Act	Title	Materials and Equipment	Qty
4	**Percent Oxygen in Air** Use an absolute pressure sensor to learn about the components of air and how to determine the percent of oxygen in air.	Data Collection System	1
		PASPORT Absolute Pressure Sensor	1
		PASPORT Sensor Extension Cable	1
		Quick-release connector[2]	1
		Tubing connector[2]	1
		Tubing, 1- to 2-cm[2]	1
		Beaker, 150-mL	1
		Test tube, 25-mm × 150-mm	1
		One-hole rubber stopper to fit test tubes	1
		Stir rod	1
		White vinegar (~5% acetic acid)	50 to 60 mL
		Steel wool, fine mesh (#000)	1 g
		Paper towels	As needed
		Glycerin	2 drops
5	**Conservation of Matter** Test the law of conservation of matter for both physical and chemical changes by finding the mass of the reactants before the chemicals are reacted and the mass of the products after the reaction has occurred.	Balance	1
		Test tube, 15-mm × 100-mm	2
		Beaker, 250-mL	1
		Plastic soda bottle (with cap), 500-mL	1
		Sodium nitrate	5 g
		0.1 M Sodium sulfate	5 mL
		0.1 M Strontium chloride	5 mL
		Sodium bicarbonate	8 g
		5% Acetic acid	30 mL
		Distilled (deionized) water	10 mL
6	**Properties of Ionic and Covalent Compounds** Use a conductivity sensor to determine if an unknown substance is an ionic, polar covalent, or non-polar covalent compound based on its physical properties.	Data Collection System	1
		PASPORT Conductivity Sensor	1
		Hot plate	1
		Graduated cylinder, 10-mL	1
		Test tube, 15-mm × 100-mm	5
		Test tube rack	1
		Stopper to fit test tubes	3
		Spatula	1
		Tongs	1
		Aluminum foil squares, 5-cm × 5-cm	6
		Masking tape	1
		Wash bottle and waste container	1
		Distilled (deionized) water	30 mL
		Table salt (sodium chloride)	1 g
		Table sugar (sucrose)	1 g
		Paraffin wax	1 g
		Unknown A (use glucose)	1 g
		Unknown B (use crayon pieces)	1 g
		Unknown C (use potassium chloride)	1 g

Act	Title	Materials and Equipment	Qty
7	**Electrolyte versus Non-Electrolyte Solutions** Use a conductivity sensor to determine which substances in sports drinks (water, sugars, or salts) are electrolytes.	Data Collection System PASPORT Conductivity Sensor *Test tube, 20-mm × 150-mm* *Beaker for collecting rinse water* *Test tube rack* *Funnel* *Wash bottle filled with distilled (deionized) water* *Sucrose solutions (0.02 M, 0.04 M, 0.06 M, 0.08 M, 0.10 M)* *Sodium chloride solutions (0.02 M, 0.04 M, 0.06 M, 0.08 M, 0.10 M)* *Distilled (deionized) water* *Sports drink*	1 1 6 1 1 1 1 10 mL of each 10 mL of each 50 mL 10 mL
8	**Boyle's Law** Use an absolute pressure sensor to determine the effect of volume on the pressure of a closed system containing a fixed amount of molecules at a constant temperature.	Data Collection System PASPORT Absolute Pressure Sensor PASPORT Sensor Extension Cable Tubing, 1- to 2-cm[2] Quick-release connector[2] Syringe, 20-mL or 60-mL[2] *Glycerin*	1 1 1 1 1 1 2 drops
9	**Gay-Lussacs's Law and Absolute Zero** Use an absolute pressure sensor and fast response temperature sensor to determine the temperature at which all motion stops (absolute zero).	Data Collection System PASPORT Absolute Pressure Sensor PASPORT Fast Response Temperature Sensor PASPORT Sensor Extension Cable Quick-release connector[2] Tubing connector[2] Tubing, 1- to 2-cm[2] *Test tube, 15-mm × 100-mm* *One-hole rubber stopper to fit test tubes* *Beaker, 250-mL* Ring stand Three-finger clamp *Glycerin* *Polystyrene cup* *Rubber band* *Crushed ice* *Room temperature water* *~45 °C water* *~55 °C water* *~65 °C water*	1 1 1 1 1 1 1 1 1 2 1 1 2 drops 2 1 300 mL 300 mL 300 mL 300 mL 300 mL

Act	Title	Materials and Equipment	Qty
10	**Phase Change** Use a fast response temperature sensor and stainless steel temperature sensor to determine how to add heat to a substance without the temperature of the substance increasing.	Data Collection System	1
		PASPORT Stainless Steel Temperature Sensor	1
		Hot plate	1
		Beaker, 150-mL or larger	2
		Graduated cylinder, 10-mL	1
		Test tube, 10-mm × 100-mm	1
		Test tube rack	1
		Ring stand	1
		Utility clamp	1
		Stir rod	1
		Tablespoon	1
		Distilled (deionized) water	104 mL
		Crushed ice to fill the beaker	1
		Rock salt	200 g
11	**Specific Heat** Use a fast response temperature sensor to determine the identity of an unknown metal by calculating the specific heat of the metal and comparing it to a list of known values.	Data Collection System	1
		PASPORT Fast Response Temperature Sensor	1
		Beaker, 250-mL	1
		Beaker, 400-mL	1
		Graduated cylinder, 100-mL	1
		Balance, centigram	1
		Thermometer (or PASPORT Stainless Steel Temperature Sensor)	1
		Hot plate	1
		Tongs	1
		Polystyrene cup	2
		Lid for the polystyrene cup	1
		Paper towels	As needed
		Water (from the tap)	250 mL
		Distilled (deionized) water	200 mL
		Metal sample, unknown, up to 4 × 4 × 4 cm	1
12	**Heat of Fusion** Use a fast response temperature sensor and calorimetry to determine the heat of fusion for water.	Data Collection System	1
		PASPORT Fast Response Temperature Sensor	1
		Graduated cylinder, 100-mL	1
		Beaker, 250-mL	2
		Hot plate	1
		Polystyrene cup	2
		Lid for polystyrene cup	1
		Paper towels	As needed
		Water	200 mL
		Ice cube	2

Act	Title	Materials and Equipment	Qty
13	**Intermolecular Forces** Use a stainless steel temperature sensor to determine the effects of molecular size and shape on the strength of intermolecular forces for different alcohols within the same homologous series and between isomeric pairs.	Data Collection System	1
		PASPORT Stainless Steel Temperature Sensor	1
		Graduated cylinder, 10-mL	1
		Test tube, 15-mm × 100-mm	7
		Test tube rack	1
		Stopper to fit test tube	7
		Wash bottle and waste container	1
		Masking tape, 6-cm strips	2
		Methanol	5 mL
		Ethanol	5 mL
		Propanol	5 mL
		Butanol	5 mL
		Pentanol	5 mL
		2-Propanol	5 mL
		2-Butanol	5 mL
14	**Concentration of a Solution: Beer's Law** Use a colorimeter to determine the concentration of a copper(II) sulfate solution.	Data Collection System	1
		PASPORT Colorimeter	1
		PASPORT Sensor Extension Cable[2]	1
		Glass cuvette with cap	7
		Beaker, 100-mL	2
		Test tube, 20-mm × 150-mm	6
		Test tube rack	1
		Volumetric pipet with bulb or a pump, 10-mL	2
		Non-abrasive cleaning tissue	1
		0.80 M Copper(II) sulfate	30 mL
		Unknown copper(II) sulfate (a solution less than 1.0 M)	10 mL
		Distilled (deionized) water	30 mL
15	**pH of Household Chemicals** Use a pH sensor and common household chemicals to relate pH and hydronium ion (H_3O^+) concentration, classifying solutions as acidic, basic, or neutral.	Data Collection System	1
		PASPORT pH Sensor	1
		Beaker, 50-mL	2
		Graduated cylinder, 50-mL	1
		Graduated cylinder, 10-mL	1
		Test tube, 15-mm × 100-mm	10
		Test tube rack	1
		Wash bottle and waste container	1
		Buffer solution pH 4	25 mL
		Buffer solution pH 10	25 mL
		White vinegar (~5% acetic acid)	5 mL
		Lemon Juice	5 mL
		Soft drink	5 mL
		Window cleaner	5 mL
		Water (from the tap)	5 mL
		Milk	5 mL
		Coffee	5 mL
		0.5 M Sodium bicarbonate	5 mL
		Liquid soap	5 mL
		Bleach	5 mL

Act	Title	Materials and Equipment	Qty
16	**Electrochemical Battery: Energy from Electrons** Use a voltage sensor to place metal reactants in their proper order on the table of standard electrode potentials.	Data Collection System	1
		PASPORT Voltage Sensor	1
		Beaker, 50-mL	2
		Alligator clip, 1 black,1 red	2
		Wash bottle and waste container	1
		Thick string or yarn	20 cm
		Knife to cut fruit	1
		Paper towels	As needed
		0.1 M Sodium chloride	5 to 10 mL
		0.1 M Hydrochloric acid	50 mL
		Copper strip	1
		Zinc strip	1
		Magnesium strip	1
		Nickel strip	1
		Iron strip	1
		Lemon	1
		Tomato	1
17	**Evidence of a Chemical Reaction** Use a fast response temperature sensor to distinguish between physical changes and chemical reactions and identify unknown changes as either physical changes or chemical reactions using evidence to support your decision.	Data Collection System	1
		PASPORT Fast Response Temperature Sensor	1
		Balance	2 or 3 per class
		Hot plate	1
		Graduated cylinder, 100-mL	1
		Graduated cylinder, 10-mL	1
		Beaker, 250-mL	2
		Test tube, 15-mm x 100-mm	7
		Test tube rack	1
		Test tube holder	1
		Stir rod	1
		Spatula	1
		Beaker for collecting rinse water	1
		Weighing paper	1
		Wash bottle filled with distilled (deionized) water	1
		Water (from the tap)	255 mL
		Calcium carbonate	~0.2 g
		White vinegar (~5% acetic acid)	2 mL
		1.0 M Citric acid	2 mL
		1.0 M Sodium bicarbonate	2 mL
		0.5 M Copper(II) sulfate	2 mL
		1.0 M Sodium hydroxide	2 mL
		0.05 M Silver nitrate	2 mL
		0.1 M Sodium chloride	2 mL
		Lauric acid	~0.5 g
		Effervescent tablet	1
		Colored drink powder	~0.2 g

Act	Title	Materials and Equipment	Qty
18	**Stoichiometry** Use a temperature sensor to determine the mole ratio between the reactants sodium hypochlorite and sodium thiosulfate.	Data Collection System PASPORT Temperature sensor[1] *Graduated cylinder, 10-mL* *Graduated cylinder, 50- or 100-mL* *Transfer pipet* *Test tube, 20-mm × 150-mm* *Test tube rack* *Wash bottle filled with water* *Waste container* *0.5 M Sodium hypochlorite* *0.5 M Sodium thiosulfate, in 0.2 M sodium hydroxide*	1 1 2 2 2 7 1 1 1 35 to 40 mL 35 to 40 mL
19	**Single Replacement Reactions** Use a colorimeter to determine the mass of copper consumed and silver deposited in a single replacement reaction.	Data Collection System PASPORT Colorimeter PASPORT Sensor Extension Cable[2] Glass cuvette with cap[2] Balance, centigram *Test tube, 20-mm × 150-mm* *Test tube rack* *Graduated cylinder, 100-mL* *Sand paper or steel wool* *Non-abrasive cleaning tissue* *0.5 M Silver nitrate solution* *Copper wire* *Paper towels*	1 1 1 1 1 1 1 1 1 1 30 mL 20 cm As needed
20	**Molar Mass of Copper** Use a voltage-current senosr to determine the molar mass of copper through electroplating in an electrolytic cell.	Data Collection System PASPORT Voltage-Current Sensor Balance, centigram *Beaker, 250-mL* Utility clamps, insulated Ring stand Magnetic stirrer Magnetic stir bar[2] DC power supply Red patch cord, 4-mm banana plug[2] Black patch cord, 4-mm banana plug Alligator clip[2] *Copper electrode* *Stainless steel spoon (or other item to electroplate)* *0.5 M Copper(II) sulfate*	1 1 1 1 2 1 1 1 1 2 1 2 1 1 150 mL

Act	Title	Materials and Equipment	Qty
21	**Double Replacement Reactions** Using a titration, determine the amount of chloride ion in water samples.	Ring stand	1
		Buret clamp	1
		Buret, 50-mL	1
		Funnel	1
		Magnetic stirrer	1
		Magnetic stir bar[2]	1
		Transfer pipet	1
		Waste container	1
		Erlenmeyer flask, 125-mL	4
		Graduated cylinder, 50-mL	1
		0.2% Disodium salt fluorescein indicator	2 mL
		1% Dexrin solution	100 mL
		0.020 M Silver nitrate	200 mL
		0.010 M Sodium chloride	100 mL
		Swimming pool water	100 mL
22	**Rates of Reaction** Use an absolute pressure sensor to determine the effect of temperature, concentration, and surface area on the rate of a chemical reaction by measuring changes in absolute pressure as a reaction proceeds.	Data Collection System	1
		PASPORT Absolute Pressure Sensor	1
		PASPORT Sensor Extension Cable	1
		Test tube, 20-mm × 150-mm	3
		Test tube rack	1
		One-hole rubber stopper to fit test tube	1
		Quick-release connector[2]	1
		Tubing, 1- to 2-cm[2]	1
		Tubing connector[2]	1
		Glycerin	1
		4.0 M Hydrochloric acid	5 mL
		2.0 M Hydrochloric acid	5 mL
		1.0 M Hydrochloric acid	20 mL
		0.1 M Hydrochloric acid	5 mL
		Warm and cold water baths	One per class
		Magnesium ribbon, 1-cm pieces	18
		Magnesium powder	0.05 g

Act	Title	Materials and Equipment	Qty
23	**Ideal Gas Law** Use an absolute pressure sensor and stainless steel temperature sensor to determine the number of moles of carbon dioxide gas generated during a reaction between hydrochloric acid and sodium bicarbonate.	Data Collection System PASPORT Absolute Pressure Sensor PASPORT Stainless Steel Temperature Sensor Blue plastic tubing for the temperature sensor[2] PASPORT Sensor Extension Cable Balance, centigram Graduated cylinder or volumetric pipet, 10-mL *Graduated cylinder, 1000-mL* *Test tube, 15-mm × 100-mm* *Plastic bottle, 300- to 500-mL* *Two-hole rubber stopper that fits the plastic bottle* Quick-release connector[2] Tubing, 1- to 2-cm[2] Tubing connector[2] *1.0 M Hydrochloric acid* *Sodium bicarbonate* *Glycerin* *Paper towels*	1 1 1 1 1 1 1 1 1 1 1 1 1 1 10 mL 0.80 g 2 drops As needed
24	**Heats of Reaction and Solution** Use a temperature sensor to determine the molar heat of solution for sodium hydroxide and ammonium chloride when they are dissolved in water, and the molar heat of reaction when magnesium reacts with hydrochloric acid.	Data Collection System PASPORT Temperature Sensor[1] *Beaker, 250-mL* *Graduated cylinder, 50-mL* Balance, centigram *Polystyrene cup* *Spatula* *Stir rod* *Paper towels* *Weighing paper* *Sand paper or steel wool* *Wash bottle and waste container* *Sodium hydroxide pellets* *Ammonium chloride* *Magnesium metal ribbon* *1.0 M Hydrochloric acid* *Distilled (deionized) water*	1 1 1 1 1 2 1 1 As needed 1 1 piece 1 1 g 1 g 0.10 g 25 mL 50 mL

Act	Title	Materials and Equipment	Qty
25	**Hess's Law** Use a temperature sensor to show that the change in enthalpy for the reaction between solid sodium hydroxide and aqueous hydrochloric acid can be determined using both a direct and an indirect method.	Data Collection System	1
		PASPORT Temperature Sensor[1]	1
		Beaker, 250-mL	1
		Graduated cylinder, 50-mL	1
		Spatula	1
		Polystyrene cup	2
		Lid for polystyrene cup	1
		Weighing paper	2
		Wash bottle and waste container	1
		1.0 M Hydrochloric acid	25 mL
		0.5 M Hydrochloric acid	50 mL
		1.0 M Sodium hydroxide	25 mL
		Sodium hydroxide pellets	2.0 g
		Distilled (deionized) water	50 mL
26	**An Acid-Base Titration** Use a drop counter and pH sensor to to determine the concentration of a hydrochloric acid solution and the concentration of an acetic acid solution by titration.	Data Collection System	1
		PASPORT Drop Counter	1
		PASPORT pH Sensor	1
		Acetic acid solution	10 mL
		Magnetic stirrer	1
		Micro stir bar[2]	1
		Beaker, 250-mL	2
		Beaker, 50-mL	2
		Graduated cylinder, 100-mL	1
		Volumetric pipet or graduated cylinder, 10-mL	1
		Buret, 50-mL	1
		Ring stand	1
		Right-angle clamp	1
		Buret clamp	1
		Funnel	1
		Transfer pipet	1
		Waste container	1
		Wash bottle filled with distilled (deionized) water	1
		Buffer solution, pH 4	25 mL
		Buffer solution, pH 10	25 mL
		Distilled (deionized) water	200 mL
		Hydrochloric acid solution (~0.1 M)	10 mL
		Acetic acid solution (~0.1 M)	10 mL
		Standardized sodium hydroxide solution (~0.1 M)	120 mL

Act	Title	Materials and Equipment	Qty
27	**Diprotic Titration: Multi-Step Chemical Reactions** Use a drop counter and a pH sensor to determine the concentration of a sodium carbonate solution, learning that chemical reactions can be the sum of several individual reactions.	Data Collection System PASPORT Drop Counter PASPORT pH Sensor Micro stir bar[2] Magnetic stirrer *Beaker, 50-mL* *Beaker, 250-mL* *Graduated cylinder, 50-mL* *Graduated cylinder, 100-mL* *Transfer pipet* *Buret, 50-mL* *Buret clamp* Ring stand Right-angle clamp *Funnel* *Waste container* *Wash bottle filled with distilled (deionized) water* Buffer solution, pH 4 Buffer solution, pH 10 *Distilled (deionized) water* *Sodium carbonate solution* *1.0 M Hydrochloric acid*	1 1 1 1 1 2 1 *1* 1 1 1 1 1 1 1 1 1 25 mL 25 mL 200 mL 40 mL 110 mL
28	**Le Chatelier's Principle** Use a pH sensor to determine the effect of concentration changes on the equilibrium of a system, relating pH values with the acid-base indicator phenolphthalein.	Data Collection System PASPORT pH Sensor *Beaker, 100-mL* *Beaker, 50-mL* *Graduated cylinder, 25-mL* *Graduated cylinder, 50- or 100-mL* *Transfer pipet* *Waste container* *Wash bottle filled with distilled (deionized) water* Buffer solution pH 4 Buffer solution pH 10 *Distilled (deionized) water* *Phenolphthalein indicator* *0.1 M Hydrochloric acid* *0.1 M Sodium hydroxide* *0.5 M Acetic acid* *0.5 M Sodium acetate*	1 1 2 2 1 1 3 1 1 25 mL 25 mL 100 mL 4 drops 5mL 5mL 50 mL 10 mL

[1]Either the PASPORT Fast Response Temperature Sensor or the PASPORT Stainless Steel Temperature Sensor can be used for this activity.

[2]These items are included with the specific apparatus or sensor used in the experiment.

Activities by PASCO Equipment

This list shows the PASCO specific equipment used in each lab activity. The Chemistry Sensor is a MultiMeasure™ sensor that contains a PASPORT Absolute Pressure Sensor, a PASPORT pH Sensor, a PASPORT Stainless Steel Temperature Sensor, and a PASPORT Voltage Sensor.

Items Available from PASCO	Qty	Activity Where Used
PASCO Density Set	1	2
PASCO Discover Density Set	1	3
PASCO Significant Figure Set	1	1
PASPORT Absolute Pressure Sensor[1]	1	4, 8, 9, 22, 23,
PASPORT Colorimeter	1	14, 19
PASPORT Conductivity Sensor	1	6, 7
PASPORT Drop Counter	1	26, 27
PASPORT Fast Response Temperature Sensor	1	9, 11, 12, 17
PASPORT pH Sensor[1]	1	15, 26, 27, 28
PASPORT Stainless Steel Temperature Sensor[1]	1	10, 13, 23
PASPORT Temperature Sensor[2]	1	0, 18, 24, 25
PASPORT Voltage Sensor[1]	1	16
PASPORT Voltage-Current Sensor	1	20

[1]This sensor is part of the Chemistry Sensor

[2]Either the PASPORT Fast Response Temperature Sensor or the PASPORT Stainless Steel Temperature Sensor can be used for this activity.

Lab Skills

1. Significant Figures

Objectives

Determine the correct number of significant figures to include when reporting a measurement or a calculated value based upon measurements. Through this investigation, students:

- Explain the difference between precision and accuracy
- Rank the precision of different instruments
- Record the values of length and volume measurements to the proper number of significant figures
- Record the result of a calculation utilizing measurements to the proper number of significant figures

Procedural Overview

Students conduct the following procedures:

- Use meter sticks with various scales to record the dimensions of different objects to the proper number of significant figures
- Use recorded measurements to calculate volumes (multiplication and addition) and report the results with the correct number of significant figures

Time Requirement

♦ Preparation time	15 minutes
♦ Pre-lab discussion and activity	45 minutes
♦ Lab activity	45 minutes

Materials and Equipment

For each group:

- Four-scale meter stick
- Graduated cylinder, 100-mL, partially filled with water
- Graduated cylinder, 10-mL, partially filled with water
- Beaker, 100-mL, partially filled with water
- Irregular-shaped object
- Regular-shaped object

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Units of measure for quantities such as volume, mass, and length
- ♦ Metric-to-metric unit conversions

Related Labs in This Guide

Since the experiments throughout this guide require that data be collected and calculations be made using significant figures, all labs in this guide are related to this one.

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Note: There are no Tech Tips to list in this section as this activity does not use a data collection system.

Background

One of the foundations of science is the collection of data by properly recording measurements. In order for a reported measurement to be useful, it needs to be reliable. Data is reliable if it is both valid and reproducible. The terms accuracy and precision are used when discussing the reliability of scientific data.

Accuracy refers to how closely a measured or calculated value agrees with an accepted value. Accurate measurements can only be recorded if the instrument used has been properly calibrated.

Precision refers to how closely individual measurements or calculations of the same item agree with each other. In other words, it refers to the ability to reproduce the same answer each time the same measurement is made. The precision of a value depends on the instrument being used and can be expressed using significant figures. Significant figures are all of the digits that are known for certain, plus a final estimated digit. This convention automatically indicates the uncertainty in the measurement. It is always important to record scientific results to the correct number of significant figures so that the uncertainty of a measurement will always be known to anyone reviewing the data.

Four rules are used to determine the number of significant figures in a measured quantity:

1. All non-zero digits are significant. (3.42 has three significant figures)

2. All zeros between two non-zero digits are significant. (303.02 has 5 significant figures)

3. All leading zeros that precede the first non-zero digit are never significant. (0.0034 has 2 significant figures) These zeros are there simply to hold place values.

4. Trailing zeros that follow the last non-zero digit are only significant if there is a decimal point in the number. For example, 3400. has 4 significant figures and 3.40 has three, but 3400 only has 2. Here, the decimal point is used to indicate if the zero is simply holding a place value (as in 3400) or whether it was actually a recorded value and has meaning (as in 3400. and 3.40). The difference between 3.40 (with three significant figures) and 3.4 (with two significant figures) is that the hundredths place in 3.40 is known to be exactly zero and not possibly a one or a nine. In 3.4, the value of the digit in the hundredths place is unknown, often because the instrument used to measure the quantity was not precise enough (for example, it lacked markings on the scale) for estimating the hundredths place.

Science often requires the mathematical manipulation of data through addition, subtraction, multiplication, and division. Calculators, as well as data collection systems, often report answers and measurements with every digit that can fit on the screen. Unfortunately, these values are often unrealistic in terms of the number of digits that actually have meaning. For this reason, when performing mathematical operations, additional rules must be followed when expressing the answer to the correct number of significant figures.

When adding and subtracting, the answer should have the same number of decimal places as the least precise measurement (the value with the least number of decimal places). When multiplying and dividing, the answer should have the same number of significant figures as the measurement with the fewest number of significant figures. Often, this requires rounding the result of a calculation to the proper number of significant figures. Although more robust rules for rounding scientific data exist, this guide uses the simpler rules that students are most likely already familiar with:

- If the digit immediately to the right of the one to be rounded is 5 or greater, the value increases.

- If the digit immediately to the right of the one to be rounded is 4 or less, the value remains unchanged.

- The results of intermediate calculations (those to be carried through as part of additional calculations) should not be rounded; any necessary rounding should be performed only on the absolute final result.

Pre-Lab Discussion and Activity

Accuracy versus Precision

Show the following figures representing accuracy and precision, and discuss the difference between accuracy and precision. Explain that the goal of an experimenter is to get all the dots into the very center of the target. Dots representing data points that are closest to the center of the target are closest to the accepted value and have high accuracy. The closer the dots are grouped to each other, the higher the precision of the measurements.

Figure 1

Figure 2

Figure 3

1. Use the terms accurate and precise to describe the results portrayed in Figure 1.

Precision refers to how close the measurements are to each other; accuracy describes how close the data are to the desired, accepted value. The dots are both accurate and precise. They are accurate, because they are all near the center of the bull's-eye (the desired result). They are precise, because they are all close to one another.

2. Use the terms accurate and precise to describe Figure 2.

In Figure 2, the dots are far apart from each other and are not in the center. Therefore, the data is neither precise nor accurate.

3. Use the terms accurate and precise to describe Figure 3.

In Figure 3, the dots are all very close to each other; however, they are not near the center of the target. Even though the dots are precise, they are not accurate.

Accuracy and Precision of Scientific Measurements

To emphasize the importance of taking consistent measurements, have the students take turns measuring the length of the same index card using rulers. Record the values on the board. While discussing the results, the students should agree there exists only one correct answer; this allows for the introduction of the topic of accuracy. Introduce the concept of precision by having the students discuss the number of decimal places the rulers were able to produce.

The accuracy and precision of data collected during an experiment depends on many factors including the procedure followed, the experimenter's technique, and the precision of the instruments used to collect the data. To assess an experiment's accuracy and precision, statistics are used. The accuracy of a result can only be determined if the average experimental value from replicate data can be compared to an accepted value. Percent error is often used to quantify a result's accuracy.

$$\text{percent error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100$$

The precision of a result is determined by the standard deviation of the average value. The smaller the standard deviation, the better the precision. The average (μ) is found by:

$$\mu = \frac{x_1 + x_2 + \ldots + x_{N-1} + x_N}{N},$$

where x is the value of an individual result and N is the number of replicates. If desired, the formula for standard deviation can be given. The standard deviation (σ) is found by:

$$\sigma = \sqrt{\frac{(x_1 - \mu)^2 + (x_2 - \mu)^2 + \ldots + (x_{N-1} - \mu)^2 + (x_N - \mu)^2}{N}},$$

where μ is the average value, x is the value of an individual result, and N is the number of replicates. The following question may be added: What is the standard deviation for the length of the index card? The standard deviation of the length of the index card to the hundredths place using the sample data below is 0.02 cm.

4. What is the length of the index card?

Values will vary depending on the size of the card and the precision of the ruler used. Example values are: 12.66 cm, 12.64 cm, 12.65 cm, and 12.68 cm.

5. If we are all measuring the same card, how many answers should there be?

One.

6. What term is used to refer to the idea that there is only one correct or accepted value?

Accuracy.

7. What is the precision of the rulers used to measure the index card?

Precision is indicated by the number of decimals places recorded. Because the rulers were able to measure to the hundredths of a centimeter, the rulers had a precision of 0.01 cm.

8. What is the average result for the length of the index card?

(12.66 cm + 12.64 cm + 12.65 cm + 12.68 cm) = 50.63 cm / 4 = 12.6575 cm → 12.66 cm

The average length of the index card using the example values to the hundredths place is 12.66 cm.

9. If the accepted or true value of the length of the index card is 12.65 cm, what is the percent error of the experimental result?

$$\textbf{percent error} = \frac{|\textbf{accepted value} - \textbf{experimental value}|}{\textbf{accepted value}} \times \mathbf{100}$$

$$\text{percent error} = \frac{|12.65\text{ cm} - 12.66\text{ cm}|}{12.65\text{ cm}} \times 100 = 0.08\%$$

Accuracy and Calibration

For this activity, one half of the class will use an 80-cm stick scaled with a 100-division label and the other half will use Side C of the four-scale meter stick (1 mm precision) to measure an object (for example, a textbook). This will allow half of the class to obtain precise but inaccurate measurements and the other half to obtain measurements that are both precise and accurate. Place a number line on the board and have the students record their answer on a data pointer and then position it on the number line. After all the groups have submitted their answers, write the true value for the length of the object on a data pointer and add it to the number line.

Teachers Tip: All the items required for this activity are available with the PASCO Significant Figures Set, or individually: Meter Stick Label 80 cm/100 div, Four-Scale Meter Stick, Number Line, and Data Pointers.

10. What is the length of the chemistry textbook? Record the result on a data pointer and place it on the number line

Samples results: 35.1 cm, 35 cm, 28.2 cm, 34.9 cm, 28.3 cm, 28.5 cm, 35.2 cm, 28.2 cm

In this example, the correct value will be taken as 28.2 cm. The results near 35 cm are from the improperly scaled meter stick (80-cm stick with 100 divisions).

11. How many correct answers should there be?

There should be only one correct answer because the same book is being measured.

12. Look at the measurements on the data pointers posted on the number line. Which measurements were precise?

Precise measurements will be clustered closely together with other measurements. Data pointers that are isolated by themselves are not precise. Answers will vary by class.

13. Which measurements on the number line were accurate?

Accurate measurements will be close to the data pointer with the correct value. Answers will vary by class.

14. Can results be precise but not accurate? Explain. Which measurements on the number line are precise but not accurate?

The groups measuring with the shorter meter stick label should give results that are precise (clustered together) but not accurate (not near the data pointer with the correct value).

Precision and Significant Figures

The number of significant figures that can be used in a measurement depends on the number of divisions on the scale of the measuring device being used.

Use magnified pictures of two different graduated cylinders, similar to the ones below, to demonstrate how to take measurements with the correct number of significant figures. Including both a 100-mL and a 10-mL graduated cylinder will produce different levels of precision based on the number of divisions on their scales. All the digits that are known for certain (the marked divisions) plus one estimated digit (between the two smallest divisions) are significant. Because of attractions between the glass and the water molecules, the surface of the water inside the graduated cylinder is curved. This curve is called a meniscus. The meniscus can extend across many division lines on the scale of the cylinder, so the measurement is read at the meniscus's lowest point (the bottom).

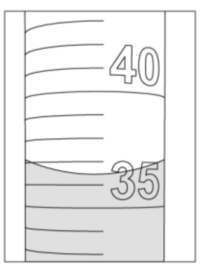


Figure 4: Graduated cylinder, 100-mL

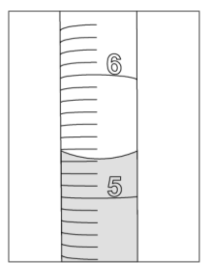


Figure 5: Graduated cylinder,10-mL

15. When using a given piece of equipment, how can the number of significant figures for a measurement be determined?

Only the significant figures should be recorded. Significant figures include all the digits that are known for certain plus one estimated digit.

16. On a graduated cylinder made from glass, the liquid may form a curved line instead of a line straight across the cylinder. What is this curve called and where should the measurement be taken?

The curve is called the meniscus, and the measurement should be taken at the bottom of the meniscus.

17. What is the volume of the liquid in Figure 4?

It is certain that the meniscus is above the 36-mL mark (the certain digits), but it is not known exactly how much beyond the mark. Because it appears approximately halfway, the final digit must be estimated as 0.5, giving a complete reading of 36.5 mL. Other acceptable answers would be 36.4 mL, or 36.6 mL.

18. What is the volume of the liquid in Figure 5?

It is certain that the meniscus is above the 5.3-mL mark (the certain digits), but it is not known exactly how much beyond the mark. Because it appears only slightly above the mark, the final digit can be estimated as 0.02; the volume should be recorded as 5.32 mL.

19. Which of the two graduated cylinders is the more precise? Explain.

The 10-mL graduated cylinder in Figure 5 is the more precise. The divisions on this cylinder mark every tenth of a milliliter, whereas the divisions on the larger graduated cylinder only mark every whole milliliter.

20. Which digit in each of the measurements is the least reliable? Explain.

The last digit in each measurement is the least reliable because it had to be estimated.

Calculations with Significant Figures

Demonstrate the need for significant figures as they apply to mathematical calculations. Carefully measure 50.0 mL of water into a graduated cylinder. Also, fill a glass jar with no measurement markings with approximately 50 mL of water (not measured). Explain that the water in the graduated cylinder is known to a greater level of precision than that in the glass jar. Add the water from the cylinder to that in the glass jar. Discuss the total amount of water and the precision of the known volume now in the jar. Explain the rules for determining the number of significant figures to be reported in the result of a mathematical calculation. Demonstrate rounding a calculated value to the proper number of significant figures.

21. Is the fifty milliliters in the graduated cylinder or the fifty milliliters in the glass jar more precise? Explain.

The 50.0 mL in the graduated cylinder is more precise because the graduated cylinder is marked with lines that indicate each milliliter, whereas the jar has no divisions at all.

22. After adding the 50.0 mL from the graduated cylinder to the glass jar, how much water is in the jar?

Even though the water from the graduated cylinder had a more precise measurement, the final volume cannot be as precise because one of the measurements was not (the initial volume of water in the glass jar). The final answer can never be more precise than the least precise measurement (measurement with the fewest number of known decimal places).

50.0 mL	(graduated cylinder, tenths place known, 3 significant figures)
50 mL	(glass jar, tens place known, 1 significant figure
100 mL	(total, tens place known, 2 significant figures)

23. What is the answer, to the correct number of significant figures, when adding 12.11, 18.0, and 1.013? Explain your reasoning.

12.11

+18.0 least precise measurement has one decimal place

<u>1.013</u>

31.123 → 31.1

The answer should only have one decimal place, because the final answer can only be reported to the precision of the least precise measurement.

24. What is the answer, to the correct number of significant digits, when multiplying 4.56 by 1.4?

When multiplying and dividing, the answer should have the same number of significant figures as the measurement with the fewest number of significant digits.

4.56 (three significant figures) x 1.4 (two significant figures) = 6.384, which should be rounded to 6.4 to have two significant figures.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. Part 1 requires an object of irregular shape. Note the length of the object for later assessment of student work.
2. Part 2 requires an object of a regular geometric shape. Note the length, width, and height of the object for later assessment of student work.
3. Part 3 requires a 100-mL beaker, a 100-mL graduated cylinder labeled “cylinder 1”, and a 10-mL graduated cylinder labeled “cylinder 2”. Fill each with various volumes of water, noting the individual volumes for later assessment of student work. Food coloring may be added to the water to assist the students in seeing the volume easier.

Teacher Tip: For the most efficient use of time and equipment, the class should be divided into small groups and cycled through the experimental procedure set up as three separate stations throughout the laboratory. Replicate stations for each part can be created for larger class sizes. Make sure that each station will produce identical results (objects are of the exact same dimensions and water volumes are exactly equivalent).

Safety

Follow all standard laboratory procedures

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	3	1	5	4
Measure the object and record all of the digits that are known for certain based on the divisions on your measuring device.	Record the last digit by estimating where the object being measured falls between two divisions on your measuring device.	Determine the scale on the measuring device you are using.	Complete the necessary calculations and report the calculated value using the correct number of significant figures.	Continue by recording all additional the measurements using the correct number of significant figures.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Collect Data

Part 1 – Precision of Instruments

1. ☐ Measure the length of the irregular shaped-object provided using each side of the four-scale meter stick: Side A has the largest divisions; Side D, the smallest. Use the proper number of significant figures, remembering to estimate your final digit. Include the proper units for each measurement. Record your results in Table 1 below.

Object measured: Xplorer GLX®

Table 1: Irregular-shaped object's measurements

Length Measured with Side A	Length Measured with Side B	Length Measured with Side C	Length Measured with Side D
0.3 m	2.3 dm	22.6 cm	22.35 cm

2. ☐ What is the value of the divisions on each side of the four-scale meter stick? Record your answers in Table 2 below.

Table 2: Four-scale meter stick divisions

Side	Size of Divisions
A	1 m
B	1 dm
C	0.5 cm
D	1 mm

Part 2 – Volume Calculations with Significant Figures

3. ☐ Measure the length of the object using side B of the four-scale meter stick. Record the length using the correct number of significant figures in Table 3.

4. ☐ Measure the width of the object using side C of the four-scale meter stick. Record the width using the correct number of significant figures in Table 3.

5. ☐ Measure the height of the object using side D of the four-scale meter stick Record the height using the correct number of significant figures in Table 3.

Object measured: _____cardboard box_______________________

Table 3: Regular-shaped object's measurements

Length (Side B of meter stick)	Width (Side C of meter stick)	Height (Side D of meter stick)
31 cm	45.3 cm	61.32 cm

Part 3 – Addition Problems with Significant Figures

6. ☐ Record the volume of the liquid in the beaker in Table 4 using the correct number of significant figures.

Table 4: Volume of liquid in the beaker

Beaker Volume	Cylinder 1 Volume	Cylinder 2 Volume
61 mL	32.1 mL	83.23 mL

7. ☐ Look at the liquid in the graduated cylinders and notice the curve on the surface of the liquid. This is the meniscus. Why does the water curve upward towards the sides of the glass? Should you measure from the top or the bottom of the meniscus?

The water curves upward towards the sides of the glass because the water molecules are attracted to the glass. The liquid should be measured from the bottom of the meniscus.

8. ☐ Measure the volume of the liquid in cylinder 1 and record the volume in Table 4 using the correct number of significant figures.

9. ☐ Measure the volume of the liquid in cylinder 2 and record the volume in Table 4 using the correct number of significant figures.

10. ☐ Clean-up your lab station according to the teacher's instructions.

Data Analysis

Part 1 – Precision of Instruments

1. ☐ Convert all the irregular-shaped object's measurements to centimeters and record them in Table 5.

Table 5: Irregular-shaped object's measurements in centimeters

Side of Ruler Measuring the Object	Show Your Work Converting to cm	Length (cm)
Side A	$0.3\ \text{m} \times \frac{100\ \text{cm}}{1\ \text{m}} = 30\ \text{cm}$	30
Side B	$2.3\ \text{dm} \times \frac{10\ \text{cm}}{1\ \text{dm}} = 23\ \text{cm}$	23
Side C	Conversion not needed	22.6
Side D	Conversion not needed	22.35

2. ☐ Record this data (Group 1) as well as the data collected by two other groups in Table 6.

Table 6: Irregular-shaped object's measurements collected by three different groups

Group	Side A of Meter Stick (cm)	Side B of Meter Stick (cm)	Side C of Meter Stick (cm)	Side D of Meter Stick (cm)
1	30	23	22.6	22.35
2	20	22	22.4	22.41
3	20	24	22.5	22.39

3. ☐ When given a group of data values, how can you determine if the data is precise?

Precise data will produce the same values every time the same measurement is taken. The closer the values are to each other, the more precise the data.

4. ☐ Which side of the meter stick allowed for the greatest precision? Explain.

Looking at the data table, the measurements taken with Side D of the meter stick were clustered closer together than the measurements taken with the other sides.

5. ☐ Which side of the meter stick showed the least amount of precision? Explain.

Looking at the data table, the measurements taken with Side A of the meter stick were the farthest apart in value.

6. ☐ Rank the sides of the meter stick in order of least to greatest precision.

Side A < Side B < Side C < Side D.

Part 2 – Volume Calculations with Significant Figures

7. ☐ Convert all the regular-shaped object's measurements to centimeters with the correct 3number of significant figures and record them in Table 7 (as Group 1).

8. ☐ Enter the data collected by two other lab groups in Table 7.

Table 7: Regular-shaped object's measurements and calculated volume

Group #	Length: Side B of Meter Stick (cm)	Width: Side C of Meter Stick (cm)	Height: Side D of Meter Stick (cm)	Volume of Object (cm^3)
1	31	45.3	61.32	86000
2	61	32.4	45.28	89000
3	44	61.4	32.08	87000

9. ☐ How can the volume of a regular-shaped object be calculated?

volume = length x height x width

10. ☐ Calculate the volume of the object with the data collected from each lab group. Record the answer in Table 7. Be sure to use the correct number of significant figures.

31 cm x 45.3 cm x 61.32 cm = 86111.676 cm^3 → 86000 cm^3

11. ☐ Explain how the number of significant figures was decided when recording the volume.

Since calculating volume requires multiplication, the answer is limited to the number of significant figures in the number with the least number of significant figures. The length, which was measured with side B, has only two significant figures; thus, the answer could only be reported with two significant figures.

Part 3 – Addition Problems with Significant Figures

12. ☐ Without actually combining the contents of the glassware, mathematically add the recorded measurements to produce a result that represents the total amount of liquid present in all three containers taken together. Record the value with the correct number of significant figures in Table 8 (as Group 1).

Table 8: Total volume of liquid

Group #	Beaker Volume (mL)	Cylinder 1 Volume (mL)	Cylinder 2 Volume (mL)	Total Volume (mL)
1	61	32.1	83.23	176
2	62	32.3	83.22	178
3	61	32.4	83.23	177

13. ☐ Explain how the number of significant figures was decided when recording the total volume.

Since combining volume requires addition, the sum must be recorded using the same number of decimal places as the measurement with the fewest decimal places (least precision). The volume in the beaker could only be measured to whole milliliters; thus, the answer could only be reported with the same precision (whole milliliters).

14. ☐ Collect the volumes recorded from two other lab groups and record them in Table 8.

15. ☐ Which of the three pieces of glassware provided the most precise measurement? Was this precision seen in the final volume?

The 10-mL graduated cylinder (Cylinder 2) provided the most precise measurement (0.01 mL). This precision was not seen in the final volume, because the final volume was limited by the volume in the beaker which was the least precise (1 mL).

Analysis Questions

1. Do significant figures relate to the accuracy or the precision of the measurement?

Significant figures are closely related to the precision of a measurement because it reports the uncertainty in a measurement. Accuracy is how close a measurement is to its true/accepted value which will depend on the calibration of the measuring device.

2. Explain the reasoning behind the rules for adding, subtracting, multiplying, and dividing with significant figures.

Significant figures reflect the amount of uncertainty in a measurement. When two or more numbers are combined in a mathematical operation, the uncertainty in the least precise measurement will carry over into the final answer. The final answer must display the same amount of uncertainty as the least precise measurement.

3. What determines the number of significant figures in a recorded value?

The number of significant figures is determined by the precision of the measuring device used to make the measurement.

4. What determines the number of significant figures in a calculated value?

If the result is from a multiplication or a division calculation, then the number of significant figures in the answer will be the same as the value with the least number of significant figures.

If the result is from an addition or a subtraction calculation, then the number of decimal places in the answer (precision of the answer) will be the same as the value with the least number of decimal places (least precise measurement).

Synthesis Questions

Use available resources to help you answer the following questions.

1. The density of copper is listed as 8.94 g/mL. Two students each make three density determinations through experimentation. Student A's measurements are 6.3 g/mL, 8.9 g/mL, and 11.1 g/mL. Student B's measurements are 8.3 g/mL, 8.2 g/mL, and 8.4 g/mL. Compare the two sets of results in terms of precision and accuracy.

Student A displayed less precision but more accuracy (due to the one measurement that was very close to the accepted density). Student B displayed less accuracy but more precision (due to having all three measurements fairly close to one another).

2. Five different students take the following measurements of the same object: 1.3 m, 1.5 m, 1.45 m, 1.47 m, and 1.453 m. Why are the measurements different? Which measurement is correct?

The measurements are different because the students used different instruments of varying precision to measure the object. All the measurements may be recorded "correctly" according to the degree of precision of each instrument. The instrument that gave the value of 1.453 however, offers the greatest amount of precision.

3. A student reported finding the mass of an object to be 350 grams. How many significant figures are in this number and which digit has uncertainty?

There are two significant digits in this number (3 and 5). There is uncertainty in the tens digit (5) because it was estimated. The zero is simply holding the place value and is not significant.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Which of the following numbers does NOT have 2 significant figures?

- **A.** 2300
- **B.** 0.000030
- **C.** **51.0**
- **D.** 30.

2. Using the rules of significant figures, calculate the following: (6.167 + 83) / 5.10

- **A.** 17.48
- **B.** **17**
- **C.** 17.5
- **D.** 20

3. The amount of uncertainty in a measured quantity is determined by:

- **A.** The skill of the observer only
- **B.** Neither the skill of the observer nor the limitations of the measuring instrument
- **C.** The limitations of the measuring instrument only
- **D.** **Both the skill of the observer and the limitations of the measuring device**

4. How many significant figures are there in 0.0503 grams?

- **A.** 5
- **B.** 4
- **C.** **3**
- **D.** 2

5. If you need exactly 7.00 mL, which measuring device would you recommend?

- **A.** A 50-mL beaker
- **B.** A 50-mL graduated cylinder
- **C.** **A 10-mL graduated cylinder**
- **D.** A 100-mL graduated cylinder

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. When collecting data for an experiment, it is important to note certain qualities of that data. The **accuracy** of the data is a measure of how close the results are to an expected or accepted true value. The **precision** of the data is how close the results are to each other and is a measure of the repeatability of the results. The precision of an instrument is reported by using **significant figures**; these consist of all the digits of a measurement that are **known** for certain plus one **estimated** digit.

2. To determine the number of significant figures in a measurement, a set of rules is followed. All **non-zero** digits are significant. Zeroes between non-zero digits **are** significant. Leading zeroes before non-zero digits **are not** significant. Zeroes that end a measurement are significant only if there is a **decimal point** in the number.

3. Knowing how many significant figures are in a number is important because they are used when **measurements** are used in calculations. In **multiplication** and **division**, the number of significant figures depends on the measurement with the **fewest** number of significant figures. In **addition** and **subtraction**, the number of digits depends on the number of **decimal places** in the **least** precise number used in the calculation. To report an answer with the correct number of significant figures often requires the final answer to be **rounded**; digits **five** or greater will **be rounded up**, while **four** or less will **remain unchanged**.

Extended Inquiry Suggestions

Measure the diameter and height of a beaker. Calculate the circumference, area, and volume of the beaker using proper significant figures.

Include mass as a quantity to be measured. Use various balances with different levels of precision.

Using a centigram balance (precision of 0.01 g), measure the weight of the same amount of water transferred using various pieces of glassware (beaker, 10-mL graduated cylinder, 50-mL graduated cylinder). Plot the results on a number line to further demonstrate the precision and accuracy of the different pieces of glassware.

2. Density

Objectives

Determine that density is an intensive property of a substance independent of the shape or size of an object. Through this investigation, students:

- Determine the volume of regular- and irregular-shaped objects using geometric calculations and water displacement methods
- Use mass and volume data to calculate density using the formula, $\text{density} = \frac{\text{mass}}{\text{volume}}$
- Distinguish between intensive and extensive properties
- Learn that density is an intensive physical property that can be used to identify unknown substances

Procedural Overview

Students conduct the following procedures:

- Determine the volume of regular-shaped objects through geometric calculation
- Determine the volume of irregular-shaped objects through water displacement
- Measure the mass and volume of various objects and calculate the density by dividing the two values
- Identify the material a plastic cylinder is made from when given a list of substances and their corresponding densities

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	25 minutes
♦ Lab activity	30 minutes

Materials and Equipment

For each student or group:

- PASCO density set
- Beaker, 150-mL
- Graduated cylinder, 50- or 100-mL
- Balance (2 to 3 per class)
- Overflow can
- Metric ruler (or calipers)
- Water, 500 mL
- String

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Mass measurements
- ♦ Volume measurements
- ♦ Length measurements
- ♦ Physical properties
- ♦ Geometric mathematical formula for a cube and a cylinder

Related Labs in This Guide

Labs conceptually related to this one include:

- ♦ Significant Figures
- ♦ Graphing Mass versus Volume to Determine Density

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Note: There are no Tech Tips to list in this section as this activity does not use a data collection system.

Background

A substance can often be identified by its physical properties. For example, a substance that melts at 0ºC and boils at 100ºC might very well be water, especially if that substance is a clear, colorless, odorless liquid at room temperature. Boiling point, melting point, color, and odor are examples of intensive properties. Intensive properties are those that are independent of how much of the substance is being measured. All water, from a small amount contained in a glass to a large amount contained in a swimming pool or a fresh water lake, freezes at 0°C and boils at 100ºC.

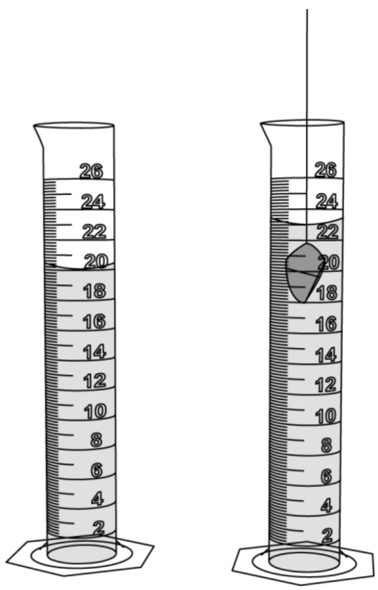
Some properties, however, depend on the amount of substance present. An extensive property changes as the amount of the substance being measured changes. Mass (weight) and volume are both examples of extensive properties. The water in a glass takes up less space and is not as heavy as the water in a swimming pool, even though both contain the same chemical substance.

Objects of the same size might not necessarily have the same mass. For example, an object made from foam weighs less than a lead object of the same size. Density is an intensive physical property of a substance that relates an object's mass to its volume. It is not an extensive property dependent on size because as a sample's mass increases, the sample's volume also increases proportionately. Many substances have similar densities, so density should be used along with other properties to positively identify a substance.

An object's density (ρ, "rho") determines whether it floats in a particular substance. To float, the density of the substance must be less than the density of the fluid in which the object is placed. At room temperature and to two significant figures, water has a density of 1.0 g/mL. This means that objects with densities less than 1.0 g/mL float in water. The fluids are not limited to liquids. For example, helium (ρ_{He} = 0.18 g/L) floats when trapped in a balloon because the density of the surrounding air is much greater (ρ_{Air} = 1.19 g/L).

To calculate density, both the mass and volume of an object must be known. Mass can be found directly using a balance (triple-beam or electronic). To calculate an object's volume, you have two options: 1) If the object has a regular shape, such as a cube or a cylinder, the volume can be calculated using the corresponding mathematical formula for that shape. 2) If the object has an irregular shape, the volume must be determined through other methods, such as water displacement.

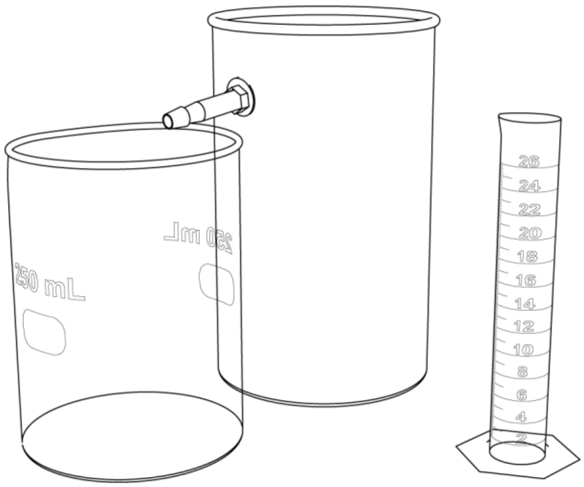
Based on the fact that two objects cannot occupy the same space at the same time, an object that sinks in water displaces a volume of liquid that is equal to its own volume. By submerging an irregular-shaped object in water, it is possible to deduce its volume by measuring the increase of the water level.



In this diagram a small rock caused the water volume to increase from 20 mL to 23 mL. The volume of the rock, therefore, is 3 mL.

Measuring volume of an irregular-shaped object using water displacement.

If the object does not fit into a graduated cylinder, the object can be placed into an overflow container. The water spills out of the container and into a collection beaker. The volume of the water collected in the beaker can then be measured using a graduated cylinder.



Overflow container with a beaker and graduated cylinder.

When the mass and volume have been determined, the density can then be calculated by dividing the two quantities.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Pre-Lab Discussion and Activity

Intensive and Extensive Properties

Use two bottles of soda (same brand and flavor) to demonstrate intensive and extensive properties. Show a small bottle (500-mL) next to a large bottle (2-L). Explain that extensive properties depend upon the amount of substance being measured—larger quantities produce greater values for the extensive property. Intensive properties, however, do not depend upon how much substance is being measured—the value of the property remains constant regardless of how large the quantity.

1. Look at the contents in the two bottles of soda. Are they the same?

If the brand and flavor are the same, the contents should be identical in all regards: taste, sweetness, color, carbonation, and so on.

2. Does the amount of soda you have matter when you decide how large a bottle should be used to contain it?

Yes. The more soda you have, the larger the bottle needs to be.

3. Is volume an intensive or extensive property of soda?

Because the volume increases as you increase the amount of soda you have, volume is an extensive property.

4. Think about the taste of the soda inside the bottles. Does the flavor of the soda change depending on the amount of soda you drink?

No. The taste remains the same no matter how much you pour into a glass.

5. Is flavor an intensive or extensive property of soda?

The taste remains the same no matter how much you drink; therefore flavor is an intensive property.

Density and Archimedes

Relate the following story of Archimedes: Around 250 B.C., King Hiero II wanted a new crown. He gave a block of gold to a goldsmith to use for the crown. The king was suspicious that the goldsmith might keep some of the gold and substitute a less expensive metal, coating only the outside of the crown with gold. The king asked the famous Greek philosopher, Archimedes, to determine if the crown was gold without damaging it. Archimedes knew if it were made of pure gold, the crown would have the same intensive properties as a bar of gold. Archimedes thought long and hard about the problem knowing he couldn't scratch the surface of the crown or melt any part of it. One day, as he stepped into his bathtub, the water spilled over the edge of the bathtub and onto the floor. This helped Archimedes figure out the solution to his problem. He was so excited he jumped out of the tub and forgot to get dressed! He ran through the streets without his clothes shouting "Eureka!" ("I have found it!") on his way to tell the king his idea.

Simulate Archimedes' problem by wrapping same sized blocks of aluminum, iron, and lead with masking tape (or paint these blocks with gold-colored paint). The masking tape or paint represents a thin layer of gold coating the less expensive metal inside. Inform the students they are not allowed to scratch through the surface coating.

6. What is the same about the blocks?

The blocks have the same size (volume) and they have the same outside appearance.

7. What is different about the blocks?

The blocks have different masses (weigh different amounts).

8. Why was the overflowing bathtub important to Archimedes' problem?

The amount of water that overflowed was equal to his volume. He reasoned that the crown's volume could be determined the same way. Finding the crown's mass was straightforward. By finding the ratio of the mass and the volume of the crown, Archimedes could compare the densities of the crown with a sample of real gold.

Regular Soda and Diet Soda

Use four bottles of soda (500-mL and 2-L of soda along with 500-mL and 2-L of diet soda) to demonstrate the concept of density. Record the weights of the different bottles and assist the students in calculating density by dividing the masses by their respective volumes. Record and display the results. Explain that for an object to float in water, it must have a density less than water (ρ_{water} = 1.0 g/mL). Ask the students to predict what will happen to each bottle (float or sink) as it is being placing into a large container of water, such an aquarium. Discuss the results by discussing the similarities and differences between the bottles.

Teacher's Tip: Avoid using the terms heavy and light when describing density. Instead, the terms more dense and less dense are more appropriate. Remember: "What weighs more, a kilogram of lead or a kilogram of feathers?" (Neither: they are both one kilogram of material.) Even though aluminum is not very dense (ρ_{Al} = 2.70 g/cm^3), a 500-kilogram block of aluminum is still very heavy. (It weighs 500 kilograms.)

9. How are the bottles of soda similar?

The bottles of soda all contain a carbonated liquid soda beverage.

10. How are the bottles of soda different?

The bottles of soda are different sizes and contain either regular or diet soda.

11. How is diet soda different from regular soda?

Regular soda contains more calories than diet soda. Regular soda contains sugar while the diet soda contains different types of sweeteners.

12. Which bottles float in water? Why?

Both bottles containing diet soda float. The diet soda has a density less than water. The regular soda sinks in water because its density is greater than water. The difference is in the sweeteners that are used in the soda. Regular soda generally contains between 35 to 45 grams of sugar, but diet soda only contains about 0.1 to 0.2 grams of artificial sweeteners. The regular soda has a much greater amount of sugar within the same volume of liquid, making it more dense.

13. The 2-L bottle of diet soda weighs more than the 500-mL bottle of regular soda, but it still floats while the other sinks. Explain.

Density is an intensive property, independent of the amount of substance present. Density is the ratio of mass to volume; as a sample's mass increases, its volume also increases proportionately to keep density constant for that substance.

Determining Volume

Hold up a box, a cylinder, and an irregular-shaped object (such as a rock) and ask how the volume of these objects could be found. Review the mathematical formulas for calculating volume of regular-shaped objects and demonstrate how to calculate the volume of an irregular-shaped object. Introduce the water displacement method by dropping a rock into an overflow can and then measuring the displaced water using a graduated cylinder.

14. How can the volume of the box be determined?

Measure the length, width, and height of the box and multiple the three measurements together.

Volume of a box = length × width × height

15. How can the volume of the cylinder be determined?

Measure the height of the cylinder as well as the diameter, then divide the diameter by two to find the radius.

Volume of a cylinder = height × πr^2

16. How can the volume of the rock be determined?

Because the rock is an irregular-shaped object, its volume is difficult to determine mathematically. It is much easier to determine the volume of the rock using water displacement. Place the rock in a container and measure the volume of water that is displaced. The volume of water displaced is equal to the volume of the rock.

17. Can the volume of regular-shaped objects be determined using water displacement?

Yes. The volume determined from water displacement should be the same as the volume calculated using the mathematical formulas.

Lab Preparation

Although this activity requires no specific lab preparation, allow 10 minutes to gather the equipment needed to conduct the lab.

Safety

Follow all standard laboratory practices.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

1	5	2	4	3
Before taking any quantitative measurements, list some qualitative properties of the object being studied.	Identify the object's material based on its density.	Measure the object's dimensions and mass.	Determine the object's density.	Calculate the object's volume.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Collect Data

Part 1 – Brass Objects

1. ☐ List at least two qualitative observations about the brass objects.

Brass is golden in color and has luster (shiny).

2. ☐ Predict how the density of the brass block compares to the density of the brass cylinder.

Because density is an intensive property of the substance, the density of the block should be the same as the density of the cylinder because both shapes are made of brass.

3. ☐ Measure the length, width, height, and mass of the brass block and the height, diameter, and mass of the brass cylinder. Record your results in Table 1 below.

Table 1: Dimensions and mass of brass objects

Object	Length (cm)	Width (cm)	Height (cm)	Diameter (cm)	Mass (g)
Brass block	2.59	1.88	1.58		66.10
Brass cylinder			6.38	2.20	208.16

Part 2 – Aluminum Objects

4. ☐ List at least two qualitative observations about the aluminum objects.

Aluminum is silver in color and has luster (shiny).

5. ☐ Predict how the densities of the three aluminum objects will compare to each other.

Density is an intensive property of the substance, therefore the densities of the three objects should be the same because all three objects are made of aluminum.

6. ☐ Measure the length, width, height and mass of the aluminum block and the height, diameter, and mass of the aluminum cylinder and record your results in Table 2 below.

Table 2: Dimensions and mass of aluminum objects

Object	Length (cm)	Width (cm)	Height (cm)	Diameter (cm)	Mass (g)
Aluminum block	4.88	3.12	1.55		64.85
Aluminum cylinder			6.35	2.20	66.44

7. ☐ Measure the mass of the irregular-shaped aluminum object.

Mass of irregular-shaped aluminum object: __________66.54 g _______

8. ☐ Complete the following steps to measure the volume of the irregular-shaped aluminum object using water displacement.

a. Put the beaker under the overflow can spout.

b. Pour water into the overflow can until it overflows into the beaker.

c. Allow the water to stop overflowing on its own and empty the beaker into the sink.

d. Place the beaker back in its position under the overflow can spout without touching the overflow can.

e. Tie a string to the irregular-shaped object and gently lower the object into the overflow can until it is completely submerged.

f. Allow the water to stop overflowing and then pour the water from the beaker into the graduated cylinder.

g. Measure the volume that was displaced by reading the water level in the graduated cylinder.

h. Record the volume of water that was displaced in units of cm^3 (1 mL = 1 cm^3).

Volume of water displaced: ___________25.3 mL ___ ___

9. ☐ Why do you need to use the water displacement method for the irregular-shaped object?

There is no mathematical formula for the volume of an irregularly-shaped object, so you must find the volume through other means.

Part 3 – Unknown Plastic Objects

10. ☐ List at least two qualitative observations about the plastic cylinder.

The plastic cylinder is white in color and is opaque.

11. ☐ Table 3 lists three common plastics and their densities. How might you determine the material that the plastic cylinder is made?

Table 3: Density of plastics

Types of Plastic	Density
Polypropylene	0.95 g/cm^3
Nylon	1.15 g/cm^3
Polyvinyl chloride	1.39 g/cm^3

Measure the volume and mass of the cylinder and then calculate its density. If the cylinder is made of one of these three plastics, its density should match one of the three densities given.

12. ☐ Measure the height, diameter, and mass of the plastic cylinder and record your results in Table 4 below.

Table 4: Dimensions and mass of a plastic cylinder

Object	Height (cm)	Diameter (cm)	Mass (g)
Plastic cylinder	6.35	2.20	23.43

13. ☐ Clean up your lab station according to the teacher's instructions.

Data Analysis

Part 1 – Brass Objects

1. ☐ Use the following equations to calculate the volumes of the brass block and brass cylinder. Show your work and record your results in Table 5 below.

Volume (block) = length × width × height   Volume (cylinder) = height × πr^2

Table 5: Volume of brass objects

Object	Show Your Work Here	Volume
Brass block	Vol = 2.59 cm × 1.88 cm × 1.58 cm =	7.69 cm^3
Brass cylinder	Vol = 6.38 cm × 3.1416 × $(1.10\ cm)^2$ =	24.3 cm^3

2. ☐ Use the following equation to calculate the densities of the brass block and brass cylinder. Show your work and record your results in Table 6 below.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Table 6: Density of brass objects

Object	Show Your Work Here	Density
Brass block	Density = 66.10 g / 7.69 cm^3	8.60 g/cm^3
Brass cylinder	Density = 208.16 g / 24.3 cm^3	8.57 g/cm^3

3. ☐ Did the shape of the brass object have an effect on the resulting density?

No. The densities of the two brass objects were nearly identical.

Part 2 – Aluminum Objects

4. ☐ Calculate the volumes of the aluminum block and the aluminum cylinder. Show your work and record your results in Table 7 below.

Table 7: Volume of aluminum objects

Object	Show Your Work Here	Volume
Aluminum block	Vol = 4.88 cm × 3.12 cm × 1.55 cm =	23.6 cm^3
Aluminum cylinder	Vol = 6.35 cm × 3.1416 × $(1.10\ cm)^2$ =	24.1 cm^3

5. ☐ Calculate the density of the aluminum block, aluminum cylinder, and the irregular-shaped aluminum object. Show your work and record your results in Table 8 below.

Table 8: Density of aluminum objects

Object	Show Your Work Here	Density
Aluminum block	Density = 64.85 g / 23.6 cm^3	2.75 g/cm^3
Aluminum cylinder	Density = 66.44 g / 24.1 cm^3	2.76 g/cm^3
Irregular-shaped aluminum object	Density = 66.54 g / 25.3 cm^3	2.63 g/cm^3

6. ☐ Did the shapes of the aluminum objects have an effect on the resulting densities?

No. The densities of the three aluminum objects were nearly identical.

Part 3 – Unknown Plastic

7. ☐ Calculate the volume of the plastic cylinder. Show your work and record your results in Table 9 below.

Table 9: Volume of plastic cylinder

Object	Show your work here	Volume
Plastic cylinder	Vol = 6.35 cm × 3.1416 × $(1.10\ cm)^2$	24.1 cm^3

8. ☐ Calculate the density of the plastic cylinder. Show your work and record your results in Table 10 below.

Table 10: Density of plastic cylinder

Object	Show your work here	Density
Aluminum block	Density = 23.43 g / 24.1 cm^3	0.97 g/cm^3

9. ☐ From which plastic is the cylinder made?

The density of the cylinder is similar to that for polypropylene. Therefore, the plastic cylinder is most likely made from polypropylene.

Analysis Questions

1. Does the shape of an object affect its density?

No. Density is an intensive property and is always constant for the same substance. The shape and size of the object does not affect its density.

2. Is it possible for two objects to have the same volume and different densities? Explain your answer and provide evidence from this experiment to support your answer.

Yes, the three cylinders had essentially the same volume (~24 cm^3), but they had three different densities (ρ_{brass} = 8.6 g/cm^3, ρ_{Al} = 2.7 g/cm^3, $\rho_{plastic}$ = 0.97 g/cm^3). The cylinders had three different densities because they were made from three different substances.

3. Which material, brass, aluminum, or plastic, was the most dense?

The brass objects have the greatest density (ρ = 8.6 g/cm^3), followed by aluminum (ρ = 2.7 g/cm^3), and then plastic (ρ = 0.97 g/cm^3).

4. Research the accepted values for the densities of aluminum and brass. How do the accepted answers compare to the values you calculated in this experiment?

Aluminum has an accepted density of 2.70 g/cm^3. Brass has an accepted density between 8.4 to 8.8 g/cm^3. Brass is an alloy of variable amounts of zinc and copper, which explains the range of possible densities. These values are similar to those found in the lab.

Teacher Tip: Percent error may be calculated using the following formula:

$$\text{Percent error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100$$

Synthesis Questions

Use available resources to help you answer the following questions.

1. Will the brass, aluminum, or plastic cylinder float in water? Explain.

The brass and aluminum cylinders will both sink in water because their densities are greater than the density of water (ρ_{water} = 1.0 g/cm^3). The plastic cylinder will float because its density is less than that of water.

2. If a company buys 200 cm^3 of aluminum, how much would you expect the aluminum to weigh?

The aluminum would weigh 540 g. This value is calculated from the density:

$$200\ \text{cm}^3\left(\frac{2.70\ \text{g}}{1\,\text{cm}^3}\right) = 540\ \text{g Al}$$

3. A 260-kg tree that is 10 m tall and 25 cm in diameter falls into a river. Explain mathematically why the tree floats, given that the density of water is 1000 kg/m³.

The tree has a volume of 0.49 m^3.

$$V_{\text{cylinder}} = h \times \pi r^2$$

$$V_{\text{cylinder}} = 10\text{ m} \times \pi\left(\frac{0.25\text{ m}}{2}\right)^2 = 0.49\text{ m}^3$$

Its density is 530 kg/m^3.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{density} = \frac{260\text{ kg}}{0.49\text{ m}^3} = 530\text{ kg/m}^3$$

The density of the tree is less than the density of the water, therefore the tree will float.

4. Can a very large object have the same density as a very small object? Explain.

Two different-sized objects can have the same density. Density is an intensive property. If the large object also has a large mass and the small object has a small mass, they could have the same density if their mass/volume ratio is the same. A small lead pipe will have the same density as a large lead pipe because it is made out of the same material.

5. A student has three silver cubes. Although the cubes look the same, one is made of zinc, another is made of lead, and third is made of aluminum. How can the student determine the material that was used to make each cube?

One way to identify the material in each cube is to determine the density of each cube by measuring its mass and volume. The calculated values can then be compared to literature values of these metals. Lead has the highest density (ρ_{Pb} = 11.34 g/cm^3) followed by zinc (ρ_{Zn} = 7.14 g/cm^3). The lowest density is aluminum (ρ_{Al} = 2.70 g/cm^3).

6. A rectangular object weighs 2445 g and its density is 12.9 g/cm³. When measured, its height is 7.43 cm and its width is 3.45 cm. How long is the object?

The object is 7.40 cm long.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

$$12.9\text{ g/cm}^3 = \frac{2445\text{ g}}{(7.42\text{ cm})(3.45\text{ cm})(\text{length})}$$

$$\text{length} = \frac{2445\text{ g}}{(7.42\text{ cm})(3.45\text{ cm})\left(\frac{12.9\text{ g}}{\text{cm}^3}\right)} = 7.40\text{ cm}$$

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Diamond has a density of 3.26 g/cm^3. What is the mass of a diamond that has a volume of 0.350 cm^3?

- **A.** 0.107 g
- **B.** **1.14 g**
- **C.** 9.31 g
- **D.** None of the above

2. What is the volume of a sample of liquid mercury that has a mass of 76.2 g, given that the density of mercury is 13.6 g/mL?

- **A.** 0.178 mL
- **B.** **5.60 mL**
- **C.** 1040 mL
- **D.** None of the above

3. Which statement about density is true?

- **A.** Two samples of nickel may have different densities
- **B.** Density is constant for all types of metals
- **C.** The density of a sample depends on its location on Earth
- **D.** **Density is a constant value for all objects made of the same material**

4. A zinc block has a mass of 20 g and a zinc cylinder has a mass of 40 g. How will the density of the two objects compare?

- **A.** The zinc block will be less dense than the zinc cylinder
- **B.** The zinc block will be more dense than the zinc cylinder
- **C.** **The zinc block and the zinc cylinder will have the same density**
- **D.** There is not enough information to answer the question

5. Density equals:

- **A.** **Mass / volume**
- **B.** Volume / mass
- **C.** Mass × volume
- **D.** Length × width × height

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Properties that depend on the amount of material present are called **extensive** properties, and include **mass** and **volume**. Those properties that are independent of the amount of substance being studied are called **intensive** properties, and include color, **boiling point**, and **density**.

2. **Density** is the amount of matter in a particular amount of space. Density is the ratio of **mass** to **volume**. Substances with large densities feel **heavy** for their size. Substances with densities less than 1.0 g/mL **float** in water. To find an object's density, a **balance** is used to determine its mass. Volume is found either by using a **mathematical formula** or by **water displacement**. Density can be used to **identify** a substance.

Extended Inquiry Suggestions

Pennies produced before 1982 are mostly comprised of copper, while pennies produced after 1982 are mostly zinc covered with a thin layer of copper. The two types of pennies, therefore, have different densities. Determine the densities of pre-1982 pennies and post-1982 pennies. Use ten pennies of each type to get accurate measurements. Use the water-displacement method to find the volume of all the pennies of a particular type at the same time.

Repeat the experiment using only the method of water displacement to determine volume.

3. Graphing Mass versus Volume to Determine Density

Objectives

Use multiple mass and volume data to graphically determine the density of a substance. Through this investigation, students:

- Learn the process of graphing data, including scale selection, axis labeling, descriptive titling, and plotting data points
- Determine trends in the data by mathematically describing the lines that best fit the data, particularly linear relationships ($y = mx + b$) and directly proportional relationships ($y = mx$)
- Discover that slope and y-intercept have meaning
- Identify variables as either independent or dependent

Procedural Overview

Students conduct the following procedures:

- Calculate the volume of cubes and cylinders
- Determine the mass of various objects
- Plot both the mass and volume data of various objects on the same graph
- Determine the mathematical relationship between the data

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	30 minutes

Materials and Equipment

For each student or group:

- Four different-sized rectangular aluminum pieces (part of PASCO's Discover Density Set)
- Balance (2 to 3 per class)
- Four different-sized rectangular plastic pieces of the same composition (part of PASCO's Discover Density Set)
- Metric ruler (or calipers)

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Mass
- ♦ Volume
- ♦ Density
- ♦ Geometric mathematical formula for a cube and a cylinder

Related Labs in This Guide

- ♦ Significant Figures
- ♦ Density
- ♦ Boyle's Law

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Note: There are no Tech Tips to list in this section as this activity does not use a data collection system.

Background

Graphically displaying data is an important process in elucidating trends and relationships between variables. Today, computers can automatically produce graphs from data. However, scientists continue to rely on their knowledge and experience to create graphs without the aid of a computer. To sketch meaningful graphs, you need to select proper scales for the axes, accurately plot the data points, including a line of best fit, select and creatively label the axes, and provide a descriptive title for the figure.

The purpose of an experiment is to produce data that will support or invalidate a hypothesis. The selection of variables that will be measured is critical to the ability of an experiment to test a hypothesis. Independent variables are those whose conditions are set by the experimenter. Dependent variables are those whose values depend upon the values of the independent variables. Often, the value of a dependent variable can be influenced by more than one independent variable. For this reason, you should design controlled experiments. Controlled experiments ensure that changes in the value of a dependent variable being measured are due only to one independent variable.

The independent and dependent variable may or may not be related to one another. If there is no relationship the independent variable will have no effect on the dependent variable. If there is a relationship between the independent and dependent variable the relationship will be either linear or non-linear. In a linear relationship as one variable increases the other variable increases or decreases at a constant rate (constant slope). In a non-linear relationship as one variable increases the other variable increases or decreases at a changing rate (slope changes). A graph is made to show the type of relationship that exists between the independent and dependent variable.

To construct a graph, you should first examine the variables that will be related to one another. The independent variable is plotted on the x-axis (the horizontal axis); the dependent variable is plotted on the y-axis (the vertical axis). Both descriptions of the data and the data points' units of measure should be included in the labels of the two axes. Along with naming the axes, you should provide a descriptive title for the overall graph. A good title reflects the purpose of the graph and not just what is being compared because this is already evident from directly examining the labels of the axes.

Before plotting the data, the scales of both axes must be determined. The overall graph should be large enough that trends are apparent, however not so large that the data points occupy only a small quadrant of the figure. To set each scale, inspect the range of data for that axis. The full range should be expanded over the entire axis. You do not need to include the origin (the point where the axes intersect at 0) unless that point is important for later analysis. Additionally, divisions along each axis should be regularly spaced, marked, and labeled. Data points should be added as "•", "+", or "×"; their marks should be dark, but not excessively large, which would obscure their precise positions.

After the data have been plotted, you must find a relationship between the points. Statistical methods to produce a "line of best fit" exist in which the distance from each data point to the line inscribed by a mathematical formula is minimized. Many types of lines exist; some are linear and some are non-linear. Non-linear lines include an inversely proportional line, an exponential line, a logarithmic line, a polynomial line such as a cubic or quadratic function, and many others. The mathematical equations of these lines are used to describe the relationship between the variables used in the experiment.

Data is often fit using linear regression, in which the resulting relationship is described by a straight line with the following formula:

$$y = mx + b,$$

where y is the value of the dependent variable, x is the value of the independent variable, m is the slope of the line

$$\frac{\text{rise}}{\text{run}} = \frac{\Delta y}{\Delta x} = \frac{\text{change in } y}{\text{change in } x} = \frac{y_2 - y_1}{x_2 - x_1},$$

and b is the y-intercept (where the line crosses the y-axis at x = 0). Linear relationships that have a y-intercept of zero (go through the origin) are called directly proportional relationships.

In this experiment, mass and volume data are collected and related to each other as they apply to density. Density is an intensive property of a substance, independent of the amount of material being studied. As an object's volume increases, so does its mass; mass and volume are directly proportional variables. By selectively choosing objects larger and larger in volume (thus making volume the independent variable, plotted on the x-axis), the mass of the object is forced to increase as well (thus making mass the dependent variable, plotted on the y-axis).

If mass (y) and volume (x) data for different-sized objects made of the same material are plotted, the resulting linear mathematical relationship produces a line ($y = mx + b$))with a slope (m) equivalent to the material's density. Because an object with a volume of zero has no mass, the y-intercept (b) is zero. The resulting equation of the line is: mass = density (volume) + 0. Rearranging the equation for density produces the accepted version of the formula to determine an object's density.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Pre-Lab Discussion and Activity

Graphing the Method of Water Displacement

Place 15 glass marbles, all the same size, into a beaker. Explain that the demonstration will relate the level of the water to the number of marbles in the cylinder. To a 100-mL graduated cylinder, add five marbles. Next, fill the graduated cylinder with enough water to bring the water level to the 70-mL mark. Display a data table with two columns: "Number of Marbles" and "Water Level". Record the initial data point of 5 marbles and a water level of 70 mL. Add an additional marble and record its new water level in the data table.

Repeat until all 15 marbles have been added to the cylinder. Use the data to demonstrate the graphing process as described in the Background section above. Prove the y-intercept value by repeating the first steps: add five marbles to a 100-mL graduated cylinder and then fill the graduated cylinder until the water level reaches 70 mL. Decant only the water into another empty 100-mL graduated cylinder. The volume should be similar to the value of the y-intercept (where x = 0 marbles in the cylinder). Prove the value of the slope by measuring the diameter of one of the marbles with a caliper. Calculate the volume of the marble using the geometric mathematical equation for the volume of a sphere. The values should be similar.

$$V_{\text{sphere}} = \frac{4}{3}\pi r^3$$

Table: Sample data of marble demonstration

Number of Marbles	Water Level (mL)
5	70.0
6	71.7
7	73.5
8	75.2
9	77.0
10	78.9
11	80.7
12	82.3
13	83.9
14	85.7
15	87.3

1. Identify the independent and dependent variables in this demonstration. Explain the reasons for each.

The independent variable is the number of marbles added to the cylinder. The marbles were added independently from any other event or situation.

The dependent variable is the water level in the cylinder. The level depended upon the number of marbles added to the cylinder.

2. Choose a label with units for both the x-axis and the y-axis. Explain your choices.

The x-axis (the horizontal axis) should be labeled "Number of Marbles". Marbles are the objects being counted, therefore the units of "marbles" could be used. The number of marbles in the cylinder should be on the x-axis because the number of marbles is the independent variable.

The y-axis (the vertical axis) should be labeled "Water Level". The markings on the graduated cylinder are labeled in milliliters. Therefore the units of "mL" could be used. The water level in the cylinder should be on the y-axis because it is the dependent variable.

3. Choose a descriptive title for the graph.

"Water Displacement by Spherical Objects" or "Determining Volume of a Spherical Object" are two of many possibilities.

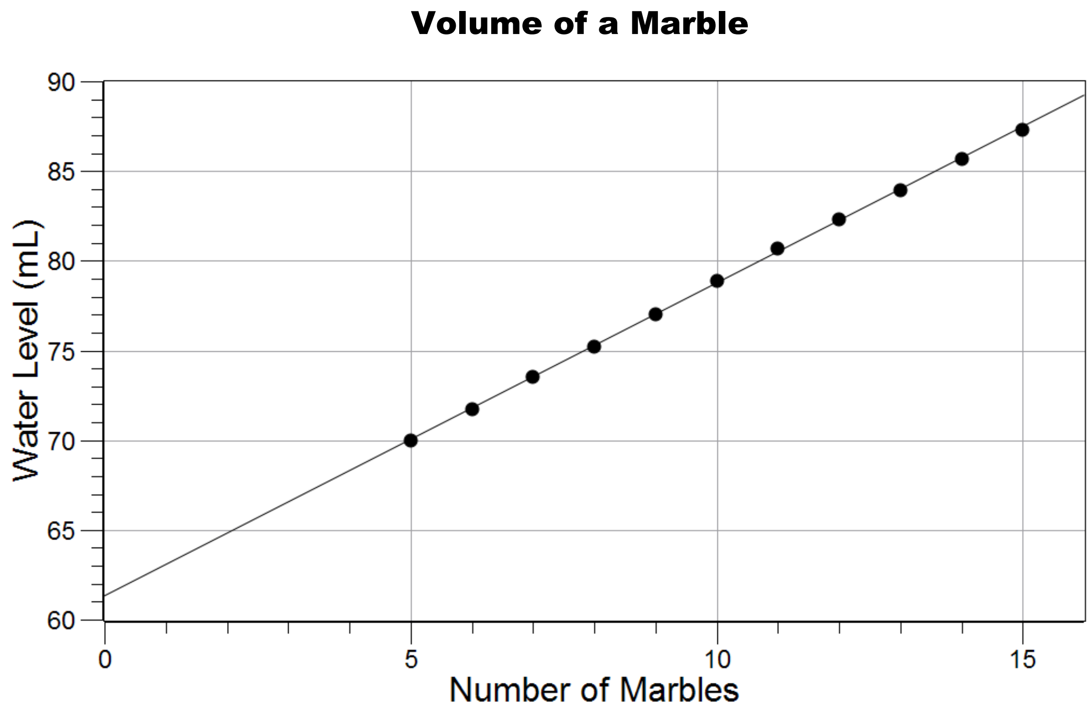
Teacher Tip: The title should convey the purpose of the graph. Avoid titles such as "Water Level versus Marbles" because it is not descriptive and could have been determined simply by looking at the labels of the axes.

4. Determine the range of data for each variable. Choose the starting and ending points for each axis such that information can be gathered from the final graph (y-intercept is important) as well as the major divisions for each axis scale.

The x-axis has a range of data spanning 5 to 15 marbles. The point at which the line crosses the y-axis (the y-intercept, where x = 0) represents the water level when no marbles are in the cylinder, therefore that point is an important location on the axis. You should show that point. The scale will, therefore, begin at 0 marbles and end at 15 marbles with equal divisions every 5 marbles.

The y-axis has a range of data spanning 70 to 90 mL. The y-intercept will be less than 70 mL, therefore the scale needs to begin at a point lower than 70 mL. You can estimate this by examining the data. The scale will, therefore, begin at 60 mL and end at 90 mL with equal divisions every 5 mL.

5. Create a graph of water level versus number of marbles. Plot the data using either "•", "+", or "×".



6. Is the relationship between the number of marbles and the water level in the graduated cylinder linear or non-linear? How do you know?

There is a linear relationship between the number of marbles and the water level in the graduated cylinder. The data fits on a straight line (constant slope), not a curved line (changing slope).

7. Use a ruler to draw the best fit line on the graph. Give the y-intercept and the slope of the line.

The y-intercept is 61.3.

The slope of the line is 1.74.

8. Give the mathematical equation for the equation of the line drawn on the graph.

General equation for a line: $y = mx + b$

Equation for the line on this graph: y = 1.74 (x) + 61.3

Water Level = 1.74 (marbles) + 61.3

9. What is the meaning of the y-intercept in this particular graph?

The y-intercept indicates the water level in the cylinder if there were no marbles at the beginning (where x = 0 marbles). In this case, where x = 0 marbles, the water level would be 61.3 mL.

10. On this graph, what does the slope mean? Hint: Use the units of the slope to determine its meaning.

The slope is calculated using rise over run. In this case, the slope, with proper units, is 1.74 mL/marble. The slope gives the average volume of the individual marbles (1.74 mL).

11. Is the number of marbles directly proportional to the water level in the graduated cylinder? How do you know?

There is a linear relationship between the number of marbles and the water level in the graduated cylinder, but the relationship is not directly proportional. It is not directly proportional because there is a y-intercept. For a linear relationship to be directly proportional it must go through the origin (0,0).

12. Calculate the volume of a marble using the geometric mathematical equation for the volume of a sphere.

The diameter of a marble: 1.49 cm

The radius of a marble: 1.49 cm / 2 = 0.745 cm

Volume: $V_{\text{sphere}} = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(0.745\text{ cm})^3 = 1.73\text{ cm}^3$

13. How does the volume calculated using water displacement and using the mathematical equation for the volume of a sphere compare?

The volume of a marble was almost exactly the same when calculated using the two different methods. Using water displacement the volume was found to be 1.74 mL per marble and using the mathematical formula the volume was determined to be 1.73 cm^3.

Density

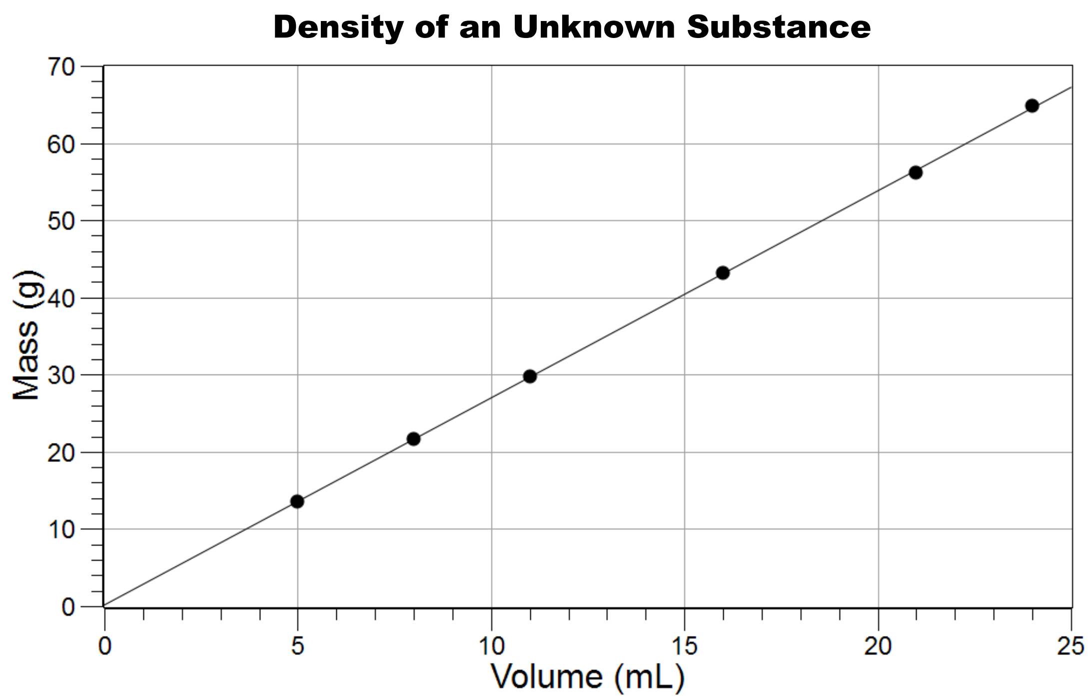
Review the concept of density. Give the formula for density.

$$\textbf{density} = \frac{\textbf{mass}}{\textbf{volume}}$$

Remind students that density is an intensive property specific for a particular material—all objects made from that same substance will always have the same density regardless of the individual shape and size of the object. This idea is the one being tested in this experiment by graphing mass and volume data. The relationship between the two variables produces a value equal to the slope of the line when the two variables are plotted. Display the following graph from the given data:

Table: Volume and mass of an unknown substance

Volume (mL)	Mass (g)
5	13.5
8	21.6
11	29.7
16	43.2
21	56.7
24	64.8



14. **What is the y-intercept for this line? Does this make sense? Explain.**

The y-intercept is 0. It makes sense because an object with no volume has zero mass.

15. **Is the mass directly proportional to the volume? How do you know?**

The mass is directly proportional to the volume. When the mass is zero the volume is also zero. Therefore, when the mass increases by a given factor the volume increases by the same factor.

16. How would the slope be calculated? What is the slope of this line?

The slope would be determined by finding $\frac{\text{rise}}{\text{run}}$.

In this case, mass is the dependent variable on the y-axis and volume is the independent variable on the x-axis. The equation is:

$$\frac{\text{rise}}{\text{run}} = \frac{\text{mass}_2 - \text{mass}_1}{\text{volume}_2 - \text{volume}_1} = \frac{56.7\text{ g} - 29.7\text{ g}}{21\text{ mL} - 11\text{ mL}} = \frac{27.0\text{ g}}{10\text{ mL}} = 2.7\text{ g/mL}$$

17. How does the equation to determine the slope compare to the formula for density? What does this mean? Do the units for the slope match those commonly used for density? Use the equation of the line to give the density of the material tested in the example graph above.

The values and the units are the same. This means that by plotting mass and volume data for different objects made of the same material, the resulting slope is equal to the density of the substance. The density of the substance tested in the example graph is 2.7 g/mL.

Lab Preparation

Although this activity requires no specific lab preparation, allow 10 minutes to gather the equipment needed to conduct the lab.

Safety

Follow all standard laboratory procedures.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

5	4	2	1	3
Use the slope of the best fit line to determine the density of the substance making up the set of objects.	Plot Mass versus Volume on a graph.	After measuring length, width, and height of each object, measure the mass of each object in the set.	Measure the length, width, and height of the objects in a set.	Once all the data has been collected, calculate the volume of the objects in the set.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Collect Data

Part 1 – Aluminum Objects

1. ☐ Look at the Table 1 below. Which values can you measure directly, and which values will you need to calculate?

Direct measurements: length, width, height, mass. Calculated values: volume

2. ☐ Predict the effect of volume on the mass of aluminum metal.

As the volume of the aluminum metal increases, its mass increases.

3. ☐ Measure the length, width, height, and mass of the four aluminum metal pieces. Be sure to record the values using the correct number of significant figures in Table 1 below.

Table 1: Aluminum objects

Aluminum Object	Length (cm)	Width (cm)	Height (cm)	Mass (g)	Volume (cm^3)
1	0.95	0.95	2.08	5.02	1.9
2	0.95	0.95	4.05	9.83	3.7
3	0.95	0.95	6.08	14.73	5.5
4	1.90	1.90	2.18	20.92	7.87

4. ☐ Calculate the volume for each of the four aluminum objects and enter them in Table 1 above.

Volume is calculated using the following equation: Volume = (length) × (width) × (height)

$$\text{Volume of object 1} = (0.95\text{ cm}) \times (0.95\text{ cm}) \times (2.08\text{ cm}) = 1.9\text{ cm}^3$$

Part 2 – Plastic Objects

5. ☐ Predict the effect of volume on the mass of the plastic objects.

As the volume of the plastic objects increases, the mass also increases. This is because the two variables are directly related.

6. ☐ Measure the length, width, height, and mass of the four plastic objects. Be sure to record the values using the correct number of significant figures in Table 2 below.

Table 2: Plastic objects

Plastic Object	Length (cm)	Width (cm)	Height (cm)	Mass (g)	Volume (cm^3)
1	1.30	1.30	1.94	4.44	3.28
2	1.30	1.30	4.44	10.37	7.50
3	1.30	1.30	6.40	14.64	10.8
4	1.61	1.61	5.59	19.52	14.5

7. ☐ Calculate the volumes for each of the four plastic objects and enter the values in the Table 2 above.

Volume is calculated using the following equation: Volume = (length) × (width) × (height)

$$\text{Volume of object 1} = (1.30\text{ cm}) \times (1.30\text{ cm}) \times (1.94\text{ cm}) = 3.28\text{ cm}^3$$

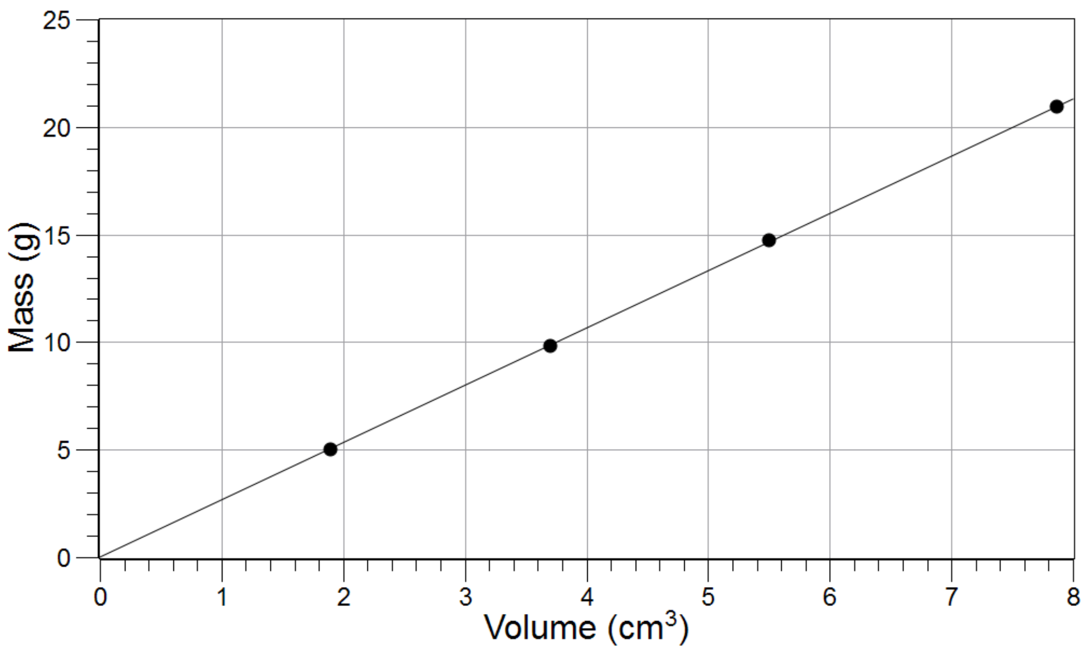
8. ☐ Clean up your lab station according to the teacher's instructions.

Data Analysis

Part 1 – Aluminum Objects

1. ☐ Plot Mass (g) versus Volume (cm^3). Consider the volume to be the independent variable (graphed on the x-axis) and the mass to be the dependent variable (graphed on the y-axis). Label the overall graph, the x-axis, the y-axis, and include units on the axes.

Density of Aluminum Objects



2. ☐ Is the relationship between mass and volume linear or non-linear? How do you know?

There is a linear relationship between mass and volume. The data can be fit using a straight line (there is a constant slope).

3. ☐ On the graph, add a line of best fit. In this case, the line of best fit should be a straight line. The data points might or might not fall exactly on the line, but there should be approximately the same number of data points above the line as below the line.

4. ☐ Find the equation of the line of best fit ($y = mx + b$) by determining the slope (m) of the line. Note that the y-intercept (b) should be zero because when the volume (x) equals zero, the mass (y) is also zero. The slope (m) of the line is found by marking two points on the line, then dividing the difference in y-coordinates (called the rise) by the difference in x-coordinates (the run).

slope of the aluminum objects = (20.92 g – 5.02 g) / (7.87 cm^3 – 1.9 cm^3) = 2.7 g/cm^3

equation: y = 2.7 (x)

mass = 2.7 (volume)

5. ☐ Calculate the individual densities of each aluminum object and the average value for the density of aluminum. Record the densities in Table 3 below.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Table 3: Density of aluminum objects

Aluminum Object	Mass (g)	Volume (cm^3)	Density (g/cm^3)
1	5.02	1.9	2.6
2	9.83	3.7	2.7
3	14.73	5.5	2.7
4	20.92	7.87	2.7
Average (g/cm^3)			2.7

6. ☐ How does the slope of the line compare to the average density calculated from the individual densities above?

The slope of the line is the same as the average density calculated from the individual densities.

7. ☐ The accepted value for the density of aluminum is 2.70 g/cm^3. How does this compare with your experimentally determined value?

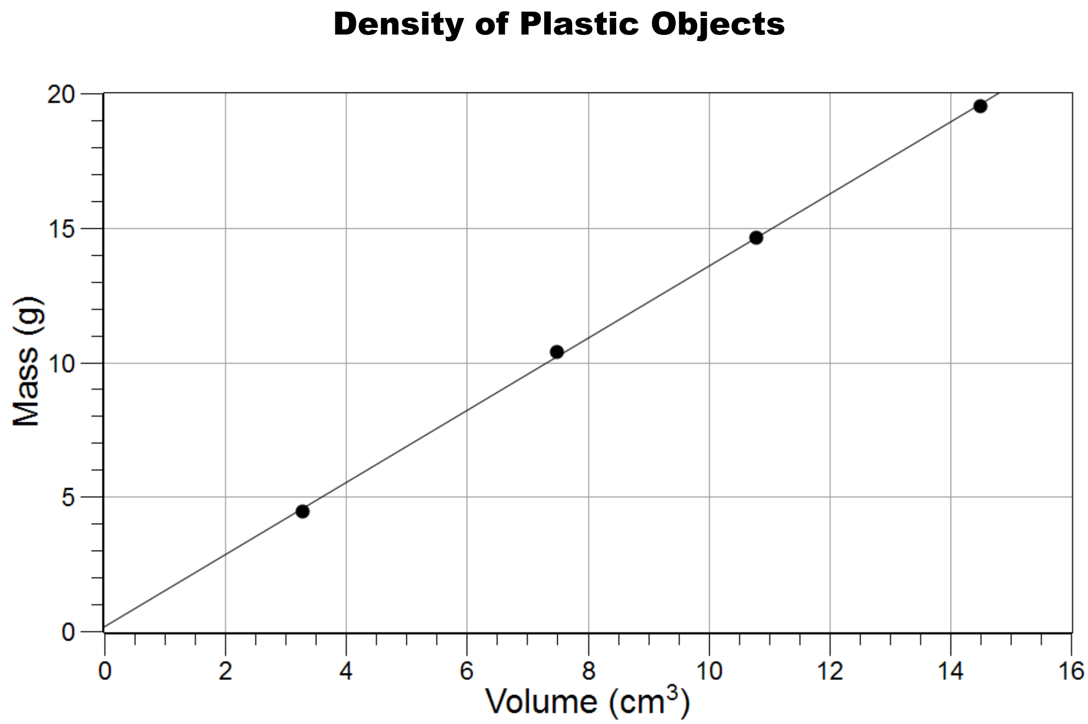
The accepted value of 2.70 g/cm^3 is the same as the values with two significant figures determined in the experiment. Some experimental error is expected, however calculated densities should be within 5% error from the accepted value.

Teacher Tip: Percent error may be calculated using the formula below.

$$\text{Percent error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100$$

Part 2 – Plastic Objects

8. ☐ Plot Mass (g) versus Volume (cm^3) for the plastic objects on the graph below. Label the overall graph, the x-axis, the y-axis, and include units on the axes.



9. ☐ Is the relationship between mass and volume for plastic objects linear or non-linear? How do you know?

There is a linear relationship between mass and volume of the plastic objects. The data can be fit using a straight line (there is a constant slope).

10. ☐ On the graph, add a line of best fit. In this case, the line of best fit should be a straight line. The data points might or might not fall exactly on the line, but there should be approximately the same number of data points above the line as below the line.

11. ☐ Find the equation of the line of best fit (y = mx + b) by determining the slope (*m*) of the line. Note that the y-intercept (*b*) should be zero because when the volume (*x*) equals zero, the mass (*y*) is also zero. The slope (*m*) of the line is found by marking two points on the line, then dividing the difference in y coordinates (called the rise) by the difference in x-coordinates (the run).

Slope of plastic objects = (19.52 g – 4.44 g) / (14.5 cm^3 – 3.28 cm^3) = 1.34 g/cm^3

Equation: y = 1.34 (x)

mass = 1.34 (volume)

12. ☐ Calculate the individual densities of each plastic object and the average value for the density of aluminum. Record the densities in Table 4 below.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Table 4: Density of plastic objects

Plastic Object	Mass (g)	Volume (cm^3)	Density (g/cm^3)
1	4.44	3.28	1.35
2	10.37	7.50	1.38
3	14.64	10.8	1.36
4	19.52	14.5	1.35
Average (g/cm^3)			1.36

13. ☐ How does the slope of the line compare to the average density calculated from the individual densities above?

The slope of the line is nearly the same as the average density calculated from the individual densities above.

14. ☐ The accepted value for the density of this particular type of plastic (polyvinyl chloride) is 1.39 g/cm^3. How does this compare with your experimentally determined value?

The accepted value of 1.39 g/cm^3 is very close to the values determined in the lab. These values ranged from 1.35 to 1.38 g/cm^3. Some experimental error is expected, however calculated densities should be within 5% error from the accepted value.

Teacher Tip: Percent error may be calculated using the formula

$$\text{Percent error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100$$

Analysis Questions

1. Explain the relationship between mass and volume (directly proportional, linear or non-linear). Are mass and volume always related this way?

The two variables are always directly proportional; as the volume of an object increases its mass also increases.

2. Identify the dependent and independent variables in this experiment.

The volume is the independent variable plotted on the x-axis; the mass is the dependent variable plotted on the y-axis.

3. If the axes upon which the variables are plotted were accidentally switched, would the slope of the line still equal the substance's density? How would you know?

If mass was plotted on the x-axis and volume on the y-axis, the resulting slope would not be equal to the density of the material. The slope is based on the rise over the run (the change in y over the change in x). If the axes were switched, the resulting slope would be based on volume/mass and have units of cm^3/g, which does not match the formula for density (density = mass/volume) with units of g/cm^3.

4. The y-intercept is not included in the formula to calculate density. Explain.

When mass and volume data are plotted, the resulting slope of the line represents the density. The y-intercept should be 0 because an object with no volume will not have any mass.

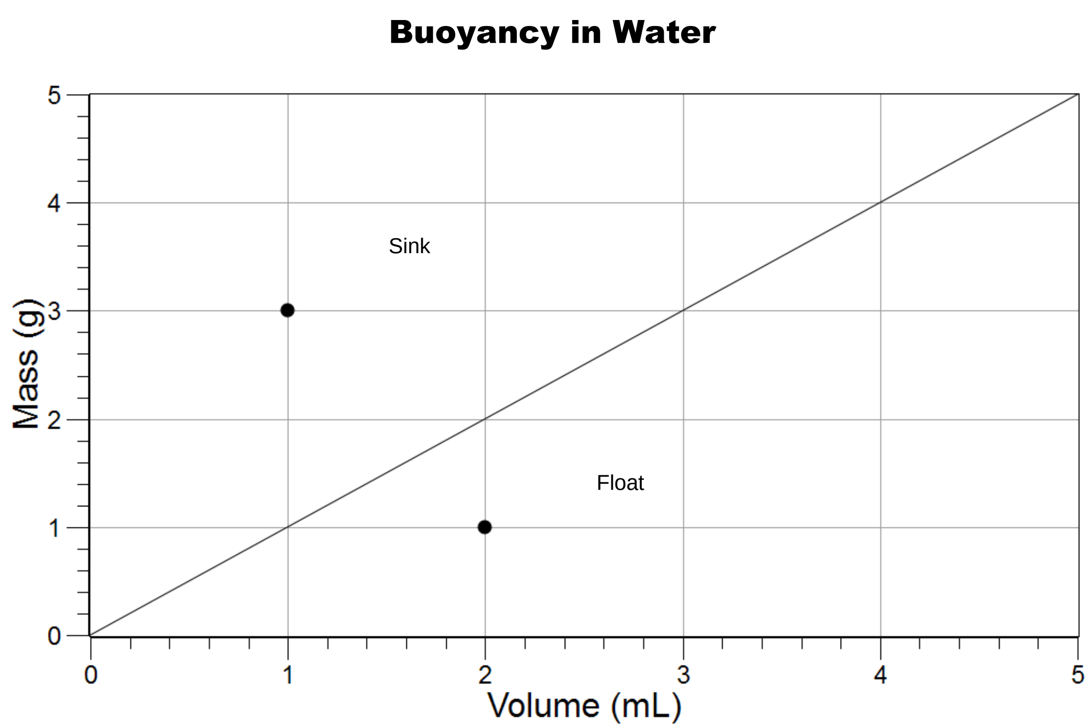
5. The values for the slopes of both the aluminum and the plastic were different. Explain how slope allows for the comparison of the densities of different materials.

Substances with greater (steeper) slopes are denser than those with smaller slope values.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Below is a graph with the density of water (ρ_{water} = 1.0 g/mL) plotted. Label the area where the data points representing objects that sink will be found. Also, label the area for objects that will float. Give an example for each.



At the point where the example has a mass of 3 g and a volume of 1 mL, the object sinks because its density is 3 g/mL, which is greater than the density of water (ρ_{water} = 1.0 g/mL). An object that will sink is a copper coin.

At the point where the example has a mass of 1 g and a volume of 2 mL, the object floats because its density is 0.5 g/mL, which is less than the density of water. An object that will float is a leaf.

2. A beaker of water is placed on a hot plate and allowed to boil. After 10 minutes of boiling, the volume in the beaker was measured every five minutes over the course of 30 minutes to produce the data shown in Table 5 below.

Table 5: Volume of water as water is allowed to boil

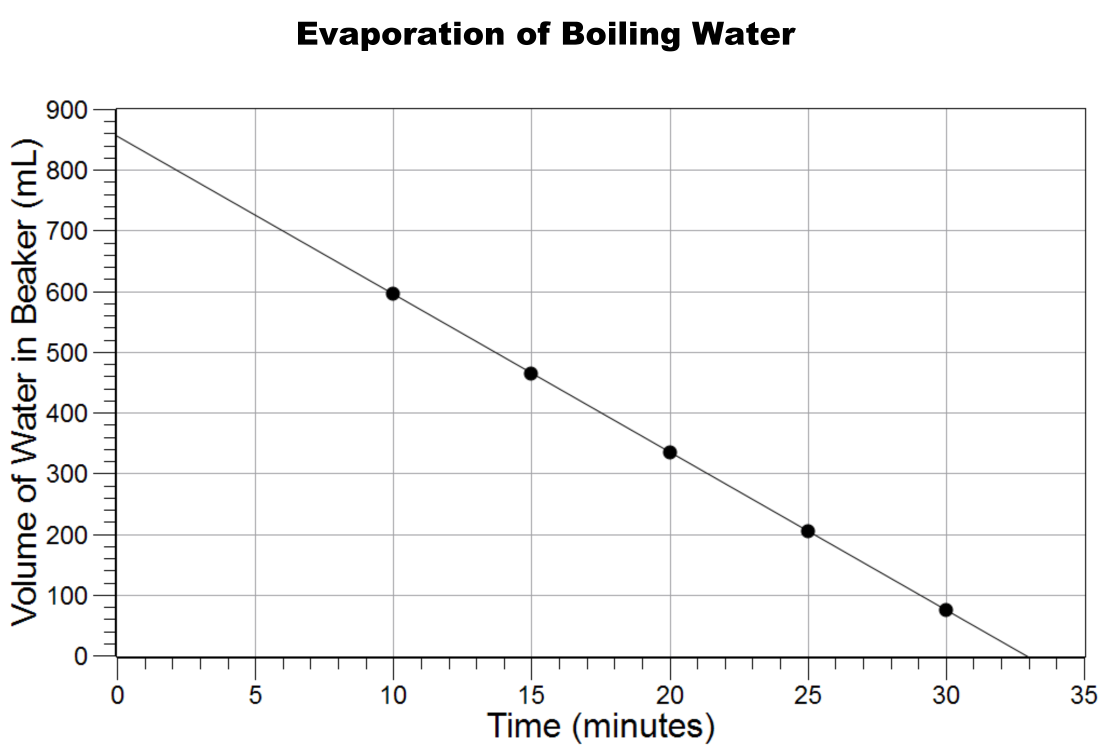
Time (minutes)	Volume (mL)
10	595
15	465
20	335
25	205
30	75

a. Identify the independent and dependent variables. Explain each choice.

The independent variable is the amount of time the solution was boiling. Time passes without influence from other factors.

The dependent variable is the amount of water in the beaker. The amount of time the water is allowed to boil determines how much water remains in the beaker.

b. Use the grid below to construct a graph of the data. Be sure to descriptively title the figure **and properly label each axis. Include a straight line to show the line of best fit.**



c. Give the equation for the line of best fit.

The equation for the best fit line is: $y = -26(x) + 855$.

d. Describe the relationship that exists between the volume of water in the beaker and the time the water is allowed to boil. Hint is there a directly proportional, linear, or non-linear relationship?

There is a linear relationship between the volume of water in the beaker and the time the water is allowed to boil. The data points fit a straight line so it is linear, but the graph does not go through the origin (0,0) so it is not directly proportional.

e. What does the value of the slope mean in this particular experiment? Use the units on the slope to help identify its meaning.

The slope has a value of –26. The slope is calculated from rise over run, or the change in volume over the change in time, the units are mL/min. The proper format for reporting the slope is: –26 mL/min. This indicates that 26 mL of water evaporates from the beaker each minute the water remains boiling.

f. What is the value for the y-intercept? What does the value of the y-intercept mean in this particular experiment?

The y-intercept is the point where the best-fit line crosses the y-axis when x = 0. If x = 0, then the value for the equation of the line is 855 mL. The point x = 0 represents the time at which the water began boiling. 855 mL is the amount of water that was in the beaker when it first started to boil.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Which statement about mass and volume data is always true?

- **A. They are directly proportional**
- **B.** They are inversely proportional (non-linear)
- **C.** They are both dependent variables
- **D.** They are both independent variables

2. A plot of mass and volume data that produces a steep slope is from a substance

- **A. With a density greater than one with a shallow slope**
- **B.** With a density less than one with a shallow slope
- **C.** That sinks in water
- **D.** That floats in water

3. A plot of mass and volume data that produces a slope less than 1

- **A.** Is very dense
- **B.** Is not dense
- **C. Floats in water**
- **D.** Sinks in water

4. The y-intercept is:

- **A.** The change in y-values over the change in x-values
- **B.** The change in x-values over the change in y-values
- **C. The y-value at the point where x = 0**
- **D.** The x-value at the point where y = 0

5. The slope is:

- **A. The change in y-values over the change in x-values**
- **B.** The change in x-values over the change in y-values
- **C.** The y-value at the point where x = 0
- **D.** The x-value at the point where y = 0

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The purpose of an **experiment** is to produce data that will prove or disprove a hypothesis. A good experiment is one that studies the most relevant variables. **Independent** variables are those whose conditions are set by the experimenter. **Dependent** variables are those whose values depend upon other variables. Variables that change in value at a constant rate and can be fit using a straight line are said to have a **linear** relationship. Linear relationships that do not have a y-intercept (and therefore pass through the origin) are called **directly proportional** relationships. Variables that change at varying rates and are fit using a curved line are **non-linear**. In **controlled** experiments, only one independent variable is changed at a time.

2. A good graph has a title that is a description of the **purpose** of the experiment. The **labels** of each axis contain the data's proper units of measure. The **independent** variable is plotted on the x-axis and the **dependent** variable is plotted on the y-axis. The **scale** of each axis depends on the range of data, and is set such that the data fills the entire graph. After the data is plotted, a line is added that mathematically relates the variables to each other. The equation of a straight line is in the form of **y = mx + b**. The **slope** of the line (*m*) is calculated by finding the change in the **y-values** over the change in the **x-values**.

Extended Inquiry Suggestions

Repeat the experiment using the water displacement method to determine the volume of each object and compare the results to results found using the geometrical mathematical formulas.

Determine if all plastics have the same density.

Determine if the shape of the material affects the density.

Repeat the experiment with different-sized blocks of wood. Compare the densities of various types of wood.

4. Percent Oxygen in Air

Objectives

Students learn about the components of air and how to determine the percent of oxygen in air. Through this investigation, students:

- Observe a chemical reaction involving different states of matter
- Describe pressure at the molecular level
- Explain how the variables temperature, volume, and concentration affect the pressure of gases

Procedural Overview

Students gain experience conducting the following procedures:

- Use an absolute pressure sensor to measure changes in pressure as atmospheric oxygen reacts with steel wool (iron)
- Determine the percent of oxygen in air from the measured pressure difference

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	45 minutes
♦ Lab activity	40 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Absolute pressure sensor
- Sensor extension cable
- Quick-release connector[1]
- Tubing connector[1]
- Tubing, 1- to 2-cm[1]
- Beaker, 150-mL
- Test tube, 25-mm x 150-mm
- One-hole stopper to fit the test tube
- Stir rod
- White vinegar (~5% acetic acid), 50 to 60 mL
- Steel wool, fine mesh (#000), 1.0 g
- Paper towels
- Glycerin, 2 drops

[1] Included with most PASCO absolute pressure sensors.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Particulate nature of matter
- ♦ States of matter

Related Labs in This Guide

Labs conceptually related to this one include:

- ♦ Gay-Lussac's Law and Absolute Zero
- ♦ Boyle's Law
- ♦ Ideal Gas Law

Using Your Data Collection System

- ♦ Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.
- ♦ Starting a new experiment on the data collection system �(1.2)
- ♦ Connecting a sensor to the data collection system �(2.1)
- ♦ Recording a run of data �(6.2)
- ♦ Displaying data in a graph �(7.1.1)
- ♦ Adjusting the scale of a graph �(7.1.2)
- ♦ Finding the coordinates of a point in a graph �(9.1)
- ♦ Saving your experiment �(11.1)
- ♦ Printing �(11.2)

Background

Air is a mixture made up of approximately 78% nitrogen, 21% oxygen, and 1% of several other gases including argon, carbon dioxide, and water vapor. In this lab, students calculate the percent of oxygen in the air by measuring initial air pressure and air pressure after oxygen has been removed from the air, and determining the difference.

Oxygen is removed from the air using a chemical reaction. The reaction between oxygen and iron was chosen because iron reacts with oxygen in the air, but not with nitrogen. Nitrogen molecules simply bounce off the iron, while the oxygen atoms collide and stick to the iron forming a new substance, iron(III) oxide (rust).

$$3O_2(g) + 4Fe(s) \rightarrow 2Fe_2O_3(s)$$

Steel wool is the source of iron. Steel is an alloy of iron with a very small amount of carbon. Most brands of steel wool are lightly coated with oil or some other rust inhibitor. The students remove the rust inhibitor by rinsing the steel wool in vinegar. The vinegar "washing" also creates a moist, slightly acidic environment that increases the reaction rate.

Solids, liquids, and gases all exert pressure on their surrounding surfaces. Since pressure is the dependent variable in this experiment, it is important that students understand pressure at the molecular level. Pressure is the force applied over an area. Pressure is measured in units of pascal (Pa) or as a newton (N) per square meter, where (1 Pa = 1 N/m^2). The newton is the standard unit for measuring force. At the molecular level, air pressure, like all pressure involving gases, results from molecules colliding with a surface. The greater the number of collisions there are per second, the greater the air pressure.

The temperature, volume, and amount of gas present all affect the frequency of gas particles colliding and thus the air pressure. When air is heated, the particles move faster (greater kinetic energy) causing more collisions per second and therefore increasing the pressure (Gay-Lussac's law). Cooling, on the other hand causes air molecules to move slower, resulting in fewer collisions and less pressure.

When the volume of a container holding a constant number of gas particles increases, the pressure decreases (Boyle's law). This is because the particles have more space in which to move and therefore collide with the walls of the container less often. The opposite is also true: less volume for the gas particles to move in results in a higher pressure because the molecules have less space in which to move and therefore hit the surfaces more frequently.

Air pressure is also dependent on the number of gas particles present (Avogadro's law). The more particles there are in a given container, the more frequently the walls of the container will be struck, causing greater pressure. The opposite holds true also: the fewer the particles present, the lower the resulting pressure.

The method used in this experiment to determine the percent of oxygen in the air works because of Dalton's law of partial pressures. Dalton's law of partial pressures states that the total pressure of a gas mixture is the sum of the partial pressures of each individual gas in the mixture. Therefore, air pressure is equal to the pressure of nitrogen plus the pressure of oxygen plus the pressure of all other gases present in the air. When oxygen is removed, the air pressure will be reduced by an amount directly proportional to the percent of oxygen in the air.

$$\%\text{oxygen} = \frac{\text{change in pressure}}{\text{initial pressure}} \times 100$$

Pre-Lab Discussion and Activity

The Concept of Pressure

Introduce the concept of pressure by comparing the pressure exerted in high heel shoes versus tennis shoes.

1. If a woman is wearing high heels on a soft dirt surface, she will sink into the ground. The same woman, however, wearing tennis shoes on the same surface will not sink at all. Why? What is the difference in the two scenarios and how does this difference explain whether or not the woman will sink?

The difference is the shoes that the woman is wearing. More specifically, the surface area touching the ground is significantly different. When the woman is wearing high heels, her body weight exerts force on the ground over a very small area. This creates a lot of pressure and she sinks. When the woman is wearing tennis shoes, she exerts the same force (her body weight) over a larger area and does not sink (she is exerting less pressure).

Calculating Pressure

Provide the students with a mathematical understanding of pressure by comparing the pressure exerted by a textbook in two different positions. Place a textbook on a table in front of the class with its largest side down. Stand a second textbook up on its edge. Write the dimensions and mass of the textbook on the board. (The following dimensions may be used: textbook dimensions, 10.0 in. x 10.0 in. x 2.0 in. textbook weight (force of gravity on the book), 5.0 lb) Explain that the idea of the "amount of force on each square" is called pressure. Write the definition of pressure on the board. Pressure is the force acting on a specific area.

2. Which book is exerting more pressure per unit area on the table? Explain.

The book standing on its edge is exerting more pressure per unit area. Both books have the same mass and are therefore exerting the same force, but the book standing on its edge is supporting the weight over a smaller area.

3. If the book is lying on its side how much of the force (weight) is felt over each square inch?

The weight of 5.0 lb is being shared over 100 square inches (the area touching the table). So each square inch of the book supports 0.05 lb.

4. How much force (weight) does each square inch support when the book is sitting on its edge?

There are fewer square inches (in.^{2}) in contact with the table so each square inch must support more weight. The book still weighs 5.0 lb, but there are only 20 in.^{2} sharing the weight. So each square inch supports 0.25 lb.

5. What is the equation for pressure? Calculate the pressure each book is exerting on the table?

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

The book lying on its side: $P = \dfrac{5.0\text{ lb}}{(10\text{ in.} \times 10\text{ in.})} = 0.05\text{ lb/in.}^2$

The book sitting on its edge: $P = \dfrac{5.0\text{ lb}}{(10\text{ in.} \times 2\text{ in.})} = 0.25\text{ lb/in.}^2$

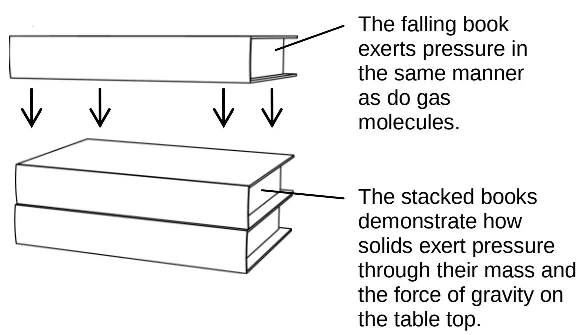
6. What are the units of pressure in the above example? What are the units in the SI system?

In the above example, pressure was measured in pounds per square inch. In the SI system, the unit of force is newton (N) and the unit for area is square meters (m^2). A newton ($(kg{\cdot}m)/s^2$) is the standard unit of force containing the units for both mass as well as the acceleration due to gravity; combined, these can be thought of as "weight". A N/m^2 is also called a pascal (Pa). One pascal of pressure is really small, so it is common to measure pressure in kilopascals (kPa).

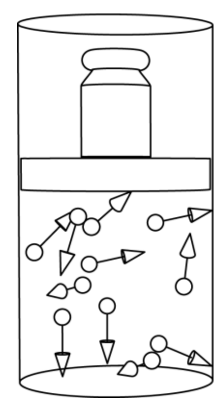
Pressure at the Molecular Level

Explain to the students that in solids and liquids, the atoms or molecules are all in contact with each other. Because of this, all the molecules contribute to the force on the table. Stack two textbooks on top of each other and hold a third text book in your hand. Explain that each book on the table contributes its weight to the total force and pressure being exerted just like each molecule in a solid or liquid contributes all of its weight to the pressure being exerted.

Molecules that make up a gas, however, are different because they are not in contact with each other. A gas particle could be thought of as a third book that is falling toward the stack of books. Drop the third book onto the stack of books. The falling book does not contribute any force to the surface of the table until it hits the stack. For gases, it is the number of collisions that create pressure. End the discussion by showing the students a picture or a video of gas molecules colliding with the surface of their container.



Demonstration of the difference between pressure exerted by solids versus pressure exerted by gases.



Gas molecules create pressure through collisions with the surfaces of their container.

7. Using the textbook example, explain pressure at the molecular level?

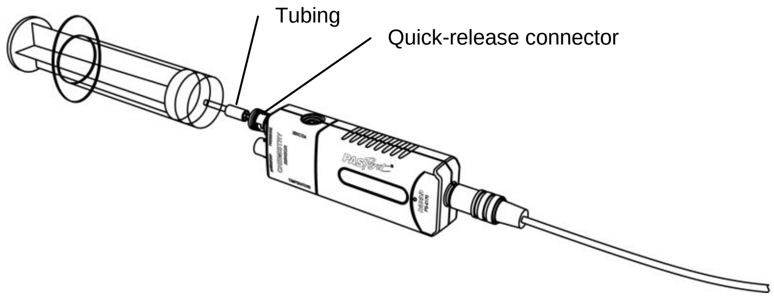
The molecules that make up the book are exerting a force (their weight) onto the table.

8. How is the pressure exerted by molecules different if the molecules are in a solid or liquid state versus a gaseous state?

In solids and liquids, all the molecules are touching each other all the time and thus always contribute to a constant force and pressure. Molecules in the gaseous state, on the other hand, are bouncing around. They contribute to the force when a collision occurs. Collisions are how gas molecules create pressure.

Temperature, Volume, and Air Pressure

Engage the students in a discussion about air pressure and the ways in which pressure can be altered by having students change the temperature and then the volume of air inside a syringe.



Have students set the volume of a syringe to about 20 mL and then connect it to a data collection system. While monitoring pressure in the digits display, have students observe what happens as they heat the air inside the syringe by holding the syringe in the palm of their hand. Next, have students change the volume of the air in the syringe by pushing the plunger in, compressing the gas. Help students analyze their results by explaining what is happening at the molecular level.

9. What are some possible ways to increase the pressure of the air inside the syringe?

Increase the number of collisions by either: 1) increasing the temperature, 2) increasing the number of molecules, or 3) decreasing the size of the container.

10. What do you notice about the pressure when you hold the syringe in the palm of your hand? Why?

The pressure goes up because the gas molecules get warmer. Warmer things move faster so the gas molecules hit the walls inside the syringe more often and with more force producing more pressure.

11. Why is it possible to compress the gas? How does it change the pressure? Why?

The gas can be compressed because there is space between the gas molecules. The molecules can move into this space and essentially squish closer together.

The pressure increased because the molecules hit the walls more often. The molecules hit the walls more often because they did not have to travel as far to hit the walls.

Types of Molecules in Air

Discuss the different types of molecules that make up air and explain that each type of molecule contributes to the pressure exerted by air. This observation is summarized in Dalton's law of partial pressures which states that the total pressure of a gas mixture is the sum of the individual pressures of each type of gas.

Introduce the idea that the number of molecules in a closed container can be increased or decreased through chemical reactions. Explain the chemical reaction that will be performed in this lab and end the discussion by having the students perform a sample calculation using provided pressure data.

Write the following equation on the board:

oxygen gas + iron (steel wool) → rust

$$3O_2(g) + 4Fe(s) \rightarrow 2Fe_2O_3(s)$$

12. What types of molecules are in the air in your syringe?

Air is a mixture made of nitrogen molecules, oxygen molecules, and a very small amount of other molecules such as argon, carbon dioxide, water vapor, and others.

13. All of the molecules in air are bouncing on the walls and contributing to the total gas pressure. Imagine there were 100 gas molecules. If 50 of them were nitrogen, then what percent of the observed pressure is due to nitrogen molecules?

Half (50%) of the collisions would be nitrogen molecules so half of the pressure would be from nitrogen.

14. We have seen that altering the temperature causes the pressure to either increase or decrease because the number of collisions and the force of the collisions both change. We have also seen that decreasing the volume of the container increases the pressure by increasing the number of collisions. The final variable that affects the pressure is the actual number of gas molecules. How can the number of gas molecules in a closed container change?

The number of gas molecules can either be increased or decreased through a chemical reaction.

15. In this lab, oxygen molecules in the air will bounce into the iron atoms in the steel wool. If the collision is just right, the oxygen reacts with the iron and creates a new molecule (rust). How will the gas pressure be affected as the reaction proceeds? Why?

The gas pressure should decrease because the number of gas molecules should decrease. Before the reaction, oxygen molecules moved freely as a gas. After the reaction, the oxygen will be bonded to the iron in a newly formed solid (rust).

16. How will measuring a change in pressure help you determine the percent of oxygen in the air?

The amount that the pressure decreases is proportional to the percent of oxygen in the air.

$$\%\text{ oxygen} = \frac{\text{change in pressure}}{\text{initial pressure}} \times 100$$

17. If the starting pressure of the gas in a test tube is 100 kPa and the final pressure is 75 kPa, what was the change in pressure? What percent of the gas molecules were removed?

The change in pressure is 25 kPa (100 kPa – 75 kPa).

$$\%\text{ gas lost} = \frac{25\text{ kPa}}{100\text{ kPa}} \times 100 = 25\%$$

Lab Preparation

Although this activity requires no specific lab preparation, allow 10 minutes to gather the equipment needed to conduct the lab.

Safety

Add this important safety precaution to your normal laboratory procedures:

♦ Vinegar is a weak acid. Avoid contact with the eyes and wash your hands after handling glassware, steel wool, and equipment.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

3	1	2	4	5
Measure and record the initial pressure.	Clean the steel wool with vinegar so that oxygen can react with the iron.	Create a closed container with air and steel wool.	Measure the final pressure value.	Calculate the percent oxygen in air.

Procedure with Inquiry

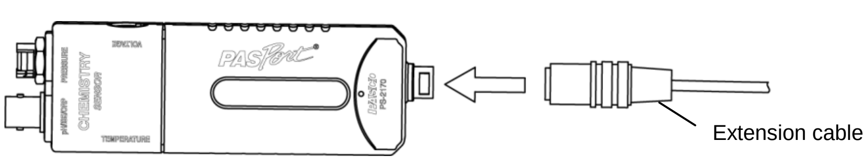
After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect the absolute pressure sensor to the data collection system using a sensor extension cable. �(2.1)

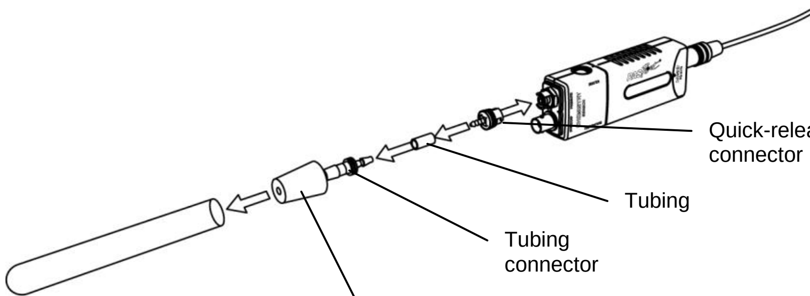


3. ☐ Connect the quick-release connector to the stopper using the tubing connector and the 1- to 2-cm piece of tubing by following the steps below. Use the picture as a guide.

a. Insert the thicker end of the tubing connector into the hole in the stopper. If this is difficult, add a drop of glycerin.

b. Connect the 1- to 2-cm piece of tubing to the other, thinner end of the tubing connector.

c. Insert the barbed end of the quick-release connector into the open end of the 1- to 2-cm piece of tubing. If this is difficult, add a drop of glycerin.



4. ☐ Insert the quick-release connector into the port of the absolute pressure sensor and then turn the connector clockwise until fitting "clicks" onto the sensor (about one-eighth turn).

5. ☐ Create a graph display of Pressure (kPa) versus Time. �(7.1.1)

6. ☐ What are the dependent and independent variables in this experiment? In what units are these variables measured?

The dependent variable is the pressure in units of kilopascals (kPa).

The independent variable is the time that the reaction is allowed to occur in units of seconds (s).

7. ☐ Predict what will happen to the pressure as the reaction occurs?

The pressure decreases as the oxygen gas is consumed by the reaction with iron.

8. ☐ Obtain enough fine mesh steel wool to fill a large test tube about ⅔ full (approximately 1.0 g).

9. ☐ Stretch the steel wool apart so that a large amount of surface area is exposed.

10. ☐ Clean the steel wool by soaking it in a 150-mL beaker containing approximately 60 mL of vinegar for about one minute. Use a stir rod to fully rinse the steel wool in the vinegar.

11. ☐ Why do we need to rinse the steel wool in vinegar?

Rinsing removes the protective coating from the iron so that the oxygen molecules can collide directly with the iron atoms. The vinegar also provides a moist, slightly acidic environment that will cause the reaction to occur faster (increase the rate of the reaction).

12. ☐ Remove the steel wool from the beaker of vinegar and wring it out, draining the vinegar into the beaker.

13. ☐ Stretch apart the steel wool and thoroughly dry it with paper towels.

14. ☐ Change the paper towels and dry it again.

15. ☐ Stretch the steel wool apart and shake it in the air to make sure it is dry.

16. ☐ Put the steel wool in a large test tube making sure that a large surface area is still exposed. Do not pack the steel wool into the bottom of the test tube.

Note: You may have to gently tap the test tube to get the steel wool to slide down into the test tube.

Collect Data

17. ☐ Place the stopper into the top of the test tube and immediately start collecting data. �[(6.2)]

Note: You may have to adjust the scale of the graph to observe any changes taking place. �[(7.1.2)].

18. ☐ What molecules are contributing to the pressure you are recording on your data collection system? Be specific.

The pressure is coming from the nitrogen, oxygen, and very small amount of other molecules that make up the air inside the test tube.

19. ☐ Write a sentence explaining the reaction occurring in the test tube. Explain where each substance comes from and its physical state (solid, liquid, or gas).

Oxygen gas from the air is reacting with solid iron in the steel wool to form rust, which is a new solid.

20. ☐ What is happening to the pressure as the reaction occurs? Why?

The pressure is decreasing because the oxygen gas is being incorporated into a solid and therefore is removed from the air.

21. ☐ Write down at least three changes you observe taking place in the test tube.

Water condenses on the side of the test tube, the test tube gets hot, and the steel wool turns brown/orange.

22. ☐ When the pressure has stabilized (after about 20 to 30 minutes), stop data collection. �[(6.2)]

23. ☐ Save the data file and clean up your lab station according to the teacher's instructions. �[(11.1)]

Data Analysis

1. ☐ Determine the initial and final pressures and write them in the Table 1 below. �(9.1)

Table 1: Initial and final pressure

Initial Pressure (kPa)	105.21
Final Pressure (kPa)	83.31

2. ☐ Calculate the change in pressure.

Initial pressure (kPa) – final pressure (kPa) = change in pressure (kPa)

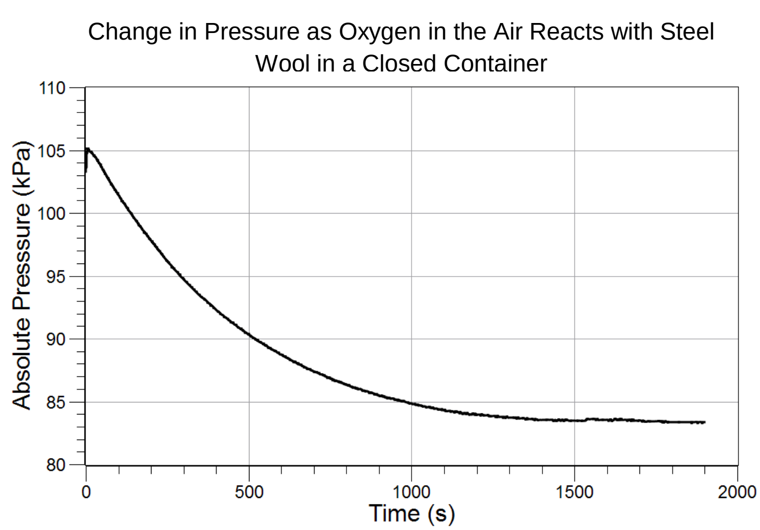
105.21 kPa – 83.31 kPa = 21.9 kPa

3. ☐ Calculate the percent oxygen in air.

$$\frac{\text{change in pressure (kPa)}}{\text{initial pressure (kPa)}} \times 100 = \%\ \text{oxygen}$$

$$\frac{21.9\ \text{kPa}}{105.21\ \text{kPa}} \times 100 = 20.8\%\ \text{oxygen}$$

4. ☐ Sketch or print a copy of the graph of Pressure (kPa) versus Time (s). Label the overall graph, the x-axis, the y-axis and include units on the axes. �(11.2)



Analysis Questions

1. Why did the pressure graph flatten out after a while? (Hint: think about what is happening to the amount of oxygen in the test tube.)

The oxygen in the test tube was being used up. When it was gone, the graph became flat.

2. Why was the pressure not reduced to zero?

The pressure did not make it to zero because there were still other molecules, including nitrogen, carbon dioxide, water vapor, and argon in the test tube. They continued bouncing into the walls, causing pressure.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Gases are often described as having no definite shape and filling the container they occupy. Explain what is happening at the molecular level to give gases these properties.

The gas molecules bounce around and do not bond together. They move through space in a straight line until they hit a wall, so they occupy the entire volume of any size container.

2. Explain why solids have a definite shape.

The molecules or atoms in a solid are bonded together in a fixed pattern. They cannot move around to fill the container.

3. Chemical reactions stop when one of the reactants is used up. This reactant is called the *limiting reactant* because it limits the amount of product that is made. In this lab, rust was the product. What was the limiting reactant?

The pressure stopped changing when the oxygen in the test tube was all used up. This makes oxygen the limiting reactant.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Which of the following variables affects the pressure of a gas?

- **A.** The number of gas molecules
- **B.** The temperature of the gas molecules
- **C.** The volume of the container the gas molecules are in
- **D.** **All of the above**

2. If you increase the temperature of a gas, what will happen to the pressure?

- **A.** It will stay the same
- **B. It will increase**
- **C.** It will decrease
- **D.** Not enough information

3. If you increase the number of gas molecules in a container, what will happen to the pressure?

- **A.** It will stay the same
- **B. It will increase**
- **C.** It will decrease
- **D.** Not enough information

4. Approximately what percentage of air is made up of oxygen gas?

- **A.** Less than 5%
- **B. 20%**
- **C.** 70%
- **D.** More than 80%

5. Pressure is best described as

- **A. A force spread out over an area**
- **B.** The motion of molecules
- **C.** The space between molecules in a gas
- **D.** A strong force

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Pressure is a force spread out over an **area**. Gas pressure is caused by gas molecules flying through space and **bouncing** off surfaces. If the collision rate increases, the **pressure** goes up. An increase in **temperature** causes greater pressure because the gas molecules are moving with more kinetic energy, and therefore are moving faster. A decrease in volume causes an **increase** in pressure because the gas molecules are closer together and have less distance to travel to hit the walls of the container, so collisions are more frequent. At a given temperature, all gas molecules contribute to the total pressure. If 70% of the gas molecules in a container are nitrogen, then **70%** of the pressure will be due to the nitrogen molecules.

Extended Inquiry Suggestions

Have students design an experiment to determine if other metals corrode in the presence of oxygen.

Have students determine the ideal amount of steel wool to use in this lab.

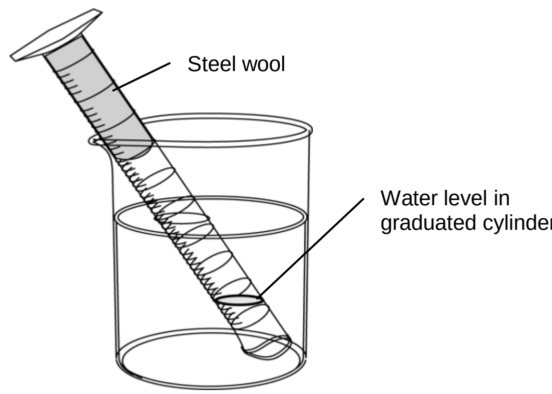
Have students determine the ideal size of a test tube to use in this experiment.

Have students determine the effects of altitude on the percent of oxygen in the atmosphere.

Discuss the following questions with students:

- Does changing the amount of steel wool change the calculated percentage of oxygen?
- Does changing the amount of air available (by using different sizes of test tubes) change the calculated percentage of oxygen?

Have students design an experiment to determine the amount of oxygen in the atmosphere using a water displacement method (see illustration) instead of the absolute pressure sensor. How do the results compare?



Setup for the water displacement method for determining the percent oxygen in air.

Structure and Properties of Matter

5. Conservation of Matter

Objectives

In this investigation, students test the law of conservation of matter for both physical and chemical changes.

Procedural Overview

Students conduct the following procedures:

- Find the mass of the reactants before the chemicals are reacted and the mass of the products after the reaction has occurred (in a chemical reaction where a precipitate forms)
- Measure the mass of a solute and a solvent separately and the mass of the solution after combining the two
- Determine the mass of gas produced during a chemical reaction by calculating the difference between the mass of the initial reactants and the mass of the final products (the final products do not include the escaped gas)

Time Requirement

♦ Preparation time	20 minutes
♦ Pre-lab discussion and activity	20 minutes
♦ Lab activity	50 minutes

Materials and Equipment

For each student or group:

- Balance
- Test tube (2), 15-mm x 100-mm
- Beaker, 250-mL
- Plastic soda bottle (with cap), 500-mL
- Sodium nitrate ($NaNO_3$), 5 g
- 0.1 M Sodium sulfate (Na_2SO_4), 5 mL[1]
- 0.1 M Strontium chloride ($SrCl_2$), 5 mL[2]
- Sodium bicarbonate ($NaHCO_3$), 8 g[3]
- 5% Acetic acid ($HC_2H_3O_2$), 30 mL[4]
- Distilled (deionized) water, 10 mL

[1]To formulate using solid sodium sulfate (Na_2SO_4), refer to the Lab Preparation section.

[2]To formulate using solid strontium chloride ($SrCl_2$), refer to the Lab Preparation section.

[3]Sodium bicarbonate ($NaHCO_3$) is household baking soda.

[4]Household vinegar can be used for the 5% acetic acid ($HC_2H_3O_2$).

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Mass
- Chemical and physical changes
- Evidence of a chemical reaction

Related Labs in This Guide

Labs conceptually related to this one include:

- Evidence of a Chemical Reaction
- Stoichiometry

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Note: There are no Tech Tips to list in this section as this activity does not use a data collection system.

Background

The law of conservation of matter states that matter cannot be created or destroyed by a physical or chemical change. In both cases, the number of atoms remains the same before and after the change. The law of conservation of matter does not apply to nuclear reactions, where matter may be changed to energy.

In a physical change, the substances before and after the change remain chemically the same, although the arrangement of the *molecules* and average motion of the particles may be different. During a chemical reaction, chemical changes occur and the *atoms* of one or more substances undergo rearrangements. The result of these rearrangements is the formation of new and different substances. The substances are made of the same atoms but are put together in a new way. All of the original atoms are still present.

Because of the law of conservation of matter we are able to write balanced chemical equations. Such equations make it possible to predict the masses of individual reactants and products that are involved in a chemical reaction.

In the first part of this experiment, sodium sulfate (Na_2SO_4) chemically reacts with strontium chloride ($SrCl_2$) to form dissolved sodium chloride (NaCl) and strontium sulfate ($SrSO_4$). Strontium sulfate is a white, solid precipitate. The balanced chemical equation for this reaction is:

$$Na_2SO_4(aq) + SrCl_2(aq) \rightarrow 2NaCl(aq) + SrSO_4(s)$$

In the second part of this experiment, a solution is made by physically dissolving sodium nitrate in water to form a solution. This becomes cold to the touch and is an example of an endothermic process.

$$NaNO_3(s) \rightarrow NaNO_3(aq)$$

In the third and fourth parts of this experiment, sodium bicarbonate ($NaHCO_3$) chemically reacts with vinegar (acetic acid, $HC_2H_3O_2$). This reaction produces sodium acetate ($NaC_2H_3O_2$), water (H_2O), and carbon dioxide gas (CO_2).

$$NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) + CO_2(g)$$

First, the reaction is performed in an open system. Then it is performed again in a closed system. In the open system, the carbon dioxide gas is allowed to escape to the atmosphere. A loss of mass between the reactants and products occurs, but is not a violation of the law of conservation of matter. This is proven when the reaction is repeated in a closed system, in which the carbon dioxide gas is retained.

Pre-Lab Discussion and Activity

Physical and Chemical Changes

Place a small beaker containing an ice cube on a balance. As the ice cube melts, record its mass periodically while engaging the students in a discussion about change. Review the differences between physical and chemical changes. Discuss whether particular properties of the substance change when the substance undergoes a physical or chemical change. Discuss the effects of chemical and physical changes on the total mass of the system.

1. What is a physical change? Give several examples.

A physical change is one in which the physical appearance of a substance changes, but the chemical structure of the substance remains the same. Water (H_2O) is an example. Melting ice, boiling water, and condensing steam are all examples of physical changes in which a specific chemical substance is in different physical forms.

2. When an ice cube melts, what remains the same? What changes?

Ice is made of H_2O molecules as is liquid water. Therefore, the chemical substance stays the same. The number of water molecules present also does not change, which means that the mass of the ice cube is the same as the mass of the liquid water.

The physical appearance of the water changes. This is because the speed of the water molecules and the distance between the water molecules changes.

3. Does the mass before and after a physical change always stay the same?

Yes, during the physical change, the number of molecules before and after the change remains constant. Therefore, the mass does not change. This aligns with the law of the conservation of matter. It states that matter can neither be created nor destroyed, only changed.

4. **What is a chemical change (or chemical reaction)? What evidence suggests a chemical change has occurred? Give several examples.**

A chemical change is one in which a completely new substance is formed. The formation of a solid product from dissolved reactants (precipitate), the evolution of gas, a significant color change, and a change in temperature all indicate that a chemical change has occurred. For example burning paper, rusting, and reacting an acid with a base are all examples of chemical changes.

5. **Does mass stay the same during a chemical change?**

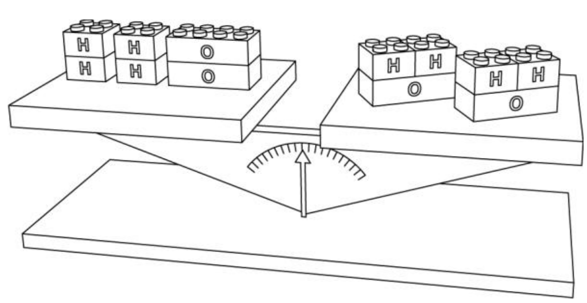
Yes, the atoms of reacting substances can rearrange and form new molecules. New substances form, even though the number of atoms of the total system remains unchanged. Often this is not obvious. For example, when burning paper, the mass of the paper seems to decrease, however the missing mass is simply changed into the gases that are released during burning. This aligns with the law of the conservation of matter which states that matter can neither be created nor destroyed.

Law of Conservation of Matter

Write the law of conservation of matter on the board: "Matter cannot be created or destroyed, only changed in form." Write the equation for the composition of water, $2H_2 + O_2 \rightarrow 2H_2O$, on the board. Ask the students what each of the numbers mean. As they are telling you, build each model using a model kit. Use an elementary school, double-pan balance and place the models on the pans. Place the reactants on one side and the products on the other. Most elementary school balances are not very precise and the students should see the reactants and products balance (even if there are minor variations in the mass of the individual components making up the models).

Teacher Tip: If you do not have a double-pan balance, you can simply measure the mass of the models on an electronic balance or single-pan balance and record the results. A less precise scale is preferred.

Teacher Tip: If you do not own a model kit, models of compounds can be built using large plastic preschool building blocks. Write the symbols for common elements on the different colored blocks.



$2H_2 \quad + \quad O_2 \quad \rightarrow \quad 2H_2O$

6. **The law of conservation of matter states that matter cannot be created or destroyed. What does this mean in your own words?**

Everything must be accounted for. Atoms cannot disappear or appear from nowhere. They can only be rearranged.

7. For the reaction of hydrogen with oxygen to produce water, what does each number in the equation mean? How can we build models of the molecules involved in this reaction?

The subscripts refer to the number of atoms of that type in each molecule. The subscript "2" in O_2 means there are two oxygen atoms in one molecule of oxygen. The coefficient (the number in front of the molecular formula) refers to the number of molecules. The "2" in front of the hydrogen molecule ($2H_2$) means there are two hydrogen molecules.

8. A different chemical reaction involves the decomposition of hydrogen peroxide to form water and oxygen gas. Beginning with two molecules of hydrogen peroxide (H_2O_2), how many water and oxygen molecules are produced to balance the pans?

The balanced chemical reaction is: $2H_2O_2 \rightarrow 2H_2O + O_2$. Two molecules of hydrogen peroxide decompose into two molecules of water and one molecule of oxygen. Notice on the reactant side, that there are a total of four hydrogen atoms and four oxygen atoms making up two hydrogen peroxide molecules. On the product side, the law of conservation of matter is obeyed. Although the four hydrogen atoms and four oxygen atoms are rearranged into different molecules, they exist on the product side as well.

Teacher Tip: Build two hydrogen peroxide molecules. Then allow the students to use the same number of each type of block to build the product molecules. Place their products on the balance and compare with the original reactant molecules.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. Prepare 100 mL of 0.1 M sodium sulfate (Na_2SO_4). This is enough for 20 lab groups.

   Dissolve 1.4 g of anhydrous Na_2SO_4 in 100 mL of distilled water.

   Alternatively, dissolve 3.2 g of sodium sulfate decahydrate ($Na_2SO_4 \cdot 10H_2O$) in 100 mL of distilled water.

2. Prepare 100 mL of 0.1 M strontium chloride ($SrCl_2$). This is enough for 20 lab groups.

   Dissolve 1.6 g of anhydrous $SrCl_2$ in 100 mL of distilled water.

   Alternatively, dissolve 2.7 g of strontium chloride hexahydrate ($SrCl_2 \cdot 6H_2O$) in 100 mL of distilled water.

3. Sodium bicarbonate ($NaHCO_3$) is household baking soda.

4. The 5% acetic acid ($HC_2H_3O_2$) may be replaced with household vinegar.

Safety

Follow all standard laboratory procedures.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	5	3	4
Determine the total mass of the starting substances and the glassware they are contained within.	Measure the starting substances and place in them into a 250-mL beaker.	Compare the mass and of the starting substances and the ending substances.	Mix the starting substances and record any observations.	Measure the mass all of the final substances and glassware.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Collect Data

Part 1 – Sodium Sulfate Solution and Strontium Chloride Solution

1. ☐ Place 5.0 mL of sodium sulfate solution (Na_2SO_4) into a test tube, and place the test tube in a 250-mL beaker.

2. ☐ Place 5.0 mL of strontium chloride solution ($SrCl_2$) into another test tube, and place it in the 250-mL beaker with the other test tube containing Na_2SO_4.

3. ☐ Determine the total mass of the solutions, test tubes, and beaker by placing them on a balance. Record this initial mass below.

Initial mass of Na_2SO_4, $SrCl_2$, and glassware (g): ____135.10 g______

4. ☐ Predict the amount of product that is produced from these reactants.

There should be the same mass of products as there are reactants.

5. ☐ Carefully pour the strontium chloride solution and the sodium sulfate solution into the beaker. Observe the chemical reaction and record your observations below.

A white solid forms (a precipitate) making the solution look cloudy.

6. ☐ How do you know a chemical reaction took place?

When mixed, the reactants formed a precipitate. The mixed solution appears milky. The bottom of the beaker may have become slightly cooler.

7. ☐ Place *both* test tubes back inside the beaker and measure the mass of the test tubes, beaker, and solution again. Record the final mass below.

Final mass of Na_2SO_4, $SrCl_2$, and glassware (g): ___________135.08 g ________

8. ☐ Why is it important to measure the mass of all the glassware together after the reaction occurs?

The initial mass also includes the mass of all the glassware. If the test tubes are removed and not measured after the reaction, the final mass is missing the mass of the original test tubes. This distorts the comparison.

9. ☐ Dispose of the solutions according to your teacher's instructions and then clean your glassware so that it may be used in the next part of this investigation.

Part 2 – Dissolving of Sodium Nitrate

10. ☐ Place approximately 5 g of solid sodium nitrate ($NaNO_3$) into a 250-mL beaker.

11. ☐ Place 10 mL of distilled water into a test tube, and place the test tube inside the 250-mL beaker containing the 5 g of $NaNO_3$(s).

12. ☐ Determine the total mass of the water, test tube, $NaNO_3$(s), and beaker by placing them on the balance. Record this initial mass below.

Initial mass of $NaNO_3$, H_2O, and glassware (g): ___________133.71 g ________

13. ☐ Predict the amount of product that is produced from these reactants.

Again, there should be the same mass of products as there are reactants.

14. ☐ Pour the water into the beaker with the solid and swirl the mixture until all of the solid dissolves. Record your observations below.

The solid dissolves and it feels cold.

15. ☐ Does a chemical reaction take place? Explain your reasoning.

No chemical reaction happens because no new substance is formed. The heat absorbed is used to break the lattice structure, but no new substance forms.

16. ☐ Place the test tube back inside the beaker and measure the mass of the test tube, beaker, and solution again. Record the final mass below.

Final mass of $NaNO_3$, H_2O, and glassware (g): ________133.70 g ________

17. ☐ Dispose of the solutions according to your teacher's instructions and then clean your glassware so that it may be used in the next part of this investigation.

Part 3 – Sodium Bicarbonate and Acetic Acid (Open System)

18. ☐ Place approximately 5 g of solid sodium bicarbonate ($NaHCO_3$) into a 250-mL beaker.

19. ☐ Pour 10 mL of 5% acetic acid ($HC_2H_3O_2$) into one test tube and an additional 10 mL into a second test tube. Place both test tubes inside the 250-mL beaker containing the $NaHCO_3$(s).

20. ☐ Determine the total mass of the $HC_2H_3O_2$, test tubes, $NaHCO_3$(s), and beaker by placing them on the balance. Record this initial mass below.

Initial mass of $NaHCO_3$, $HC_2H_3O_2$, and glassware (g): ______153.23 g_______

21. ☐ Predict the amount of product that is produced from these reactants.

The mass of the products will appear to be less than expected. This is because the reaction takes place in an open system. One of the products of the reaction is a gas. The gas escapes to the atmosphere and therefore is not measured. This is not a violation of the law of conservation of mass. Instead, it is an example of not measuring all of the components together. (This is similar to neglecting to add the test tubes when measuring the mass of the products.)

22. ☐ Pour the acetic acid from one of the test tubes into the beaker with the solid and swirl the mixture until the reaction subsides.

23. ☐ Pour the acetic acid from the second test tube and swirl the mixture until the reaction stops. Record your observations below.

Bubbles form.

24. ☐ Does a chemical reaction occur? Explain.

Yes, a chemical reaction occurs. The presence of bubbles indicates that a new substance (a gas) is produced.

25. ☐ Place the test tubes back inside the beaker and measure the mass of the test tubes, beaker, and solution again. Record the final mass below.

Final mass of $NaHCO_3$, $HC_2H_3O_2$, and glassware (g): _________152.50 g_______

26. ☐ Dispose of the solutions according to your teacher's instructions and then clean your glassware so that it may be used in the final part of the experiment.

Part 4 – Sodium Bicarbonate and Acetic Acid (Closed System)

27. ☐ Place approximately 3 g of solid sodium bicarbonate ($NaHCO_3$) in a clean 500-mL plastic soda bottle.

28. ☐ Pour 10 mL of 5% acetic acid ($HC_2H_3O_2$) into a test tube.

29. ☐ Carefully slide the test tube inside the plastic bottle containing the $NaHCO_3$(s), making sure not to spill any of the acetic acid.

30. ☐ Screw on the cap of the plastic bottle tightly.

31. ☐ While still being careful not to spill the acetic acid, place the bottle and its contents on a balance. Record this initial mass below.

Initial mass of the soda bottle and its contents (g): ___________46.48 g________

32. ☐ Why must you be careful not to spill the acetic acid at this point in the experiment?

The point of this experiment is to compare the mass of the system before and after a chemical change occurs. If even a small amount of acetic acid and sodium bicarbonate mix, they partially react.

33. ☐ Gently tip the bottle until the test tube inside spills the acetic acid into the sodium bicarbonate. Record your observations below.

Bubbles form and the pressure within the bottle increases.

34. ☐ Once the reaction is complete, and without unscrewing the cap, measure the mass of the bottle and its contents. Record the mass below.

Final mass of the closed soda bottle and its contents (g): _______46.45 g________

35. ☐ Remove the cap from the bottle and allow the gas to escape.

36. ☐ Why did the pressure inside the bottle increase?

One of the products of the reaction between sodium bicarbonate and acetic acid is a gas (carbon dioxide). Because the gas is unable to escape, it remains trapped inside the bottle, causing the pressure to increase.

37. ☐ Compared with the mass of the bottle and its contents before unscrewing the cap, do you expect the mass to be greater, less, or the same after the gas escapes?

Carbon dioxide gas is matter. It has mass and occupies space. Opening the bottle allows the carbon dioxide to escape. This causes the mass of the system to be less than when the carbon dioxide is trapped inside the bottle.

38. ☐ Screw on the cap and measure the mass of the bottle and its contents. Record the final mass below.

Final mass of the open soda bottle and its contents (g): _______46.17 g________

39. ☐ Dispose of the solutions and clean up according to your teacher's instructions.

Data Analysis

1. ☐ Determine the change in mass for each process. Record the results in Table 1 below.

Table 1: Initial mass, final mass, and change in mass

Experiment	Initial Mass (g)	Final Mass (g)	Change in Mass (g)
Part 1: $Na_2SO_4 + SrCl_2$	135.10	135.08	0.02
Part 2: Dissolving $NaNO_3$	133.71	133.70	0.01
Part 3: $NaHCO_3 + HC_2H_3O_2$ (open system)	153.23	152.50	0.73
Part 4: $NaHCO_3 + HC_2H_3O_2$ (closed system before opening the bottle)	46.48	46.45	0.03
Part 4: $NaHCO_3 + HC_2H_3O_2$ (closed system after opening the bottle)	46.48	46.17	0.31

2. ☐ How many grams of gas (CO_2) were formed in part 3 and part 4 of this investigation? How do you know?

In part 3 of this experiment, 0.73 g of CO_2 gas were formed. In part 4 of this experiment, 0.31 g of gas were formed. The amount of gas released is equal to the difference between the mass of the reactants and the products.

3. ☐ Calculate the percent change in mass for each part of the experiment and record them in Table 2.

$$\text{percent change} = \left|\frac{\text{change in mass}}{\text{initial mass}}\right| \times 100$$

Table 2: Percent change in mass

Experiment	Show Your Work Here	Percent Change in Mass
Part 1: $Na_2SO_4 + SrCl_2$	$\left\|\frac{0.02\text{ g}}{135.10\text{ g}}\right\| \times 100 = 0.007\%$	0.01%
Part 2: Dissolving $NaNO_3$	$\left\|\frac{0.01\text{ g}}{133.71\text{ g}}\right\| \times 100 = 0.007\%$	0.007%
Part 3: $NaHCO_3 + HC_2H_3O_2$ (open system)	$\left\|\frac{0.73\text{ g}}{153.23\text{ g}}\right\| \times 100 = 0.48\%$	0.48%
Part 4: $NaHCO_3 + HC_2H_3O_2$ (closed system before opening the bottle)	$\left\|\frac{0.03\text{ g}}{46.48\text{ g}}\right\| \times 100 = 0.06\%$	0.06%
Part 4: $NaHCO_3 + HC_2H_3O_2$ (closed system after opening the bottle)	$\left\|\frac{0.31\text{ g}}{46.48\text{ g}}\right\| \times 100 = 0.67\%$	0.67%

Analysis Questions

1. Why is the percent change in mass not always exactly 0%?

Balances are very sensitive pieces of equipment. Slight variations in the environment of the balance as well as unavoidable losses or gains in mass during the experimental procedure can change the final, displayed results.

2. What happens to the mass in part 3 and the second part of 4? Is this a case where the law of conservation of matter is untrue? Explain.

In both part 3 and the second part of 4, the reaction is open to the atmosphere. The gaseous products are allowed to escape from the container. The mass of the escaped gas could not be measured. Accounting for the mass of any gas in the reaction helps support the law, not refute it.

3. Do your results confirm the law of conservation of matter? Why or why not?

Yes, the results confirm the law of conservation of matter. In parts 1, 2, and the first part of 4, the masses before and after the change are essentially the same (different by less than 0.1%). In part 3 and the second part of 4, the final mass is less than the beginning mass, but these differences are due to the gas that escaped.

4. Does the law of conservation of matter apply to both physical and chemical changes?

Yes. In this experiment, part 1 is a chemical change and part 2 is a physical change. In both cases matter is conserved.

Synthesis Questions

Use available resources to help you answer the following questions.

1. In the process of electrolysis, electricity is used to convert water into its gaseous elements, hydrogen (H_2) and oxygen (O_2): $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$. If electrolysis is performed using 30 grams of water, how many grams of gas are produced?

After all of the water reacts, 30 grams of gas are produced. The law of conservation of matter states that you cannot gain or lose any mass during the reaction.

2. Pyrite is a shiny yellow mineral also known as "fool's gold". It is composed of iron and sulfur. If a 36.4 gram sample of pyrite is broken down into its elemental components and 17.3 grams of iron are produced, how many grams of sulfur are produced?

There are 19.1 grams of sulfur produced. That is the difference between the pyrite (36.4 g of combined iron + sulfur) and the iron (17.3 g).The law of conservation of matter states that you cannot gain or lose any mass during the reaction.

3. When a log burns, the resulting ash has less mass than the log. Why does this loss of mass not violate the law of conservation of matter?

When a log burns, gases (mainly carbon dioxide and water vapor) are formed and escape into the atmosphere. The ash (carbon) that remains after burning has less mass than the original log.

$$\text{Log}(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g) + C(s)$$

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. In a chemical reaction how does the mass of the products compare with the mass of the reactants?

- **A.** Greater than
- **B.** Less than
- **C.** **Equal to**
- **D.** Depends on if the reaction is endothermic or exothermic

2. If 7 grams of sodium (Na) reacts with 12 g of chlorine (Cl_2), how much table salt (sodium chloride, NaCl) is produced?

- **A.** 5 grams
- **B.** 13 grams
- **C.** **19 grams**
- **D.** 26 grams

3. What is the mass of the resulting gas when 3 grams of dry ice (solid carbon dioxide, CO_2) sublimes to gaseous CO_2?

- **A.** 0 grams
- **B.** 2 grams
- **C.** **3 grams**
- **D.** 5 grams

4. During a chemical reaction how does the total number of atoms of the reactants compare with the total number of atoms of the products?

- **A.** **Equal to**
- **B.** Greater than
- **C.** Less than
- **D.** Depends on the type of reaction

5. Which of the following states that matter cannot be created or destroyed?

- **A.** Kinetic molecular theory
- **B.** Collision theory
- **C.** Atomic theory
- **D.** **Law of conservation of matter**

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The l**aw of conservation of matter** states that **matter** cannot be created or destroyed, only changed in form. This means that any atoms present at the beginning of a reaction must also be present at the end of the reaction. The **number** of atoms can be counted in the laboratory by using a **balance** and measuring the **mass** of the reactants and the products. The mass before a change and after a change is **the same**.

2. During **physical** changes atoms do not rearrange to form new substances even though the **appearance** of the substance changes. An example is ice melting into liquid water. During **chemical** changes the atoms do rearrange to form new substances. If any of these new substances are gaseous, they may escape from a reaction. This occurs in **open** systems. C**losed** systems seal reactions from their surroundings so that gaseous products can be trapped and measured.

Extended Inquiry Suggestions

Perform an experiment in which either oxygen or carbon dioxide gas is one of the products. Use either a carbon dioxide gas sensor or an oxygen gas sensor to measure the ppm (parts per million) being produced. Calculate the mass of the gas along with the other products to determine if the mass of the reactants is the same as the mass of the products.

Determine the effect of melting or freezing on the mass of a substance.

Determine the effect of melting and freezing multiple times on the mass of a substance.

6. Properties of Ionic and Covalent Compounds

Objectives

Determine if an unknown substance is an ionic, polar covalent, or non-polar covalent compound based on its physical properties. Through this investigation, students:

- Review physical properties, including conductivity, solubility, hardness, and melting point
- Determine differences in physical properties for ionic and molecular covalent compounds
- Explain the differences between intramolecular and intermolecular forces

Procedural Overview

Students conduct the following procedures:

- Test the conductivity, solubility, hardness, and melting point of ionic, polar covalent, and non-polar covalent compounds
- Identify an unknown substance as an ionic, polar covalent, or non-polar covalent compound

Time Requirement

Preparation time	10 minutes
Pre-lab discussion and activity	30 minutes
Lab activity	45 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Conductivity sensor
- Hot plate
- Graduated cylinder, 10-mL
- Test tube (5), 15-mm x 100-mm
- Test tube rack
- Stopper, (3) to fit the test tubes
- Spatula
- Tongs
- Aluminum foil square (6), 5-cm x 5-cm
- Masking tape
- Wash bottle and waste container
- Distilled (deionized) water, 30 mL
- Table salt ($NaCl$), 1 g
- Table sugar ($C_{12}H_{22}O_{11}$), 1 g
- Paraffin wax, 1 g
- Unknown A, 1 g[1]
- Unknown B, 1 g[1]
- Unknown C, 1 g[1]

[1] Use glucose (also called dextrose, $C_6H_{12}O_6$) for unknown A; use crayon pieces for unknown B; and use potassium chloride (KCl) for unknown C.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Physical properties
- Types of bonds
- Polar versus non-polar molecules

Related Labs in This Guide

Labs conceptually related to this one include:

- Density
- Electrolyte versus Non-Electrolyte Solutions
- Intermolecular Forces

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor �(2.1)
- Monitor live data �(6.1)

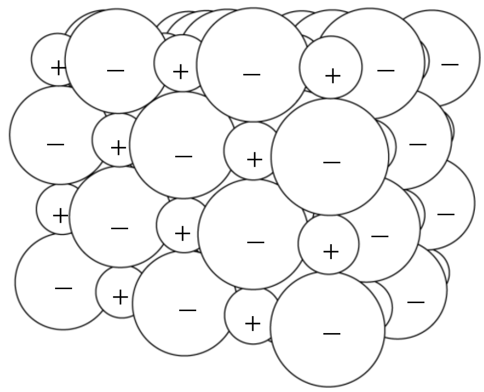
Background

In nature, a wide variety of forces exist. In chemistry, different electrical forces of attraction hold substances together. These electrostatic forces are the result of atoms interacting with electrons (subatomic particles with negative charge). Within the nucleus exists protons (subatomic particles with positive charge); it is the attraction between the positively charged nucleus and the outer negatively charged electrons that result in electrostatic forces of attraction.

Atoms are chemically bonded to each other as a result of intramolecular forces. These bonds exist within molecules because of a sharing of electrons between two bonding atoms. When each atom contributes an electron to the pair being shared between two atoms, it is called a covalent bond. If one atom shares a pair of electrons originally belonging to only itself with another atom, the result is a coordinate covalent bond. Metallic bonding occurs when electrons are equally distributed between a number of metal atoms, such that they are surrounded by a “sea of electrons.”

In some pairings of atoms (those between a metal and a non-metal), the electron is not shared. Instead, it is completely transferred from the metal to the non-metal. In such an instance, the metal completely loses one or more electrons and becomes positively charged, forming an ion (cation), while the non-metal completely gains the electrons and becomes negatively charged forming an ion (anion). Here, opposite charges attract the ions together much like a collection of magnets. The resulting electrostatic interaction forms an ionic bond.

The mass of alternating positive and negatively charged ions forms a crystal lattice. Because the units are not discrete, compounds formed between ions are not able to be called molecules. Instead, they are referred to as ionic compounds.

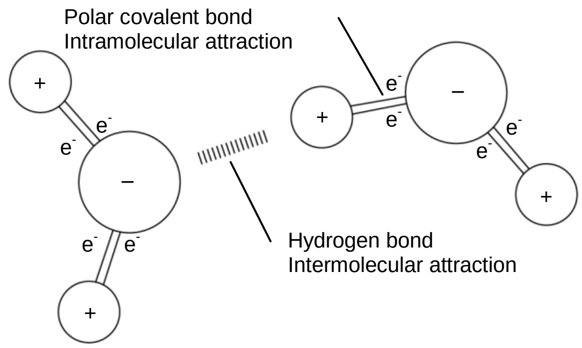


Sodium chloride in a solid crystal lattice arrangement.

Differentiating between covalent and ionic bonds depends upon the relative degree that the electrons are shared between atoms. This differentiation is based on the individual atoms' tendencies to gain or lose electrons, measured as electronegativity. Electronegativity is ordered on the Pauling scale (named after the American chemist, Linus Pauling) and has a range of 0 to 4. The larger the electronegativity value, the greater the atom's tendency to gain an electron.

Electronegativity differences between two atoms determine the bond classification: those with differences greater than 1.7 are considered ionic and those with differences less than 1.7 are considered covalent. With covalent bonding, even though electrons are shared an atom in the pair might have a greater affinity for the electrons. In these instances (electronegativity differences between 0.5 and 1.7), the electron is not equally shared and results in a polar bond, where the negative charge of the electron resides heavily with the atom with the greater electronegativity. For those pairings of atoms where there is not a great difference in tendency to gain or lose electrons (electronegativity values less than 0.5), the electron is equally distributed between the two atoms and results in a non-polar bond.

Neighboring molecules may also feel the attractions between the nuclei within its own molecule and the electrons of the molecule next to it. Intermolecular forces hold molecules together to form liquids and solids and are much weaker than the forces holding atoms together to form individual molecules. These intermolecular forces can be overcome with the input of energy. This causes solids to melt and liquids to boil, simply by overcoming the intermolecular forces holding the molecules to one another.



Water molecules

The properties of a substance are a direct result of the different intramolecular and intermolecular forces that exist because of the substance's composition and structure. Because atoms and molecules cannot be seen directly (even with the most powerful microscopes), chemists must rely upon macroscopic properties, such as conductivity, hardness, melting point, and solubility to determine the types of attractions present for a particular substance.

Pre-Lab Discussion and Activity

Demonstrations

Engage the students in a discussion about why different substances behave differently by demonstrating the following situations. Hold up a piece of chalk and a piece of plastic in front of the class and drop both of them. Next, place a packaging peanut in a beaker of water and another packaging peanut in a beaker of acetone (or nail polish remover). Discuss the results, emphasizing that different materials behave in different ways.

1. What happened to the chalk and the plastic when dropped? Why?

The chalk broke into pieces but the plastic did not change. Chalk and plastic consist of different substances that are held together with different types of chemical bonds.

2. What are packaging peanuts made up of?

Packaging peanuts are made up of plastic (polystyrene) and air.

3. Do packaging peanuts behave the same way in water and acetone?

Packaging peanuts remain unchanged when placed in water. When placed in acetone, however, the packaging peanuts dissolve and the air inside of them is released.

4. Why do packaging peanuts behave differently in water and in acetone?

Water molecules and acetone molecules are different and contain different types of bonds. Water is a polar substance and polystyrene is a non-polar substance so polystyrene does not dissolve. Acetone, on the other hand, has non-polar bonding and, therefore, dissolves the polystyrene.

Teacher's Tip: The rule "like dissolves like" can be included as a discussion topic here.

Ways That Atoms Combine to Form Macroscopic Matter

Review that the properties of matter are explained by the atoms that make them up and the way in which these atoms bond. Show a model of an ionic lattice structure, a covalent lattice structure, and a model (or drawing) of how molecular substances establish intermolecular attractions to form macroscopic matter.

5. What makes up all the different types of matter?

There are 117 known atomic elements which all have different properties and behaviors. In addition to their individual atomic properties, these atoms can also bond with other atoms to form the vast array of different types of compounds making up matter.

6. If atoms are so small they cannot be seen, even with a microscope, how is it possible that we can see matter?

Matter is made up of lots of atoms. We may not be able to see one atom, but when we have enough of them we can see them.

Teacher's Tip: Demonstrate this by putting one small grain of sand into a large jar and holding it at a distance. Compare this to when the jar is filled with sand.

7. How do atoms combine? What are these structures called?

Atoms combine in different ways. Some atoms bond to other atoms in lattice structures to form ionic compounds, while other atoms combine together into molecules and the molecules are held to each other by intermolecular forces.

8. Based on your understanding of how atoms combine, explain why the chalk broke when it was dropped and the plastic did not?

Chalk breaks when it falls because the ions in the lattice structure are shifted so that like charges are next to each other and they repel, causing the chalk to break. The plastic on the other hand did not break because it does not contain charged particles. Plastic is made up of molecules (macromolecules) that are intertwined with each other. When plastic hits the ground, the group of molecules shifts, but there is no sudden repulsion and, therefore, no breakage.

9. Why is water a liquid at room temperature, but oxygen is a gas?

Both water and oxygen are covalent molecular substances. The difference in their state of matter at room temperature is due to the strength of the intermolecular forces holding the molecules together. Water molecules are polar and oxygen molecules are non-polar. The polar water molecules are held together with stronger intermolecular attractions than the non-polar oxygen molecules.

Properties of Matter Explained

Substances that contain similar bonds have similar properties. So if you know the type of bonding in a substance, the material properties can be predicted. The reverse is also true. If you know the properties of a certain substance, then the bonding involved can be inferred.

10. How can you explain why certain substances have characteristic properties?

The properties of the substance are based on the type of atoms in the substance and the way they are bonded.

11. How can you predict the types of bonding in a given substance?

If you know the general properties of certain types of bonding, you can use the properties to predict the type of bonds in the substance.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Place the unknown substances in containers labeled "unknown A", "unknown B", and "unknown C".

Recommended unknowns are:

Unknown A: a polar covalent compound such as glucose (commonly called dextrose, $C_6H_{12}O_6$) or aspartame ($C_{14}H_{18}N_2O_5$), an artificial sweetener

Unknown B: a non-polar covalent compound such as crayons

Unknown C: an ionic compound such as potassium chloride (KCl)

Safety

Add these important safety precautions to your normal laboratory procedures:

- The hot plate gets extremely hot. Avoid contact with the hot plate until it has completely cooled.
- Keep all materials, especially electrical cords and paper, away from the hot plate while it is hot.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

4	3	1	2
Identify the bonding characteristics of the unknown compounds.	After testing the known compounds, test the physical properties of each of the unknown substances.	Obtain the equipment and label all materials that will be used in the experiment.	Test 4 different physical properties for each of the known compounds.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

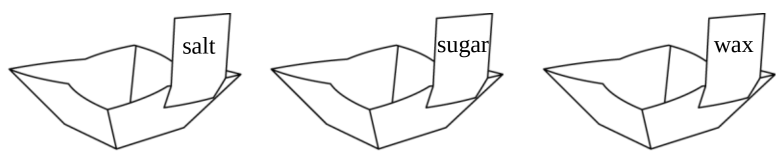
Set Up

1. ☐ Plug in the hot plate and set it to its highest setting.

2. ☐ Place four test tubes in a test tube rack. Label the test tubes “salt”, “sugar”, “wax”, and “distilled water”.

3. ☐ Create and label three aluminum foil dishes.

- **a.** Fold three pieces of aluminum foil (5-cm × 5-cm squares) into small dishes.
- **b.** Place a piece of tape on each dish and label the dishes “salt”, “sugar”, and “wax”. These dishes will eventually be placed on the hot plate, so make sure that the label is positioned so that it will not directly touch the heating surface.



4. ☐ Why is it important to label the test tubes and aluminum dishes?

Many substances have similar appearances. By labeling their containers, you avoid confusion.

5. ☐ Use a spatula to place a pea-sized sample of each substance in the appropriately labeled aluminum dish and another pea-sized sample of each substance in the appropriately labeled test tube.

Collect Data

6. ☐ Test the hardness of each compound by rubbing a small sample between your fingers. Record the hardness as either soft and waxy, or brittle and granular. Record your observations in Table 1 below. Wash your hands after testing.

Table 1: Observed physical properties of salt, sugar, and wax

Physical Property	Ionic Compound: salt (sodium chloride)	Covalent Compound Polar Molecular: sugar (sucrose)	Covalent Compound Non-polar Molecular: wax
Hardness (soft and waxy or brittle and granular)	brittle and granular	brittle and granular	soft and waxy
Melting point (high or low)	high	low	low
Soluble in water (yes or no)	yes	yes	no
Conductivity in water (µS/cm)	56,223	59	14
Conductor or non-conductor	conductor	non-conductor	non-conductor

7. ☐ Place the aluminum dishes containing the samples onto the hot plate and heat them for a maximum of three minutes. If a substance melts, use tongs to carefully remove the aluminum dish from the hot plate, allow it to cool, and record the melting point as low (in Table 1 above).

8. ☐ After three minutes of heating, turn the hot plate off. If a substance did not melt, record its melting point as high in Table 1 above.

9. ☐ Explain the types of bonds that are being overcome during the melting of ionic and covalent molecular compounds.

For ionic compounds, the ionic bonds must be overcome for the substance to melt.

For covalent molecular compounds, the intermolecular attractions must be overcome for the substance to melt.

10. ☐ Fill each of the test tubes containing the separate samples with approximately 5 mL of distilled water.

11. ☐ Stopper each test tube and gently shake the test tubes for two minutes or until dissolved.

12. ☐ Observe each test tube and record whether the substance dissolved or not. Record your observations in Table 1 above.

13. ☐ Start a new experiment on the data collection system. �(1.2)

14. ☐ Connect the conductivity sensor to the data collection system. �(2.1)

15. ☐ Configure your data collection system to monitor conductivity in a digits display. �(6.1)

16. ☐ Set the range of the conductivity to its lowest setting (0 to 1000 µS/cm) by pressing the green button marked with .

17. ☐ Test the conductivity of the distilled water by placing the conductivity sensor in the test tube filled with distilled water. Record the results below.

Conductivity of distilled water (µS/cm): 14 µS/cm

18. ☐ Test the conductivity of the three remaining samples by following the steps below:

a. If the substance did not completely dissolve, decant the solution into another test tube.

b. Place the conductivity sensor in the test tube containing the decanted liquid.

c. Start with the conductivity sensor at its lowest setting: (0 to 1000 µS/cm). If the conductivity sensor is saturated (reads 1000 µS/cm), then change to the middle setting (0 to 10,000 µS/cm). If the conductivity sensor is saturated at the middle setting (reads 10,000 µS/cm), then change to the highest setting (0 to 100,000 µS/cm).

d. Record the conductivity (µS/cm) in Table 1 above.

e. Clean the conductivity sensor using distilled water and then repeat for the next sample.

19. ☐ If the conductivity is similar to distilled water, record the sample as a non-conductor in Table 1 above. If the conductivity of the sample is much greater (100 times or more) than the distilled water, record the sample as a conductor in Table 1 above.

20. ☐ What does conductivity indicate about the molecular structure of a compound?

Conductivity occurs when there is a flow of charged particles. If a substance conducts electricity, then it must be composed of ions that are free to move.

21. ☐ Dispose of the solutions and solids and clean the glassware so that it can be used to test the unknowns.

22. ☐ Obtain the unknown samples and repeat the experiment to find the properties of each of the unknown substances. Record the results in Table 2 below.

Table 2: Observed physical properties of unknowns

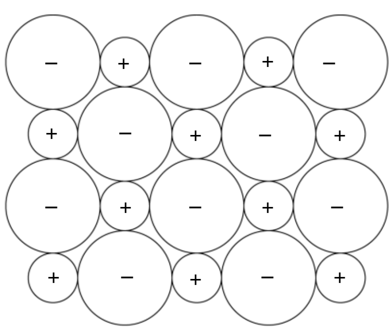
Physical Property	Unknown A	Unknown B	Unknown C
Hardness (soft and waxy or brittle and granular)	hard and brittle	soft and waxy	hard and brittle
Melting point (high or low)	low	low	high
Soluble in water (yes or no)	yes	no	yes
Conductivity in water (µS/cm)	25	32	43,011
Conductor or non-conductor	non-conductor	non-conductor	conductor

23. ☐ Clean the lab station according to the teacher's instructions.

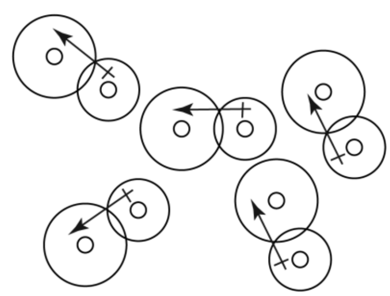
Data Analysis

1. ☐ Draw pictures that illustrate the molecular structure of each of the following:

Ionic compounds:



Molecular Polar Covalent Compounds:



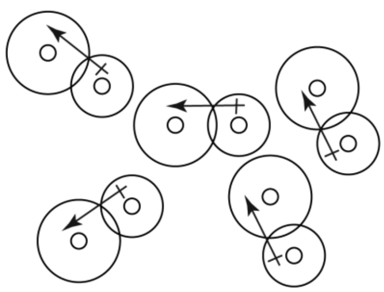
Molecular Non-polar Covalent Compounds:

2. ☐ Identify the molecular bonding (compound type) of unknown A, unknown B, and unknown C, and briefly explain the evidence supporting your decision.

Unknown	Type of Bonding	Evidence
Unknown A	polar molecular covalent compound	low melting point, soluble in water, low conductivity
Unknown B	non-polar molecular covalent compound	low melting point, insoluble in water, low conductivity
Unknown C	ionic compound	high melting point, soluble in water, high conductivity

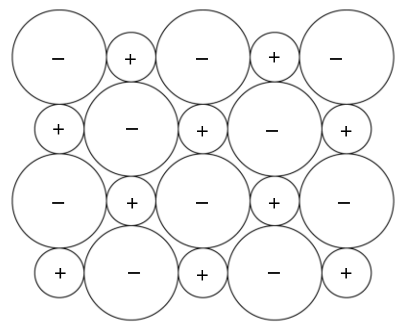
3. ☐ Draw pictures that illustrate the molecular structure of unknown A, of unknown B, and of unknown C.

Unknown A



Unknown B

Unknown C



Analysis Questions

1. How can you determine if an unknown substance is an ionic compound or a molecular covalent compound (either polar or non-polar)?

Observe the macroscopic properties of the substance. Substances that dissolve in water, conduct electricity, and have a high melting point are ionic compounds. Substances that do not dissolve in water, do not conduct electricity, and have low melting points are molecular covalent compounds.

2. What properties are the same for ionic and molecular covalent compounds (either polar or non-polar)?

Both may be hard and granular, and both may be soluble in water.

3. What is the difference between an ionic bond and an ionic compound?

Ionic bonds exist between two atoms. Ionic compounds are an entire network (lattice structure) of ionic bonding between many atoms.

4. What properties can be used to determine if a molecular covalent is polar or non-polar?

Polar covalent molecules will have high melting points and will be soluble in water while non-polar covalent molecules will have low melting points and will not be soluble in water.

Synthesis Questions

Use available resources to help you answer the following questions.

1. What are the two main chemical components of air? Predict the type of bonding for each. Explain your reasoning.

Air is approximately 78% nitrogen and 21% oxygen. The remaining 1% is other gases, including argon, carbon dioxide, and water vapor.

Both nitrogen and oxygen gases are non-polar covalent compounds. They both must have very low melting points because they exist as gases at room temperature.

2. Odor is another physical property that can be tested. Which type of compounds (ionic or molecular covalent) would you expect to have a stronger odor? Why?

Covalent compounds have lower melting points and require less energy to overcome the intermolecular forces holding the molecules together in solid form. This also applies to the intermolecular forces that must be overcome for the molecules to vaporize into individual gas molecules of the substance. The substance must be in vapor form so that the molecules can enter the nose and be smelled.

3. Oil does not dissolve in water. Based on this observation, would you classify oil as a non-polar covalent, polar covalent, or ionic compound? Explain.

Oil is a non-polar covalent compound. If oil were ionic or polar covalent, it would dissolve in water.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Which of the following is a property of an ionic compound?

- **A.** Conducts electricity in the solid state
- **B.** **Conducts electricity when dissolved in water**
- **C.** Has a low melting point
- **D.** Is soft and waxy

2. How are molecules in covalent compounds held together?

- **A.** **Intermolecular forces**
- **B.** Intramolecular forces
- **C.** Ionic bonds
- **D.** Covalent bonds

3. What substance most likely exists as a gas at room temperature?

- **A.** Ionic compound
- **B.** Polar covalent compound
- **C. Non-polar covalent compound**
- **D.** Metal

4. What substance most likely does not dissolve in water?

- **A.** Ionic compound
- **B.** Polar covalent compound
- **C. Non-polar covalent compound**
- **D.** Gaseous compound

5. What substance most likely dissolves in water but in solution does not conduct electricity?

- **A.** Ionic compound
- **B. Polar covalent compound**
- **C.** Non-polar covalent compound
- **D.** Metal

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. **Intramolecular** forces cause atoms to be attracted to one another within molecules. These forces result when the **positive** charged nuclei of two atoms are attracted to the same set of **negative** charged electrons. If the electrons are shared between the atoms, then a **covalent** bond is formed. If the electrons are taken completely by one atom, the atom becomes an ion and forms **ionic** bonds.

2. Atoms do not have the same tendencies to gain and lose electrons. **Electronegativity** is the measure of how much an atom tends toward gaining an electron. When one atom in a **covalent** bond has a greater affinity for the electron, a **polar covalent** bond is formed because the electron is **unequally** shared between the bonding atoms. If the electrons are **equally** shared, then the negative charge is distributed between the atoms equally forming a **non-polar covalent** bond.

3. When forces of attraction exist between neighboring molecules, they are called **intermolecular** forces. These forces are **weaker** than the forces holding atoms together and cause the molecules to form solids and liquids. The types of forces holding atoms and molecules together give substances their different **physical** properties, such as **melting point**, electrical **conductivity**, and **solubility** in water.

Extended Inquiry Suggestions

Investigate whether the amount of the substances tested affects the physical properties measured in this investigation (intensive versus extensive properties).

Investigate a variety of liquids to determine which is the most polar.

Ask students to draw and/or build molecules of these compounds to help them visualize the unequal distribution of charge in polar bonds and polar molecules.

Explore the effects of temperature on the solubility of an ionic substance.

7. Electrolyte versus Non-Electrolyte Solutions

Objectives

In this activity, students determine which substances in sports drinks (water, sugars, or salts) are electrolytes. Through this investigation, students:

- Differentiate an electrolyte solution from a non-electrolyte solution
- Describe the effect concentration has on the conductivity of an electrolyte solution
- Determine the approximate concentration of electrolytes in a sports drink

Procedural Overview

Students conduct the following procedures:

- Measure the conductivity of salt solutions, sugar solutions, and a sample of a sports drink
- Graph conductivity versus concentration
- Use the graph to determine the concentration of electrolytes in a sports drink

Time Requirement

♦ Preparation time	30 minutes
♦ Pre-lab discussion and activity	20 minutes
♦ Lab activity	50 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Conductivity sensor
- Test tube (6), 20-mm x 150-mm
- Beaker to collect rinse water
- Test tube rack
- Funnel
- Wash bottle filled with distilled (deionized) water
- Sucrose solutions (0.02 M, 0.04 M, 0.06 M, 0.08 M, 0.10 M), 10 mL of each concentration[1]
- Sodium chloride solutions (0.02 M, 0.04 M, 0.06 M, 0.08 M, 0.10 M), 10 mL of each concentration[2]
- Distilled (deionized) water, 50 mL
- Sports drink, 10 mL

[1] To formulate sucrose solutions using table sugar (sucrose, $C_{12}H_{22}O_{11}$), refer to the Lab Preparation section.

[2] To formulate sodium chloride solutions using table salt (sodium chloride, NaCl), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Physical properties of matter
- Ions and molecules
- Solutions

Related Labs in this Guide

Labs conceptually related to this one include:

- Properties of Ionic and Covalent Compounds
- Concentration of a Solution: Beer's Law
- Double Replacement Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Manual sampling mode with manually entered data �(5.2.1)
- Monitoring live data �(6.1)
- Starting a manually sampled data set �(6.3.1)
- Recording a manually sampled data point�(6.3.2)
- Stopping a manually sampled data set �(6.3.3)
- Displaying data in a graph �(7.1.1)
- Displaying multiple data runs on a graph �(7.1.3)
- Naming a data run �(8.2)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

We use physical properties to identify and describe matter. Physical properties of matter include color, density, odor, boiling point, melting point, solubility, conductivity, and many others. This investigation uses conductivity to differentiate between different types of matter.

Electrical conductivity is a measure of how easily electric current flows through a substance. An electric current is the flow or movement of charged particles (either electrons or ions). If a substance conducts electricity, it means the substance contains either free-moving electrons or free-moving charged particles (ions).

Electricity can conduct through both solids and liquids. Metals are good electrical conductors because metallic bonds contain delocalized electrons that can move freely throughout the solid metal. Liquids can also conduct electrical current, but only if they contain a sufficient number of ions. Electrolytes are substances that dissociate (produce ions) in solution to form electrolyte solutions. Electrolytes can be classified as strong electrolytes (substance that dissociate completely) or weak electrolytes (substances that partially dissociate). Strong electrolytes include ionic compounds, such as sodium chloride and sodium nitrate, as well as strong acids and bases, such as sulfuric acid and sodium hydroxide.

However, not all substances dissociate when they dissolve into solution. If the substance does not dissociate and produce free ions, the solution will not conduct electricity, even if the substance fully dissolves. Substances that do not conduct an electric current in solution are known as non-electrolytes. Non-electrolytes include pure water, oil, phenol, alcohol, and sugar, among others.

Determining whether or not a solution conducts electricity allows a deeper understanding of what happens at the molecular level. If a solution conducts electricity, we can conclude that the solute dissociated and ions are present in the solution. A solution that does not conduct electricity, on the other hand, does not have ions present. This means the solute dissolved as molecules.

A conductivity probe measures the conductivity of a solution. The probe contains two electrodes: a positively charged anode and a negatively charged cathode. With the probe placed in an electrolyte solution, the positive ions (cations) and negative ions (anions) move toward their oppositely charged electrodes, creating a flow of charged particles that complete the circuit. The concentration of an electrolyte solution is directly proportional to the solution's conductivity. We measure the electrical current created in units of microsiemens per centimeter (μS/cm). The more ions present in the solution, the greater the conductivity value. Extrapolating from a plot of concentration versus conductivity of samples of known concentrations, we can determine the concentration of an unknown solution from its measured conductivity value.

Pre-Lab Discussion and Activity

Metal versus Polystyrene

Use a metal pie plate and a polystyrene plate to engage the students in a discussion of lightning storms and safety. Bring up the idea of electrical circuits and their use in everyday life.

1. Which of these, a metal pie plate or a polystyrene plate, would you willingly hold over your head during a lightning storm?

Students should know the metal pie plate will conduct electricity and should not be held during a lightning storm.

Conductivity of Various Liquids

Using a conductivity tester, with a light bulb whose intensity varies with changes in conductivity, test the conductivity of the metal pie plate, polystyrene plate, distilled water, tap water, bottled spring water, and a sports drink. First, discuss which sample will cause the light bulb to be brightest (most conductivity), dimmest, and have no light at all. Next, have the students list these from brightest to dimmest either on paper or on the board. Finally, test the samples with the conductivity tester, and compare the results to the students' predications as you progress through the questions below.

2. What will happen when you place the electrodes on the conductivity tester in contact with the metal pie plate? With the polystyrene plate?

The light bulb illuminates when placed against the metal pie plate because electrons can flow through it. The light bulb remains dark when placed against the polystyrene plate because electrons are locked in bonds and cannot move. If the electrons are not free to move, electricity cannot flow.

3. Distilled water contains only H_2O molecules. Do you think distilled water will conduct electricity?

A lot of students think that all forms of water, including distilled water, conduct electricity because they have been warned about not using electricity near water. Such warnings include not using electrical appliances in the bathtub and not swimming during a lightning storm. The correct answer is no, distilled water does not conduct electricity because it does not contain ions.

4. Will tap water or bottled spring water conduct electricity? Explain your reasoning.

Both tap water and bottled spring water will slightly conduct electricity because they contain dissolved ions.

5. Explain why distilled water does not conduct electricity but tap water and bottled spring water do conduct electricity.

Freely-moving, charged particles (ions or electrons) are necessary for electricity to flow. The beaker containing distilled water does not conduct electricity because no ions are dissolved in it. Tap water and bottled spring water contain some ions and, therefore, conduct electricity.

Electrolytes

Introduce the term electrolyte. Students may be familiar with sports drinks as electrolyte replacements drinks. Electrolytes are substances that dissociate to form ions when dissolved. Non-electrolytes are substances that do not dissociate when dissolved.

6. Does tap water or bottled spring water contain electrolytes?

Both tap water and bottled spring water must contain electrolytes (dissolved ions) because they both conduct electricity.

7. Do you expect a sports drink to conduct electricity less than, the same as, or better than tap water and bottled spring water? Why?

The sports drink will conduct electricity better than tap water and bottled spring water because the sports drink has electrolytes added to it.

8. Why are electrolytes important in your body? Why must they be replaced during exercise?

Electrolytes are vital for many bodily functions. For example, the body requires electrolytes to conduct the electrical impulses along the nerve pathways between your brain and the rest of your body. Electrolytes are also vital for proper maintenance of blood volume and blood pressure. Muscles require electrolytes for proper function. Electrolytes must be replaced because they are lost through sweating during exercise.

Measuring Conductivity

The light-bulb conductivity tester provides students with qualitative data on the extent to which a solution conducts electricity. Quantitative data can be determined using a conductivity sensor. Re-test each solution, this time using a conductivity sensor, and add the quantitative conductivity values to the table written on the board. Convey that conductivity is measured in microsiemens per centimeter (µS/cm) and that small values (0 to 20 µS/cm) correspond to the light not turning on, and increasingly larger numbers correspond to an increasingly brighter light.

Construct the following table on the board, and have the students help fill it in.

Method	Distilled Water	Tap Water	Bottled Spring Water	Sports Drink
Qualitative Conductivity	dark	dim	dim	very bright
Quantitative Conductivity (µS/cm)	3	102	58	2550

9. Does the light-bulb conductivity tester provide us with qualitative or quantitative conductivity data? Explain your reasoning.

The light-bulb conductivity tester produced qualitative data because the data are descriptive words, such as ‘dark’, ‘dim’, and ‘very bright’. Quantitative data is numerical.

10. Does tap water and bottled spring water contain the same amount of electrolytes? Is qualitative or quantitative data easier to use when answering this question?

In this example, tap water has more electrolytes than the bottled spring water. This is more easily determined from the quantitative conductivity data.

11. What causes the light to illuminate? Why do different solutions produce different amounts of light?

Ions in the solution determine the brightness of the light. The more ions, the brighter the light.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. Prepare 100 mL of 0.00 M, 0.02 M, 0.04 M, 0.06 M, 0.08 M, and 0.10 M sucrose solutions. This is enough for 10 lab groups.

   a. Make 30 mL of 1.0 M sucrose solution by dissolving 10.27 g of sucrose in enough distilled water to produce 30 mL of solution.

   b. Use the 1.0 M sucrose solution to create 100 mL of each solution needed in this investigation. Use the volumes listed Table A below to create the solutions.

Table A: Serial dilution volumes

Solution Concentration (M)	Volume of 1.0 M Sucrose Solution (mL)	Volume of Distilled Water (mL)
0.00	0	100
0.02	2	98
0.04	4	96
0.06	6	94
0.08	8	92
0.10	10	90

2. Prepare 100 mL of 0.00 M, 0.02 M, 0.04 M, 0.06 M, 0.08 M, and 0.10 M sodium chloride. This is enough for 10 lab groups.

   a. Make 30 mL of 1.0 M sodium chloride (NaCl) by dissolving 1.75 g sodium chloride in enough distilled water to produce 30 mL of solution.

   b. Use the 1.0 M sodium chloride solution to create 100 mL of each solution needed in this investigation. Use the volumes listed Table A above to create the solutions.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Do not eat, drink, or taste materials in the lab.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

3	1	5	2	4
Dispose of the sodium chloride solutions, clean out each test tube, and repeat steps 1 and 2 using sucrose solutions.	Pour each sodium chloride solution into a test tube with the concentration clearly labeled.	Determine the concentration of electrolytes in the sports drink by comparing it to the conductivity versus concentration data.	Test the conductivity of each sodium chloride solution.	Dispose of the sucrose solutions, clean the test tubes, and then test the conductivity of a sample of sports drink.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Part 1 – Sodium chloride (salt) solutions

Set Up

1. ☐ Label the test tubes “0.00 M NaCl”, “0.02 M NaCl”, “0.04 M NaCl”, “0.06 M NaCl”, “0.08 M NaCl”, and “0.10 M NaCl”.

2. ☐ Use a funnel to pour 10 mL of distilled water into the test tube labeled 0.00 M.

3. ☐ Use a funnel to pour 10 mL of 0.02 M NaCl into its labeled test tube.

4. ☐ Rinse the funnel with distilled water.

5. ☐ Continue to fill each test tube with the appropriate solution, and rinse the funnel between each solution.

6. ☐ Why do you have to rinse the funnel with distilled water before using it again?

This lab deals with small amounts of electrolytes and very little conductivity. Without rinsing the funnel, the next solution will be contaminated with the previous solution. This will affect the accuracy of your results for conductivity.

7. ☐ Start a new experiment on the data collection system. �(1.2)

8. ☐ Connect the conductivity sensor to the data collection system. �(2.1)

9. ☐ Configure the data collection system to manually collect conductivity and concentration data in a table. Define concentration as a manually entered data set with units of molarity (M). �(5.2.1)

10. ☐ What do you expect to happen to the conductivity as the concentration of sodium chloride increases?

Sodium chloride should dissociate into the ions Na^+ and Cl^-. As the concentration of sodium chloride increases, the number of ions in solution will also increase. Therefore, conductivity should increase as the concentration of sodium chloride increases.

11. ☐ Identify the dependent and independent variables as well as their units used in this part of the experiment.

The dependent variable is conductivity in units of microsiemens per centimeter (μS/cm).

The independent variable is the concentration of sodium chloride in units of molarity (M).

Collect Data

12. ☐ Start a new manually sampled data set. �(6.3.1)

13. ☐ Test the conductivity of each salt solution starting with 0.00 M and moving up in concentration. To test the conductivity follow these steps:

- **a.** Set the range of the conductivity sensor to its lowest setting (0 to 1000 μS/cm) by pressing the green button marked with .
- **b.** Place the conductivity sensor in the test tube containing the sample you are testing.
- **c.** If the conductivity sensor is saturated (reads 1000 μS/cm), then change to the middle setting (0 to 10,000 μS/cm). If the conductivity sensor is saturated at the middle setting (reads 10,000 μS/cm), then change to the highest setting (0 to 100,000 μS/cm).
- **d.** Record the conductivity of the sample. �(6.3.2)
- **e.** Remove the conductivity sensor from the sample and clean the sensor by thoroughly rinsing it with distilled water.
- **f.** Repeat the steps above to test the conductivity of the next sample.

14. ☐ Why do you have to rinse the conductivity sensor with distilled water before using it again?

Without rinsing the conductivity sensor, the next solution will be contaminated with the previous solution and will affect the accuracy of your results for conductivity.

15. ☐ When you have recorded all of your data, stop the data set. �(6.3.3)

16. ☐ Copy the salt solutions conductivity data from your data collection system to Table 1 in the Data Analysis section.

17. ☐ Dispose of your solutions according to your teacher's instructions.

18. ☐ Clean all of your test tubes by rinsing them thoroughly with distilled water so that they can be used in the next part of this investigation.

Analyze Data

19. ☐ Create a graph of Conductivity (μS/cm) versus Concentration (M) on your data collection system. �(7.1.1)

20. ☐ Name the data run "salt solutions". �(8.2)

21. ☐ Explain what happens to the conductivity of the sodium chloride solution as the concentration of sodium chloride increases.

The conductivity increases as the concentration of sodium chloride increases.

Part 2 – Sucrose (sugar) solutions

Set Up

22. ☐ Label the test tubes "0.00 M sucrose", "0.02 M sucrose", "0.04 M sucrose", "0.06 M sucrose", "0.08 M sucrose", and "0.10 M sucrose".

23. ☐ Use a funnel to fill each test tube with 10 mL of the indicated sucrose solution.

24. ☐ What do you expect to happen to the conductivity as the concentration of sucrose increases?

The sucrose solutions will not conduct electricity because sucrose dissolves as molecules not ions. As the concentration of sucrose increases, there will be no change in conductivity.

25. ☐ Identify the dependent and independent variables as well as their units used in this portion of the experiment.

The dependent variable is conductivity in units of microsiemens per centimeter (μS/cm).

The independent variable is the concentration of sucrose in units of molarity (M).

26. ☐ Return to the table display on your data collection system.

Collect Data

27. ☐ Start a new manually sampled data set. �(6.3.1)

28. ☐ Record the conductivity of each salt solution. Be sure to rinse the conductivity sensor with distilled water after each test. �(6.3.2)

29. ☐ When you have recorded all of your data, stop the data set. �(6.3.3)

30. ☐ Copy the sucrose solutions conductivity data from your data collection system to Table 1 in the Data Analysis section.

31. ☐ Dispose of your solutions according to your teacher's instructions.

32. ☐ Clean all of your test tubes by rinsing them thoroughly with distilled water so that they can be used in the next part of this investigation.

Analyze Data

33. ☐ Return to your graph of Conductivity (μS/cm) versus Concentration (M) on your data collection system.

34. ☐ Name the run "sugar solutions". �(8.2)

35. ☐ Explain what happens to the conductivity of the sucrose solution as the concentration of sucrose increases.

The conductivity of the sucrose solution does not change when the concentration of sucrose increases.

Part 3 – Sports drink

Set Up

36. ☐ Use a funnel to fill a test tube approximately halfway with a sports drink (at room temperature).

37. ☐ Why does the procedure specify that the sports drink should to be at room temperature?

Temperature may have an effect on conductivity. You can eliminate this variable by ensuring that all of the solutions are at the same temperature.

38. ☐ Configure your data collection system to monitor live conductivity data in a digits display. �(6.1)

Collect Data

39. ☐ Place the conductivity sensor in the test tube containing the sports drink and allow the conductivity reading to stabilize.

40. ☐ Record the brand and flavor of sports drink tested and its conductivity below.

Sports Drink: Brand "X", lemon-lime____________________________

Conductivity (µS/cm): 2210 µS/cm___________________________________ _

41. ☐ Save the data file and clean up your lab station according to the teacher's instructions. �(11.1)

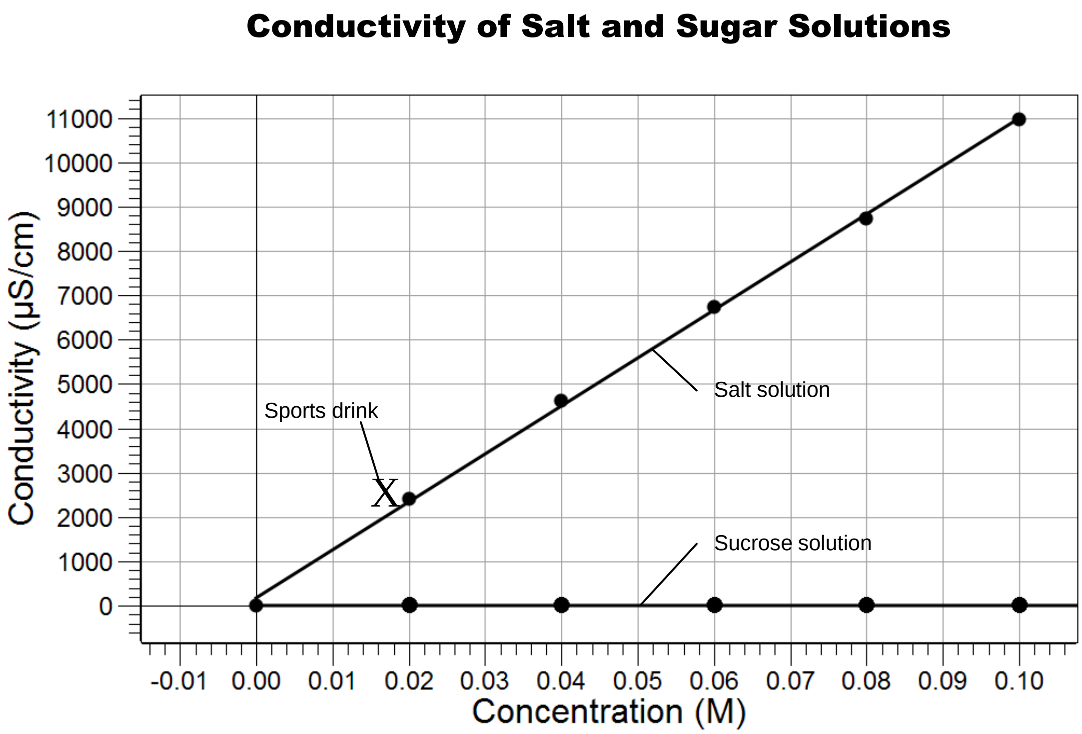
Data Analysis

Table 1: Measured conductivity of salt and sucrose solutions

Concentration of Salt and Sugar Solutions (M)	Conductivity of Salt Solutions (µS/cm)	Conductivity of Sucrose Solutions (µS/cm)
0.00	12	13
0.02	2,414	12
0.04	4,629	12
0.06	6,737	12
0.08	8,740	13
0.10	10,962	13

1. ☐ Return to your graph of Conductivity (µS/cm) versus Concentration (M) on your data collection system and display both the salt solutions data set and the sucrose solutions data set on the graph. �(7.1.3)

2. ☐ Plot or print a graph of Conductivity (µS/cm) versus Concentration (M). Include both the salt solutions data set and the sucrose solutions data set on the same set of axes (clearly label each). Label the overall graph, the x-axis, the y-axis, and include units on the axes.



3. ☐ Where does the sports drink you tested fit on the graph of Conductivity (µS/cm) versus Concentration (M) above? Place an "X" on the graph to mark this location and label it "Sports drink".

Analysis Questions

1. Explain the difference between an electrolyte solution and a non-electrolyte solution.

An electrolyte solution conducts electricity, and the conductivity increases as the concentration increases.

A non-electrolyte solution does not conduct electricity (or is a very poor conductor).

2. Which compounds in a sports drink are electrolytes? Which compounds in a sports drink are non-electrolytes?

The salts are electrolytes. The sugars and water are non-electrolytes.

3. What is the approximate concentration of electrolytes in the sports drink you tested? Explain how you determined this concentration.

Answers will vary (for the sample data, the concentration is between 0.01 and 0.02 M). The concentration can be estimated from the conductivity versus concentration graph.

4. Which effect does concentration have on the conductivity of an electrolyte solution?

Conductivity increases as the concentration of ions in solution increases.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Explain the difference at the molecular level between what happens when salt and sugar dissolve.

Salt dissociates into ions when it dissolves. Sugar dissolves as molecules.

2. Does crystalline table salt (sodium chloride) conduct an electric current?

No. In order to conduct electricity, there must be free-moving ions. In the solid state, the ions that make up sodium chloride are locked into position and cannot move. Because the electrons cannot move, they cannot conduct electricity.

3. Which type of compounds (ionic or covalent) generally makes better electrolyte solutions? Why?

Ionic compounds are generally better electrolytes because they dissociate into ions when dissolved in water.

4. Is human blood an electrolyte or non-electrolyte solution? Explain your answer.

Human blood is an electrolyte solution. Human blood contains ions such as sodium, potassium, calcium, chloride, hydrogen carbonate, and phosphate.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. What are solutions that conduct electricity called?

- **A. Electrolyte solutions**
- **B.** Non-electrolyte solutions
- **C.** Dilute solutions
- **D.** Concentrated solutions

2. What do solutions that conduct electricity contain that differentiates them from solutions that do not conduct electricity?

- **A.** Molecules
- **B.** Atoms
- **C. Free-moving ions**
- **D.** Solutes

3. Which of the following solutions would conduct electricity the best?

- **A.** A sugar solution
- **B.** **A salt solution**
- **C.** Distilled water
- **D.** All of the above

4. What would you expected to happen as the concentration of electrolytes in a solution increases?

- **A.** **An increase in the conductivity of the solution**
- **B.** A decrease in the conductivity of the solution
- **C.** A decrease followed by an increase in the conductivity of the solution
- **D.** The conductivity of the solution will remain the same

5. What forms of sodium chloride will conduct electricity?

- **A.** In a solid crystalline state
- **B.** In a molten liquid state
- **C.** When dissolved in water
- **D.** **Both B and C**
- **E.** A, B, and C will all conduct electricity

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. We use physical properties to observe and describe matter. **Physical properties** of matter include color, density, odor, boiling point, melting point, solubility, and conductivity (among others). This investigation uses conductivity to differentiate between different types of matter. **Conductivity** is the ability of an electric current to pass through a substance. An electric current is the flow or movement of charged particles (either electrons or ions).

2. Just as some solid substances, like metal wires, are good conductors of electricity, some liquids can also conduct electrical current. Such liquids are called **electrolyte solutions**. In order to be conductive, a solution must contain charged particles, or **ions**, which are free to move. Ions are formed if a solute **dissociates**, or splits into ions, when dissolved in a solvent (usually water). The free-moving ions enable an electric current to pass through the solution.

3. Not all substances ionize when they **dissolve** into solution. Many substances dissolve as molecules. Because there are no charged particles, the solution will not conduct electricity. Solutions that do not conduct an electric current are known as **non-electrolyte solutions**. Substances such as pure water, oil, phenol, and alcohol are examples of substances that are non-electrolytes.

Extended Inquiry Suggestions

Determine the concentration of sodium chloride in contact lens saline solutions. Do different brands have the same concentration?

Investigate reasons why the conductivity of freshwater lakes, streams, or ponds could change over time.

Compare various brands and flavors of sports drinks.

Compare various brands of bottled water.

Test the effect of temperature on conductivity.

Test water samples from various locations (tap water from different parts of the school or city).

Test environmental water samples from various locations or sources.

8. Boyle's Law

Objectives

Determine the effect of volume on the pressure of a closed system containing a fixed amount of molecules at a constant temperature. Through this investigation, students:

- Show that an inversely proportional relationship exists between pressure and volume for a gas at a constant temperature
- Apply Boyle's law $PV = k$ and $P_1V_1 = P_2V_2$
- Differentiate between real and ideal gases

Procedural Overview

Students conduct the following procedures:

- Use an absolute pressure sensor to determine the pressure of a fixed sample of air molecules at various volumes
- Collect data in replicate and calculate averages
- Graph volume versus pressure of a fixed amount of gas molecules at a constant temperature

Time Requirement

Preparation time	10 minutes
Pre-lab discussion and activity	30 minutes
Lab activity	30 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Absolute pressure sensor
- Sensor extension cable
- Syringe, 20-mL or 60-mL[1]
- Tubing, 1- to 2-cm[1]
- Quick-release connector[1]
- Glycerin, 2 drops

[1]Included with most PASCO absolute pressure sensors

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Kinetic molecular theory
- ♦ Pressure

Related Labs in This Guide

Labs conceptually related to this one include:

- ♦ Percent Oxygen in Air
- ♦ Gay-Lussacs's Law and Absolute Zero
- ♦ Ideal Gas Law

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- ♦ Starting a new experiment on the data collection system �(1.2)
- ♦ Connecting a sensor to the data collection system �(2.1)
- ♦ Manual sampling mode with manually entered data �(5.2.1)
- ♦ Starting a manually sampled data set �(6.3.1)
- ♦ Recording a manually sampled data point�(6.3.2)
- ♦ Stopping a manually sampled data set �(6.3.3)
- ♦ Displaying data in a graph �(7.1.1)
- ♦ Adjusting the scale of a graph �(7.1.2)
- ♦ Showing and hiding connecting lines between data points �(7.1.8)
- ♦ Manually entering data into a table �(7.2.3)
- ♦ Showing and hiding data runs in a table �(7.2.4)

- Applying a curve fit �(9.5)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

The kinetic molecular theory explains that particles in all forms of matter are in constant motion. In solids and liquids, the atoms or molecules are moving and are positioned very close to each other. In gases, the atoms or molecules are also moving, but they are positioned far apart from each other and thus have properties different from those of solids or liquids.

The kinetic molecular theory of gases makes a few basic assumptions. First, it assumes that gases are made up of small, hard particles that are separated from each other. Because the particles are far apart, their volume is considered insignificant and there are no attractive or repulsive forces between the gas particles. The large amount of empty space between the gas particles explains why gases, unlike solids or liquids, are compressible. The lack of attractive or repulsive forces explains that gases take the shape and volume of containers they occupy.

The second assumption is that gas particles are in constant, random motion. The gas particles move independently from one another in straight lines and only change direction when they collide with other particles or the walls of a container. Gas pressure is related to how often and how hard gas molecules collide with surfaces.

The third assumption is that the collisions between gas particles are elastic. This means that no energy is lost when particles collide. The total kinetic energy remains the same before and after collisions. A gas that behaves according to all of the assumptions is called an ideal gas. Although there is no such thing as an ideal gas, at many conditions of temperature and pressure, real gases behave ideally.

There are four variables that are used to describe a gas. These variables are volume (V), pressure (P), temperature (T), and number of moles (n).

Robert Boyle discovered the mathematical relationship between gas pressure and volume. Boyle's law states that for a given number of gas molecules at constant temperature, the volume of the gas varies inversely with pressure. Because it is an inversely proportional relationship, the product of the pressure (P) and volume (V) is constant (k); as one variable increases, the other must decrease. Boyle's law can be represented by the following two equations:

1) $PV = k$

2) $P_1V_1 = P_2V_2$.

When a given number of gas molecules at constant temperature are forced into a smaller volume, the pressure will increase. The increase in pressure is due to an increase in collisions between the gas particles and the container. The number of collisions increases because it takes less time for the particles to travel across a smaller space.

Boyle's law, like all the gas laws, holds true only for ideal gases. At higher pressures and lower volumes, the assumption that gas particles are so far apart that their volume is considered insignificant no longer holds true, and the gases therefore do not follow Boyle's law as well. The closer the atoms are forced together, the more important their volume becomes. As the gas molecules become even closer, attractive forces also begin to take effect. This is what enables gases to condense into liquids.

Pre-Lab Discussion and Activity

Directly Proportional (Linear) Relationships

Using two bags of candy, one large and one small, engage students in a discussion about comparing the amount of candy one buys from a candy bin compared to the total price of the bag of candy. Explain that the term directly proportional is used to describe a linear relationship between two quantities that vary in the same way, as one variable increases, the other increases as well (with one variable being a constant multiple of the other).

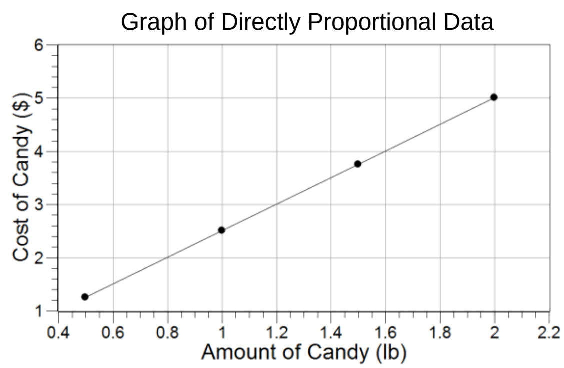
1. How does the amount of candy you buy affect the price of your bag of candy?

The more candy you buy, the more expensive the bag of candy will be.

2. What does a graph of directly proportional data look like? Does a line or a curve fit the data better?

Example table for points to plot on the graph

Amount of Candy	Cost of Candy
0.5 pounds	$1.25
1.0 pounds	$2.50
1.5 pounds	$3.75
2.0 pounds	$5.00



A line fits the data better. The slope is constant. In this example the slope is positive, but it is possible to have a directly proportional relationship with a negative slope.

Teacher Tip: You may further explain that the slope of the line is equal to the cost of candy per pound. In this case the candy costs $2.50 per pound.

Inversely Proportional Relationships

Engage students in a discussion about how the number of students affects the time it takes for 100 pieces of candy to be eaten. Explain that the term inversely proportional is used to describe two quantities that vary in opposite directions. If one quantity increases, the other quantity decreases. When two variables are inversely proportional, their product is always a constant.

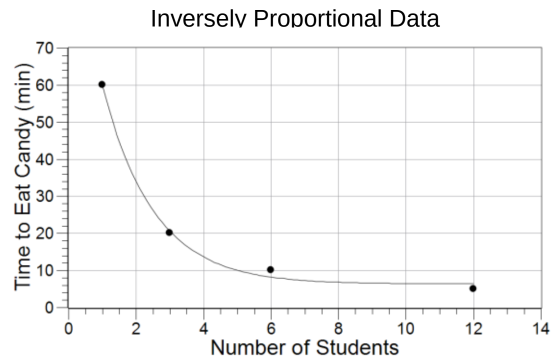
3. How does the number of students eating candy affect the time it takes to eat 100 pieces of candy?

The larger the number of students there are eating the candy, the less time it will take for the candy to be eaten.

4. What does a graph of inversely proportional data look like? Does a line or a curve fit the data better?

Example table for points to plot on the graph

Number of Students Eating Candy	Time it Takes to Eat 100 Pieces of Candy
1	60 min
3	20 min
6	10 min
12	5 min

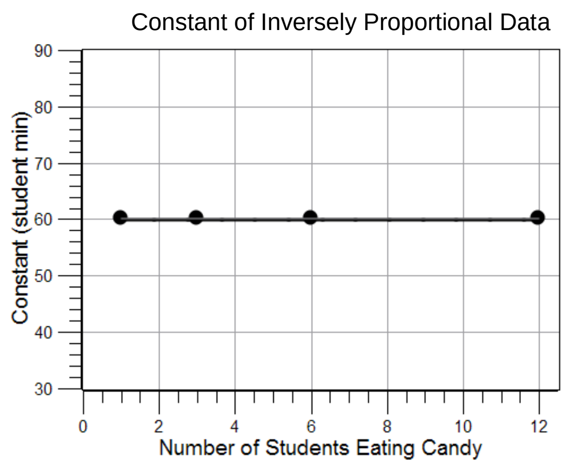


A curve fits the data better. The slope of the line is constantly changing.

5. The product of inversely proportional variables is constant. What does a graph of inversely proportional products look like? Describe the slope of the line.

Example table for points to plot on the graph

Number of Students Eating Candy	Time it Takes to Eat 100 Pieces of Candy	Constant (student·min)
1	60 min	60 student·min
3	20 min	60 student·min
6	10 min	60 student·min
12	5 min	60 student·min



The slope of the line is 0 because the product of the two inversely proportional variables is constant. In this example, students × time = constant = 60 student·min.

Marshmallow Faces in a Syringe

Provide a mini-marshmallow and a syringe to each group of students. Instruct the students to draw a face on the marshmallow, put the marshmallow inside the syringe, hold a finger over the tip of the syringe, and observe what happens to the marshmallow as they slowly depress and release the plunger. Engage the students in a discussion about their predictions and then have them explain their observations by describing the behavior of the molecules inside the syringe.

6. What will happen to the size of the marshmallow (and its face) when the plunger is depressed?

There will be a directly proportional relationship between the marshmallow's size and the volume inside the syringe. The marshmallow will become smaller and the face will shrink as the volume inside the syringe becomes smaller.

7. What did you observe?

The marshmallow became smaller and the face shrunk when the plunger was depressed. When the plunger was released the marshmallow became a little larger, but did not return to its original size.

It became difficult to push the plunger as the volume inside decreased.

8. Explain these observations by describing the behavior of the molecules inside the marshmallow and the syringe.

Marshmallows are mostly made of whipped sugar. They are soft and spongy because they contain a lot of air mixed throughout the sugar. The area inside the syringe and around the marshmallow contains particles of air. As the syringe is depressed the air particles are forced closer together causing the marshmallow to get smaller. When the plunger is released, the air particles spread back out to take up the complete area of the syringe. The marshmallow does not go back to its original size because the air molecules that were forced out of the marshmallow, do not go back inside.

The plunger gets more difficult to depress because the molecules are colliding with the plunger more often (there is an increase in pressure).

Kinetic Molecular Theory of Ideal versus Real Gases

Review the kinetic molecular theory, states of matter, and the differences between ideal and real gases with the students.

9. Compare and contrast solids, liquids, and gases.

Solids, liquids, and gases are all made up of particles that are constantly moving. In solids, the particles are held together tightly by attractive forces giving solids a fixed shape. Like solids, the particles in a liquid are very close together and have a definite volume. The particles in a liquid are not held together as tightly by the attractive forces and are therefore able to flow and take the shape of their container. Particles in a gas are very far apart and are not attracted to each other at all. This allows gases to both fill the entire volume and take the shape of their container.

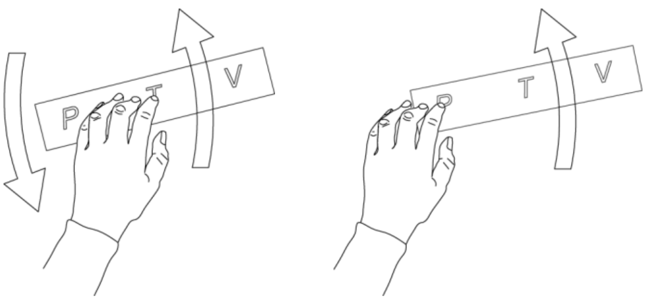
10. What would happen inside the syringe if you were able to continue compressing the plunger?

The gas particles would eventually touch one another and could not be compressed any further. The volume of the syringe would stop at the point equal to the volume of the particles all packed together with no space between them. The attractive forces between the molecules would cause the molecules to stick to each other forming a liquid.

11. As the plunger becomes more difficult to depress, is the gas inside behaving more or less like an ideal gas? Why?

The particles are getting closer together, and the volume of the particles is becoming significant. The smaller distances between the particles are allowing the attractive forces to take hold. This results in the gas behaving less like an ideal gas and more like a real gas.

Teacher Tip: A method for helping students to remember the relationships between pressure, volume, and temperature is to have the students write on a strip of paper "*P T V*" (as shown). Demonstrate to the students how to place a finger on the variable that is being held constant and push up or down on the variable that is increasing or decreasing. They can then observe what the other variable will do. You may wish to do this using an overhead projector with "*P T V*" written on a piece of transparency.



Lab Preparation

Although this activity requires no specific lab preparation, allow 10 minutes to gather the equipment needed to conduct the lab.

Safety

Add this important safety precaution to your normal laboratory procedures:

♦ To minimize the risk of injury or damage to the equipment, avoid over-compressing the air in the syringe.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	5	4	3
While viewing a data table display, connect a syringe, with 20 mL of air in it, to the absolute pressure sensor.	Connect the absolute pressure sensor to the data collection system and set up manual sampling.	Calculate the average pressure at each volume. Draw conclusions about the relationship between pressure and volume.	Collect two additional runs of absolute pressure data at each volume.	Record the absolute pressure at 20 mL, 18 mL, 16 mL, 14 mL, 12 mL, 10 mL, 8 mL, and 6 mL.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

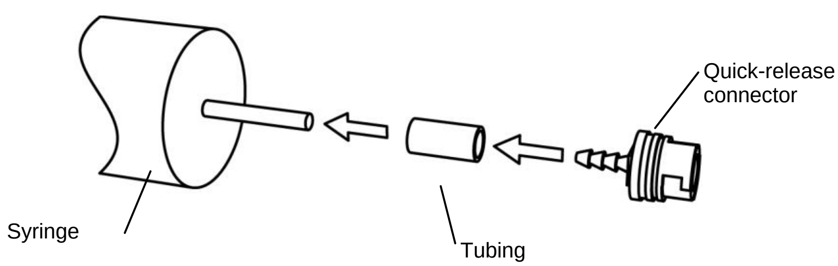
Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect the absolute pressure sensor to the data collection system using a sensor extension cable. �(2.1)

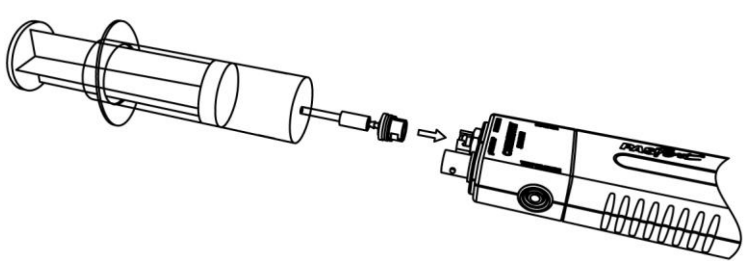
3. ☐ Configure the data collection system to manually collect pressure (kPa) and volume (mL) data in a table. Define volume as a manually entered data set with units of milliliters (mL). �(5.2.1)

4. ☐ Connect the syringe to the quick-release connector using a 1- to 2-cm piece of tubing. Put a drop of glycerin on the connecting pieces as necessary.



5. ☐ Withdraw the plunger so there is 20 mL of air in the syringe.

6. ☐ Insert the quick-release connector to the port of the absolute pressure sensor and then turn the connector clockwise until the fitting “clicks” (about one-eighth turn).



7. ☐ What do you expect to happen to the pressure as the volume decreases?

The pressure should increase as the volume decreases.

8. ☐ List the dependent variable, the independent variable, and the units for each used in this experiment.

The dependent variable is pressure (*P*) measured in units of kilopascals (kPa).

The independent variable is volume (*V*) measured in units of milliliters (mL).

Collect Data

9. ☐ Start a new manually sampled data set. �(6.3.1)

10. ☐ Ensure the syringe volume is still set to 20 mL, record the pressure, and enter the volume. �(6.3.2)

11. ☐ Record the pressure at each of the following manually entered volumes: 18 mL, 16 mL, 14 mL, 12 mL, 10 mL, 8 mL, and 6 mL. �(6.3.2)

CAUTION: To minimize the risk of injury or damage to the equipment, avoid over-compressing the air in the syringe.

12. ☐ When you have recorded all your data, stop the data set. �(6.3.3)

13. ☐ Remove the syringe from the absolute pressure sensor using the quick-release connector.

14. ☐ Set the syringe plunger to 20 mL and re-connect the syringe to the absolute pressure sensor.

15. ☐ Start a new manually sampled data set. �(6.3.1)

16. ☐ Record the pressure at each of the indicated volumes. �(6.3.2)

17. ☐ When you have recorded all your data, stop the data set. �(6.3.3)

18. ☐ Collect a third data set.

19. ☐ Explain when and why it becomes difficult to depress the syringe.

It is difficult to depress the syringe at low volumes such as 8 mL and 6 mL. It is difficult because the air molecules are colliding with the syringe more frequently because they are trapped in such a small volume.

20. ☐ Why is it important to collect more than one set of data?

Collecting multiple sets of data ensures that you are able to reproduce the same results. Reproducibility is necessary to ensure reliable data.

21. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

Table 1: Measured pressure data and calculated averages

V (mL)	$P_{\text{Set 1}}$ (kPa)	$P_{\text{Set 2}}$ (kPa)	$P_{\text{Set 3}}$ (kPa)	P_{Average} (kPa)	k_{Average} (kPa·mL)
20	101	101	101	101	2020
18	113	113	112	113	2034
16	127	126	126	126	2016
14	145	144	143	144	2016
12	168	166	167	167	2004
10	197	195	197	196	1960
8	246	242	242	243	1944
6	321	325	324	323	1938

1. ☐ Copy the three sets of pressure (P) data collected from your data collection system to Table 1 above.

Note: The previous runs of data collected can be viewed by changing the data set that is displayed. �(7.2.4)

2. ☐ Calculate the average pressure for each volume and record them in Table 1 above.

$$P_{\text{Average}} = \frac{\left(P_{\text{set}1} + P_{\text{set}2} + P_{\text{set}3}\right)}{3}$$

3. ☐ Enter the P_{Average} values into a data table on your data collection system. �(7.2.3)

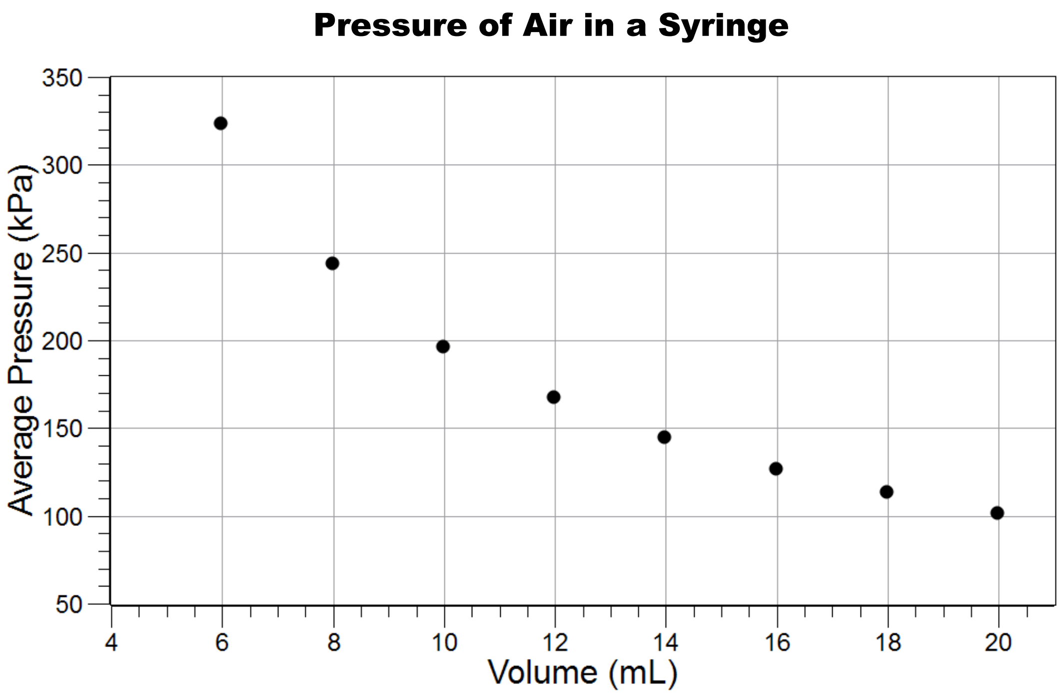
4. ☐ Create a graph of P_{Average} (kPa) versus Volume (mL) on your data collection system. �(7.1.1)

Note: To graph a scatter plot of the data points you may choose to hide the connecting lines between data points. �(7.1.8)

5. ☐ Are pressure and volume directly proportional or inversely proportional? How do you know?

Volume and pressure are inversely proportional. The data fits a curve better than a straight line because as the volume decreased the pressure increased faster and faster.

6. ☐ Plot or print a graph of Average Pressure (kPa) versus Volume (mL). Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



7. ☐ Use the average pressure (P_{Average}) to calculate the constant k_{Average} for each volume and record your results in Table 1 above.

$$P_{\text{Average}}V = k_{\text{Average}}$$

8. ☐ What should the graph of a constant (k) versus volume look like? What type of gases will display this behavior?

A constant should not change and therefore the graph should show a horizontal line. Ideal gases will display this behavior.

9. ☐ Enter the k_{Average} values into a data table on your data collection system. �(7.2.3)

10. ☐ Create a graph of k_{Average} versus Volume (mL) on your data collection system. �(7.1.1)

Note: To graph a scatter plot of the data points you may choose to hide the connecting lines between data points feature. �(7.1.8)

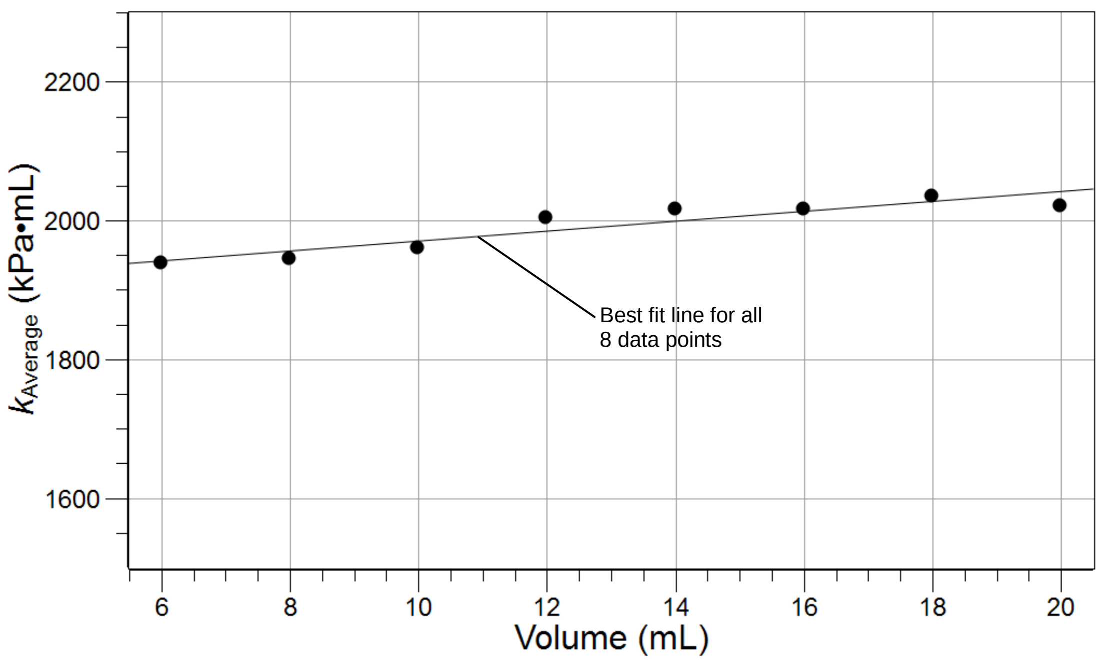
11. ☐ Adjust the scale of the y-axis so that it ranges between 1500 to 2300 kPa•mL. �(7.1.2)

12. ☐ Apply a linear fit to the data. � (9.5)

13. ☐ Plot or print a copy of the graph of Average Pressure (kPa) versus Volume (mL) including the linear line of best fit. Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

Teacher Tip: The graph is used to determine if the gas is behaving ideally or not. A perfectly ideal gas will have a completely horizontal line (slope of 0). Since the data points at the smaller volumes deviate from the approximately-horizontal line established by the larger volumes, the gas is not behaving as an ideal gas at these smaller volumes, whereas at the larger volumes, the behavior is closer to ideal.

Boyle's Law Constant



Analysis Questions

1. Does your data support Boyle's law that *PV* = *k* (pressure × volume = constant)? Explain any discrepancies in your data.

Boyle's law states that *P* × *V* should equal a constant. The data shows that the product of *P* × *V* is constant at the larger volumes, but at the smaller volumes, it deviated from the constant. This occurred because Boyle's law is for ideal gases, but air is a real gas. The deviation at smaller volumes is expected because ideal gas molecules are assumed to have no volume. However, with a real gas, as the volume of the container decreases, the volume that the gas molecules actually do occupy becomes more significant.

2. Does air behave more like an ideal gas or a real gas in different regions of the graph? Explain your analysis.

Air behaves like an ideal gas at low pressure and high volume.

Air behaves like a real gas at high pressure and low volume. It was at low volumes (6 mL and 8 mL) that there was a larger change in the constant (*P* × *V*).

3. What is the difference between an ideal and a real gas?

An ideal gas is assumed to have negligible volume and a real gas has volume. At high pressure (low container volume), the actual volume of the gas becomes significant and therefore does not follow Boyle's law.

4. What is the constant (*k*) for air as revealed by your data?

The value should be the average of all the runs. For the sample data provided in this lab: k = 1990 kPa·mL.

5. Calculate the pressure you would expect at 15.0 mL. Show your work.

$$PV = k$$

$$P = \frac{k}{V}$$

$$P = \frac{1990 \text{ mL·kPa}}{15.0 \text{ mL}} = 132.67 \text{ kPa} = 133 \text{ kPa}$$

Synthesis Questions

Use available resources to help you answer the following questions.

1. Explain why it is possible to write Boyle's law as both $PV = k$ and $P_1V_1 = P_2V_2$.

When given one set of conditions, P_1V_1 equals a constant. When the conditions change, the new conditions, P_2V_2, must equal the same constant. Since P_1V_1 equals the same constant as P_2V_2, then P_1V_1 must equal P_2V_2.

P_1V_1 = 20 mL x 100 kPa = 2000 kPa·mL

P_2V_2 = 16 mL x 125 kPa = 2000 kPa·mL

$P_1V_1 = P_2V_2$

2. How could you change the experimental design to get results more consistent with an ideal gas?

Use a larger syringe. Real gases behave ideally at low pressures.

Replace air with a gas that contains smaller particles such as helium.

3. A helium balloon is released into the atmosphere. As it rises, atmospheric pressure decreases. What do you expect will happen to the volume of the balloon?

The volume of the balloon will increase because pressure and volume are inversely proportional.

4. A cylinder containing 250 mL of a gas has a pressure of 350 kPa. If the gas was compressed to a volume of 45 mL what would the pressure change to?

P_1 = 350 kPa P_2 = ?

V_1 = 250 mL V_2 = 45 mL

$P_1V_1 = P_2V_2$

$$P_2 = \frac{P_1V_1}{V_2} = \frac{(350 \text{ kPa})(250 \text{ mL})}{45 \text{ mL}} = 1944.444 \text{ kPa} = 1900 \text{ kPa}$$

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

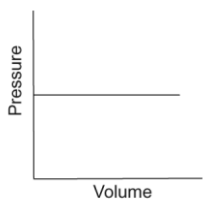
1. What conditions will cause the volume of a gas to decrease?

- **A.** An increase in the amount of gas
- **B.** An increase in temperature
- **C.** **An increase in pressure**
- **D.** A decrease in pressure

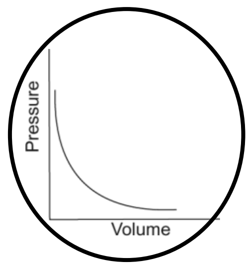
2. At constant temperature, the relationship between the volume (*V*) of a gas and its pressure (*P*) is

- **A.** $V = (\text{constant})P$
- **B.** $P = (\text{constant})V$
- **C.** $\boldsymbol{PV} = \textbf{constant}$
- **D.** $V/P = \text{constant}$

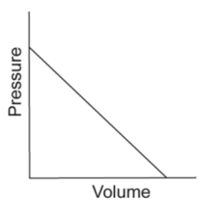
3. Which graph shows the relationship between the pressure and volume of nitrogen gas at a constant temperature?



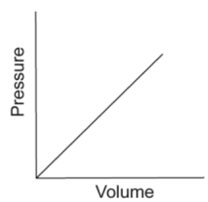
A.



B.



C.



D.

4. At room temperature, air behaves like an ideal gas at

- **A.** **Low pressures**
- **B.** High pressures
- **C.** All pressures
- **D.** Air never behaves ideally

5. A gas contained in a 3.0 L cylinder has a pressure of 120 kPa. What will the new volume be if the pressure is increased to 240 kPa?

- **A. 1.5 L**
- **B.** 3.0 L
- **C.** 4.5 L
- **D.** 6.0 L

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Answers section.

1. Gases can be **compressed** into smaller volumes because there is **a lot of** space between gas particles. Solids and liquids, on the other hand, have fixed **volumes** and cannot be compressed. When a fixed amount of gas is forced into a smaller volume, the pressure of the gas will **increase**. This happens because volume is **inversely** proportional to pressure. This relationship is known as **Boyle's** law.

2. An ideal gas is composed of a collection of perfectly hard **spheres** that are so far apart, their **volume** is assumed to be insignificant. The gas particles move in constant **random** motion and only change direction when they **collide** with another particle or the walls of the container. The collision of gas particles with a surface creates **pressure**. Real gases can behave ideally at certain temperatures and pressures. However, real gases are not ideal because they actually do have volume. Their volume is considered significant at **high** pressures.

Extended Inquiry Suggestions

Repeat the experiment using a different known gas in the syringe. Compare the results with air.

Conduct the experiment at two different temperatures and compare the constant value. On a particularly hot or cold day the students could do the experiment outside and compare it to the results of conducting the experiment inside.

Determine the effect of using different lengths of tubing to connect the syringe to the pressure sensor on the pressure × volume constant.

Determine the effect of using different sizes of syringes on the product of the pressure and volume.

Investigate other gas laws, such as determining the effect of temperature on volume at constant pressure (Charles' law), or determining the effect of temperature on pressure at constant volume (Gay-Lussac's law).

9. Gay-Lussac's Law and Absolute Zero

Objectives

Determine the temperature at which all motion stops. Through this investigation, students:

- Show that a directly proportional relationship exists between pressure and temperature for a gas at constant volume
- Relate the Kelvin temperature scale to the motion of molecules
- Show that Gay-Lussac's law, $\frac{P}{T} = k$ and $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, is valid only with Kelvin temperatures

Procedural Overview

Students conduct the following procedures:

- Use an absolute pressure sensor to determine the pressure of a fixed volume of air at various temperatures
- Record collected temperature data in both Celsius and Kelvin units
- Graph pressure versus temperature data showing the line of best fit (linear regression)
- Use the equation for the line of best fit to determine absolute zero

Time requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	45 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Absolute pressure sensor
- Fast response temperature sensor
- Sensor extension cable
- Quick-release connector[1]
- Tubing connector[1]
- Tubing, 1- to 2-cm[1]
- Test tube, 15-mm × 100-mm
- One-hole stopper to fit the test tube
- Beakers (2), 250-mL
- Ring stand
- Three-finger clamp
- Glycerin, 2 drops
- Polystyrene cups (2)
- Rubber band
- Crushed ice, 300 mL
- Room temperature water, 300 mL
- ~45 °C water, 300 mL[2]
- ~55 °C water, 300 mL[2]
- ~65 °C water, 300 mL[2]

[1] Included with most PASCO absolute pressure sensors.

[2] Refer to the Lab Preparation section for tips on how to manage hot water baths in the classroom.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Kinetic molecular theory
- Temperature as the average kinetic energy of the molecules of a substance
- Pressure

Related Labs in This Guide

Labs conceptually related to this one include:

- Percent Oxygen in Air
- Boyle's Law
- Ideal Gas Law

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)

- Connecting multiple sensors to the data collection system �(2.2)
- Manual sampling mode without manually entered data �(5.2.2)
- Changing the units of a measurement �(5.3)
- Starting a manually sampled data set �(6.3.1)
- Recording a manually sampled data point�(6.3.2)
- Stopping a manually sampled data set �(6.3.3)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Showing and hiding connecting lines between data points �(7.1.8)
- Displaying data in a table �(7.2.1)
- Applying a curve fit �(9.5)
- Finding the slope and intercept of a best-fit line �(9.6)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

There are four variables used to describe a gas: volume (V), pressure (P), temperature (T) and the number of moles (n).

In 1802, the French chemist Joseph Gay-Lussac discovered the mathematical relationship between the pressure of a gas and its temperature in Kelvin. Gay-Lussac's law states that for a given number of gas molecules at constant volume, the pressure of the gas varies directly with its temperature. Because it is a direct relationship, the pressure (P) divided by the temperature in Kelvin (T) is constant (k); as one variable increases, the other must also increase. Gay-Lussac's law can be represented by the two equations listed below.

1) $$\frac{P}{T} = k$$

2) $$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A graph of pressure versus temperature will show a directly proportional relationship. When a given number of gas molecules at constant volume are heated, the pressure will increase. Conversely, when a gas is cooled, the pressure will decrease. The decreased pressure is due to a decrease in both the number of collisions as well as the force of each collision between the gas particles and the walls of the container. The number of collisions decreases, because the particles are moving slower, so it takes more time for the particles to travel across the container. The particles collide with less force because they have less momentum as a result of having less kinetic energy.

If it were possible to continue removing kinetic energy by cooling, a point should be reached where the gas particles would be so cold they would have no kinetic energy at all. If the particles have zero kinetic energy, they are no longer moving and, therefore, do not possess momentum. Without kinetic energy, the particles cannot collide with the walls of the container and create pressure.

The theoretical point at which kinetic energy has been completely removed from a system (particles stop moving and the pressure equals zero) is known as absolute zero. Because temperature is a measurement of molecular motion (kinetic energy), the Kelvin temperature scale sets the point of no kinetic energy as its starting point, zero (0) K. This means there are no negative values on the Kelvin scale; you cannot get slower than stopped.

It is important for students to realize that at 0 °C, the gas molecules still have a great deal of kinetic energy and, therefore, a pressure not equal to zero. The Celsius scale uses a different set of reference points: the freezing and boiling points of pure water at a pressure of 1.00 atm. The size of the divisions between the Celsius degrees was transferred to the Kelvin scale, which means an increase of 1 °C is the same as an increase of 1 K— the two scales just have different starting points. This results in an easy conversion between the units: K = °C + 273.15, thus 0 K is –273.15 °C, 298.15 K is 25 °C, 0 °C is 273.15 K, and 100 °C is 373.15 K. (Incidentally, there exists an analogous absolute scale based on Fahrenheit divisions called the Rankine temperature scale.)

Gay-Lussac's law requires that temperature measurements be strictly in Kelvin. This is true for two reasons. First, there exists the 273.15 scalar difference between Kelvin and Celsius. To use Celsius in Gay-Lussac's law would introduce an offset proportional to this difference. Second, the Kelvin temperature scale is an absolute scale (all values are positive). The Celsius scale allows for negative temperature values. The values 25 °C and –25 °C would return the same value for pressure using Gay-Lussac's law, but with opposite sign; however, we know this cannot be true since pressure, just like Kelvin, cannot be negative in an absolute sense.

Pre-Lab Discussion and Activity

Gas Pressure in a Syringe

Review the kinetic molecular theory of gases with your students.

Connect an absolute pressure sensor and data collection device to a syringe and project a digits display of pressure (kPa) for your students to see. Warm a sample of air by holding the syringe in the palm of your hand and observe the corresponding change in pressure; then cool the syringe with an ice cube. Remind students of the connection between temperature and kinetic energy. Explain that continuing to cool the gas will remove more and more kinetic energy. Eventually, the system will run out of kinetic energy and the particles will stop moving. This point is called absolute zero. Because the particles are no longer moving, they cannot collide with the walls of the container to produce any pressure; the pressure is zero at absolute zero. This value has never been achieved, but scientists have come extremely close!

1. What causes the air pressure at the molecular level?

Air pressure is caused by air molecules colliding with a surface.

2. What happens to the air pressure in the syringe when the syringe is placed in the palm of my hand? Why?

The air pressure increases because heat is being transferred from the hand to the air particles in the syringe causing them to warm. Warmer gas particles move faster, so they hit the walls inside the syringe more often and with more force. This produces more pressure.

3. What variables affect gas pressure?

Volume, temperature, and the amount of gas (number of moles) all affect gas pressure.

4. What is temperature?

Temperature is a measurement of the average motion (kinetic energy) of particles in a system.

5. What is the relationship between pressure and temperature? Explain your hypothesis.

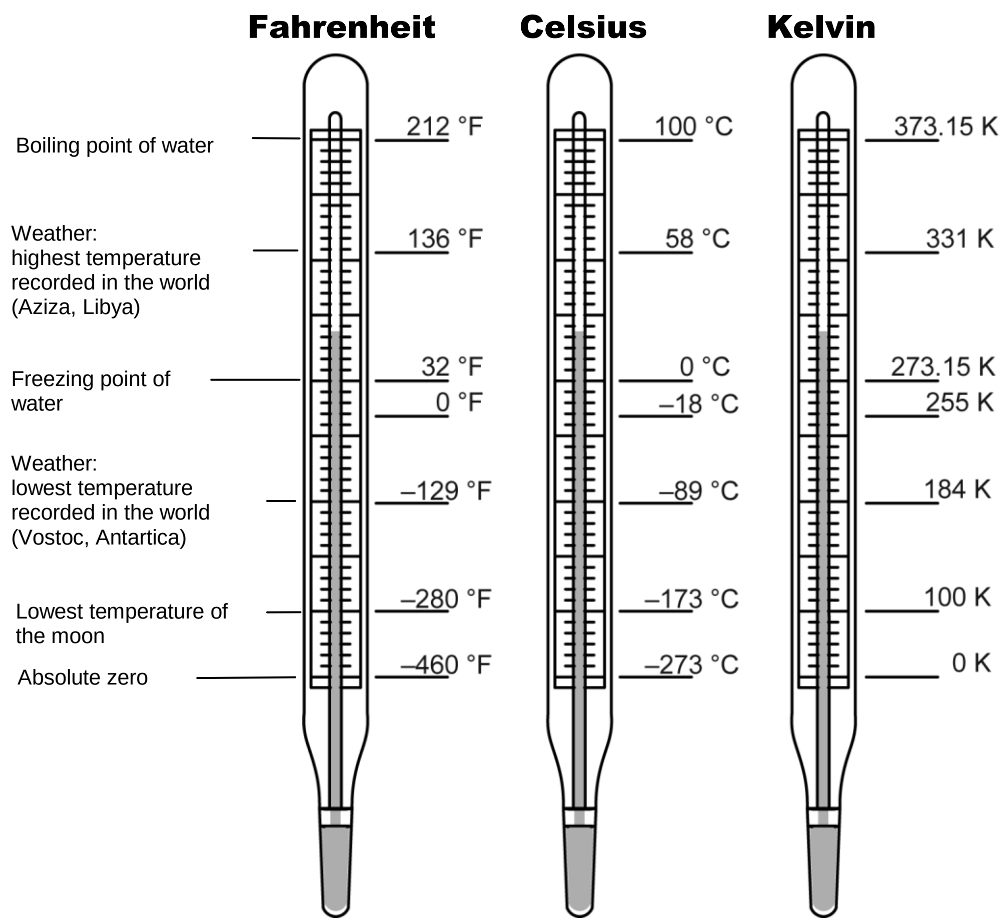
There is a directly proportional relationship between pressure and temperature. As the temperature of the molecules increases, the force the molecules exert on their container (pressure) will also increase.

6. Can particles ever stop moving?

Theoretically, particles can stop moving at a temperature called absolute zero.

Kelvin Temperature Scale

Discuss whether or not particles at 0 °C are moving. Use this discussion to review the basis of the Celsius temperature scale. Explain that the Kelvin temperature scale is based on the movement (kinetic energy) of molecules. Zero Kelvin is the temperature at which particles are no longer in motion (have no kinetic energy) and is called absolute zero. The relationship between the temperature and pressure of a gas requires the temperature to be measured using the Kelvin scale. Substituting temperatures with values in Celsius will lead to incorrect results; however, it is possible to convert between the two units using K = °C + 273.15.



7. Are the particles in a substance at 0 °C moving? Explain.

Student predictions may vary, especially with the natural tendency to automatically associate 0 °C with ice. Remind students that many substances, including air, are not solid at 0 °C. The students will be calculating the temperature at which molecules stop moving (absolute zero) in this lab. Absolute zero equals –273.15 °C.

8. What is the Celsius temperature scale based on?

The Celsius temperature scale is based on the freezing (0 °C) and boiling (100 °C) points of water.

9. What is the Kelvin temperature scale based on?

The Kelvin temperature scale is based on the motion of particles (kinetic energy). Zero Kelvin is the point where particles stop moving (have no kinetic energy). This makes all temperature values positive. The size of the Kelvin temperature scale divisions are the same as on the Celsius temperature scale.

10. Does it matter which temperature scale is used when using the gas laws?

Students address this during the lab analysis questions. Temperatures must have the units Kelvin for gas law calculations. Both Celsius and Fahrenheit temperature scales will give incorrect results if used with the gas laws.

Extracting Meaning from Graphs

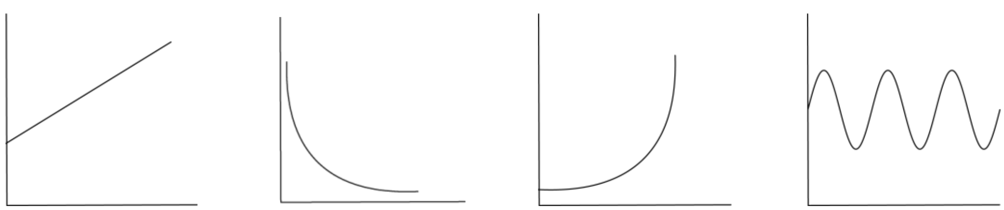
Engage students in a discussion about why we graph data and how to show the different relationships that exist between variables. Focus the discussion on linear graphs and how mathematical concepts they are familiar with, such as the slope, axis intercepts, and the equation of a line, are all used to describe the relationships between variables in science. Start by using a topic that is familiar to the students such as distance, time, and speed. Then go over the same concepts using a pressure versus temperature graph similar to the one the students will be creating in the lab.

11. Why do we graph our data?

Graphs provide a visual display of data. Data visually displayed can show relationships between variables that may not be obvious when the data is listed in data tables.

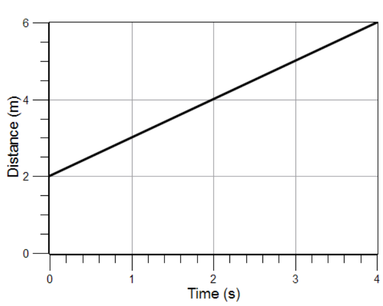
12. What types of relationships can graphs display?

The most common relationships graphs display are linear relationships ($y = mx + b$), some of which are directly proportional ($y = mx$). Other relationships include inverse proportions $\left(y = m\frac{1}{x} + b\right)$, exponential $\left(y = x^{\alpha}, y = \log x\right)$, and periodic $\left(y = \sin x, y = \cos x, \text{and others}\right)$.



Linear Inverse Exponential Periodic

13. What information can be extracted from this distance versus time graph?



The graph shows the position (distance) of an object at different points in time. This is determined by looking at the labels and units of the axes.

- The slope of the line gives the speed the object is traveling in meters per second. Slope is calculated using the formula

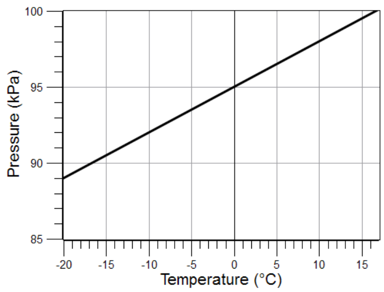
$$\text{slope} = \frac{y_2 - y_1}{x_2 - x_1}$$

$$\frac{6\text{ m} - 2\text{ m}}{4\text{ s} - 0\text{ s}} = \frac{4\text{ m}}{4\text{ s}} = 1\text{ m/s}$$

- The y-intercept (where the line "intercepts", or crosses, the y-axis) is the value of y when the value of x is zero. In this graph, the y-intercept is 2 m, which means the object was 2 meters away when the time was zero.

- The equation of the line is y = 1x + 2. This comes from the equation for a line, $y = mx + b$, where *m* is the slope and *b* is the y-intercept.

14. What information can be extracted from this pressure versus temperature graph?



- The graph shows the pressure of a substance at different temperatures. This is determined by looking at the labels and units of the axes.

- The slope of the line gives the change in pressure for each unit of temperature, kPa/°C.

$$\text{slope} = \frac{y_2 - y_1}{x_2 - x_1}$$

$$\frac{100\text{ kPa} - 95\text{ kPa}}{15\ ^\circ\text{C} - 0\ ^\circ\text{C}} = \frac{5\text{ kPa}}{15\ ^\circ\text{C}} = 0.33\text{ KPa/°C}$$

- The y-intercept is the value of y when the value of x is zero. In this graph, the y-intercept is 95 kPa, which means the substance has a pressure of 95 kPa at 0 °C.

- The equation of the line is $y = \left(\frac{0.33\text{ kPa}}{^\circ\text{C}}\right)x + 95\text{ kPa}$.

  This comes from the equation for a line $y = mx + b$, where *m* is the slope and *b* is the y-intercept.

15. **How can absolute zero be determined from the information provided in the graph in the preceding question?**

Absolute zero is the temperature at which there is no motion of particles. If there is no particle motion, then there can be no pressure. Therefore, pressure must equal zero (pressure is on the y-axis, so y = 0). The point where y = 0 is also called the x-intercept (where the line “intercepts”, or crosses, the x-axis). This can be found by using the linear equation and substituting *y* = 0.

$y = \left(\dfrac{0.33\ \text{kPa}}{°\text{C}}\right)x + 95\ \text{kPa}$, where *y* = pressure (*P*) and *x* = temperature (*T*)

$$0\ \text{kPa} = \left(\dfrac{0.33\ \text{kPa}}{^\circ\text{C}}\right)T + 95\ \text{kPa}$$

$$-95\ \text{kPa} = \left(\dfrac{0.33\ \text{kPa}}{^\circ\text{C}}\right)T$$

$$T = \dfrac{-95\ \text{kPa}}{\left(\dfrac{0.33\ \text{kPa}}{^\circ\text{C}}\right)}$$

$$T = -290\ ^\circ\text{C}$$

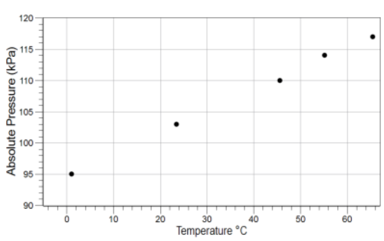
Note: This example uses idealized sample data. Working carefully, the students should be able to determine a more accurate value for absolute zero that is closer to the accepted value of –273.15 °C.

Best Fit Lines versus Connecting Lines

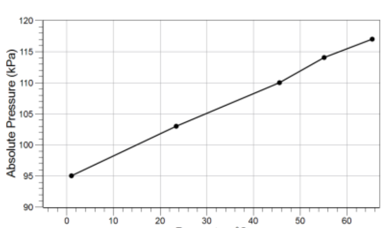
In this lab, each lab group will collect pressure values at five different temperatures. Discuss with students the best way to create a line from the five data points they collect. Differentiate the terms scatter plot, connected line, and best fit line. Emphasize that the goal of a graph is to illustrate the relationship between the variables.

16. **What is the difference between a scatter plot, a graph with a connected line, and a graph with a line of best fit?**

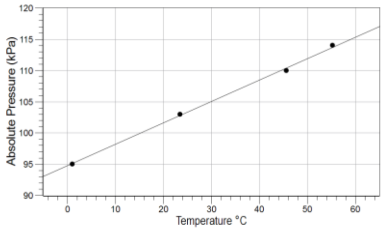
A scatter plot graphs each data point. No lines are drawn.



A connected line plot graphs each data point and then connects each point to the one that is next to it using a straight line but does not produce an “overall” trend for the data as a whole.



A best fit plot graphs each data point and then draws a line that passes as close to as many data points as possible. It does not go through each data point because there may be a small amount of random error associated with each data point. This graph attempts to minimize this error.



17. Which type of plot will best illustrate the relationship between pressure and temperature? Why?

A best fit line using linear regression (a straight line) will be the most appropriate. Using a scatter plot will make it difficult to see a trend, and a connected line plot does not take into account the error associated with each point and does not produce a mathematical relationship over the entire range of data.

18. How does the data collection system automatically graph the data points? Why do you think it does this? Should it be changed?

The data collection system automatically connects the data points. This is done to make the data points easier to see. The connected line should be turned off in this lab and the data points should be fit with a line of best fit using linear regression (a straight line).

Lab Preparation

These are the materials and equipment to set up prior to the lab. There are several ways that a variety of water temperatures (~45 °C, ~55 °C, ~65 °C) can be prepared and made available to the students for this experiment:

- Maintain three different hot water baths in different locations in the classroom. Monitor the temperature of each hot water bath with a temperature sensor and a data collection system. The students can obtain their water from these hot water baths.
- Have the students heat water on individual hot plates at their lab stations. If this is done at the beginning of the lab period the water should be hot by the time the students need the hot water samples. The students can use a second temperature sensor (stainless steel) to monitor the water temperature before submerging their test tube. Room temperature water can be mixed with the hot water to cool it as needed.
- Use an electric kettle or microwave oven to boil water and have the students mix the hot water with room temperature water to produce the desired temperature. Have temperature monitoring stations set up next to the electric kettle or microwave oven so the students can check the temperature of their water before taking it to their lab stations.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Use care when working with hot water.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	5	3	4
Place the closed test tube into a polystyrene cup containing ice water and record the pressure and temperature.	Set up the equipment to measure the pressure and temperature of air trapped in a test tube.	Use the equation of the best fit line to determine absolute zero (temperature at which pressure equals zero).	Replace the ice-water with water at 25 °C, 45 °C, 55 °C, and 65 °C. Record the pressure and temperature for each.	Graph Pressure versus Temperature data and create a straight line of best fit.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect the fast response temperature sensor to the to the data collection system. �(2.1)

3. ☐ Connect the absolute pressure sensor to the data collection system using a sensor extension cable. �(2.2)

4. ☐ Display Pressure (kPa) and Temperature (°C) in a table. �(7.2.1)

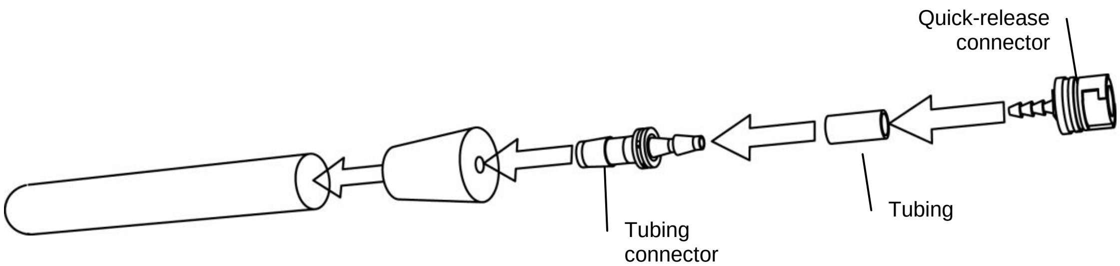
5. ☐ Put the data collection system into manual sampling mode without manually entered data. �(5.2.2)

6. ☐ Connect the quick-release connector to the rubber stopper using the tubing connector and the 1- to 2-cm piece of tubing by following the steps below. Use the picture as a guide.

a. Insert the thicker end of the tubing connector into the hole in the rubber stopper. If this is difficult, add a drop of glycerin.

b. Connect the 1- to 2-cm piece of tubing to the other, thinner end of the tubing connector.

c. Insert the barbed end of the quick-release connector into the open end of 1-to 2-cm piece of tubing. If this is difficult, add a drop of glycerin.

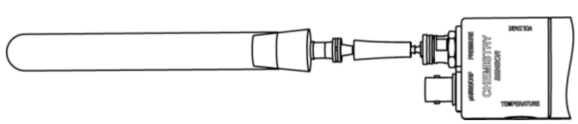


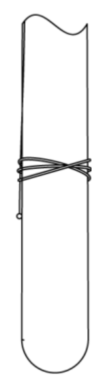
7. ☐ Why is glycerin added when it is difficult to connect pieces of plastic together?

Glycerin makes connecting the pieces much easier because it lubricates the ends.

8. ☐ Fit the rubber stopper tightly into the test tube.

9. ☐ Insert the quick-release connector to the port of the absolute pressure sensor and then turn the connector clockwise until the fitting "clicks" onto the sensor (about one-eighth turn).





10. ☐ Use the rubber band to attach the quick-response temperature sensor to the outside of the test tube. The sensor should be about halfway down the test tube.

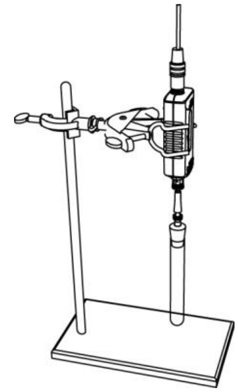
11. ☐ The temperature and pressure of what substance are being measured during this experiment?

The temperature and pressure of the air inside the test tube are being measured.

12. ☐ The temperature sensor is placed on the outside of the test tube. Is this a problem?

It would be better if the temperature sensor could be placed inside the test tube since it is the temperature of the air inside the test tube that is of interest. This is not much of a problem, however, because the heat quickly transfers from the water bath to the air inside the test tube until both are the same temperature.

13. ☐ Attach a three-finger clamp to a ring stand. Use the three-finger clamp to securely hold the absolute pressure sensor in a vertical position.



14. ☐ List the dependent and independent variables used in this experiment. Also give their units of measurement.

The independent variable is temperature measured in units of degree Celsius (°C). The dependent variable is pressure measured in units of kilopascals (kPa).

15. ☐ What do you expect to happen to the pressure as the temperature increases?

The pressure should increase as the temperature increases.

16. ☐ Place a polystyrene cup into a 250-mL beaker.

17. ☐ Fill the polystyrene cup to the top with ice and add water to make an ice bath.

18. ☐ Why is a polystyrene cup used instead of just a beaker?

Polystyrene is an insulator and will minimize heat transfer to or from the water. This is necessary in order to maintain a constant temperature.

19. ☐ Place the beaker containing the polystyrene cup and ice water underneath the absolute pressure/test tube apparatus and slowly lower the test tube into the ice water.

20. ☐ Angle the test tube in the cup so the entire test tube is covered with the ice water. Place more ice on top of the test tube to ensure that the entire test tube is covered with ice.

Collect Data

21. ☐ Ensure that your data collection system is in manual sampling mode and you are viewing the table display. Start a manually sampled data set. �(6.3.1)

22. ☐ Wait about 2 minutes to allow the temperature of the air in the test tube to become the same temperature as the ice water surrounding it and then record a data point. �(6.3.2)

23. ☐ Remove the test tube from the ice water bath.

24. ☐ Place the second polystyrene cup into the second 250-mL beaker (the polystyrene cup and beaker should both be at room temperature).

25. ☐ Fill the polystyrene cup with water at room temperature (~25 °C).

26. ☐ Angle the test tube in the cup so the entire test tube is covered with water.

27. ☐ Wait about 2 minutes for the temperature and pressure to stabilize and then record the data point. �(6.3.2)

28. ☐ Collect three more data points by replacing the room temperature water with water at: 45 °C, 55 °C, and 65 °C. For each sample, wait about 2 minutes for the readings to stabilize and then record the data points. �(6.3.2)

29. ☐ What is the role of the water inside the polystyrene cup? Is it necessary to know the exact amount of water added? Explain your reasoning.

The water added to the polystyrene cup is being used to change the temperature of the air inside the test tube. The exact amount added is not important, but it is important to have the test tube completely covered so the air inside the test tube is the same temperature as the water.

30. ☐ When all five data points have been recorded, stop the data set. �(6.3.3)

31. ☐ Copy the pressure and temperature data from your data collection system to Table 1 in the Data Analysis section.

32. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

Table 1: Pressure and temperature

Sample	Pressure (kPa)	Temperature (°C)	Temperature (K)
Ice water	95	1.0	274.2
Room temperature	103	23.5	296.6
~45 °C	110	45.6	318.8
~55 °C	114	55.2	328.4
~65 °C	117	65.5	338.6

1. ☐ On your data collection system, change the units of temperature to Kelvin (K). �(5.3)

Note: You will have to switch back and forth between units several times while analyzing the data.

2. ☐ Copy the Kelvin temperatures from your data collection system to Table 1 above.

3. ☐ Create a graph of Pressure (kPa) versus Temperature (°C) on your data collection system. �(7.1.1)

Note: To graph a scatter plot of the data points hide the connecting lines between data points feature. �(7.1.8) Adjust the scale of the graph to show all data as needed. �(7.1.2)

4. ☐ Create a linear line of best fit through the data. �(9.5)

5. ☐ Determine the slope, y-intercept, and equation of the best fit line. Include units with each number. �(9.6)

slope = 0.34 kPa/°C

y-intercept = 95 kPa (determined from the graph where temperature = 0 °C)

equation of the line: $y = mx + b$, where y = pressure, m = slope, x = temperature, and b = y-intercept

$$y = \left(\frac{0.34\ \text{kPa}}{°\text{C}}\right)x + 95\ \text{kPa}$$

6. ☐ Use the equation of the best-fit line to solve for absolute zero. Absolute zero is the temperature when the pressure is zero. Show your work.

$$y = \left(\frac{0.34\ \text{kPa}}{°\text{C}}\right)x + 95\ \text{kPa}$$

$$0 = \left(\frac{0.34\ \text{kPa}}{°\text{C}}\right)x + 95\ \text{kPa}$$

$$\frac{-95\ \text{kPa}}{\left(\frac{0.34\ \text{kPa}}{°\text{C}}\right)} = x$$

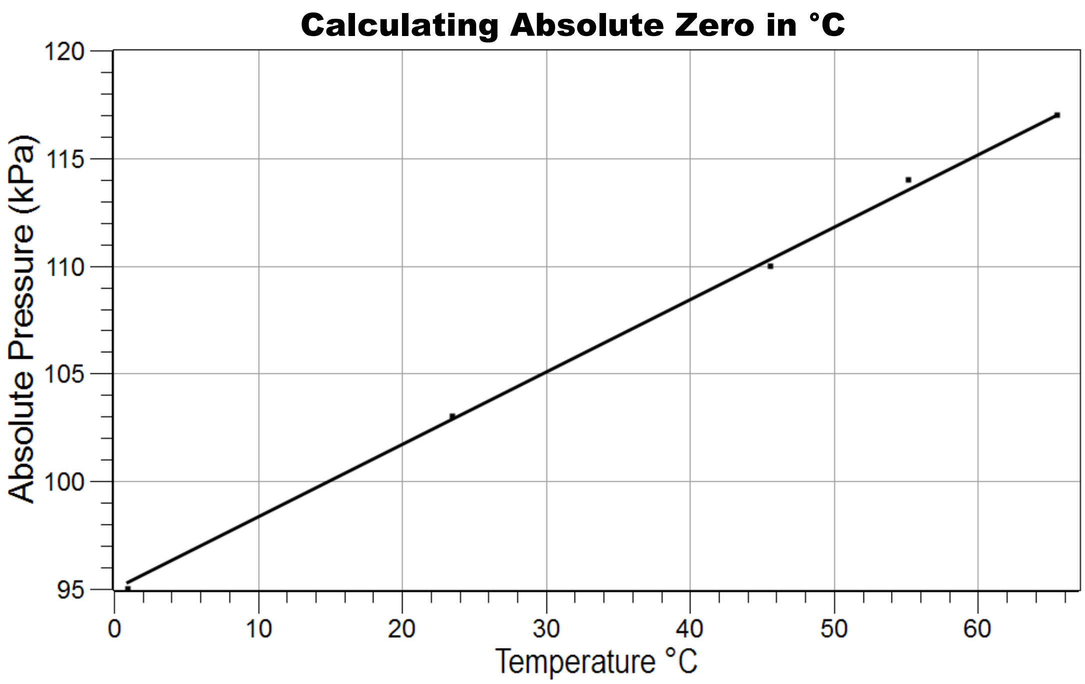
$$-280\ °\text{C} = x$$

7. ☐ Calculate the percent error of your experimental value of absolute zero (the accepted value of absolute zero is –273.15 ˚C). Show your work.

$$\text{Percent error} = \left|\frac{\text{accepted value} - \text{experimental value}}{\text{accepted value}}\right| \times 100$$

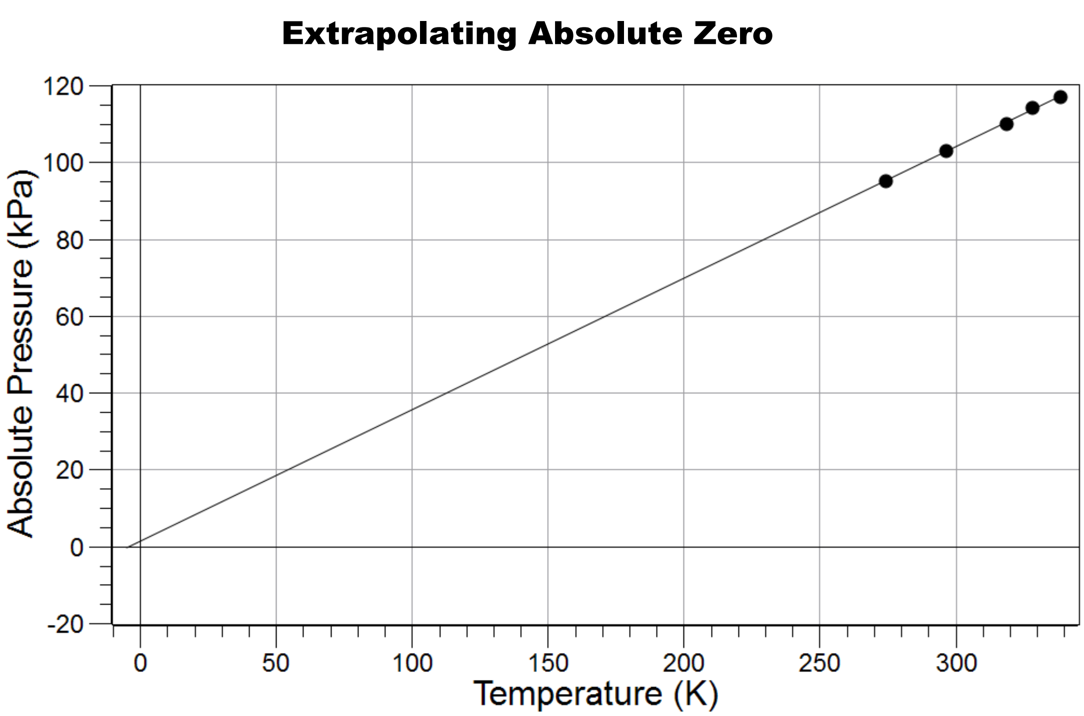
$$\text{Percent error} = \left|\frac{-273.15 - 280}{-273.15}\right| \times 100 = 2.5\%$$

8. ☐ Plot or print a copy of the graph of Absolute Pressure (kPa) versus Temperature (°C). Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



9. ☐ Change the units of temperature to Kelvin (K) on the graph displayed on your data collection system and create a linear line of best fit through the data. �(5.3)(9.5)

10. ☐ Plot or print a copy of the graph of Absolute Pressure (kPa) and Temperature (K). Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



11. ☐ Complete Table 2 below by determining the constant in Gay-Lussac's law for three data points using temperature values in °C and the same temperature values in Kelvin. From these calculations explain whether °C, K, or both °C and K support Gay-Lussac's law.

$$\frac{P}{T} = k$$

Table 2: Calculations of the constant in Gay-Lussac's law

Pressure (kPa)	**Celsius Temperatures**		**Kelvin Temperatures**	
	***T* (°C)**	**Calculation for *k***	***T* (K)**	**Calculation for *k***
95	1.0	$\frac{95\text{ kPa}}{1.0\text{ °C}} = 95\text{ kPa/°C}$	274.2	$\frac{95\text{ kPa}}{274.2\text{ K}} = 0.35\text{ kPa/K}$
110	45.6	$\frac{110\text{ kPa}}{45.6\text{ °C}} = 2.4\text{ kPa/°C}$	318.8	$\frac{110\text{ kPa}}{318.8\text{ K}} = 0.35\text{ kPa/K}$
117	65.5	$\frac{117\text{ kPa}}{65.5\text{ °C}} = 1.79\text{ kPa/°C}$	338.6	$\frac{117\text{ kPa}}{338.6\text{ K}} = 0.346\text{ kPa/K}$

Gay-Lussac's law is only supported using Kelvin temperature units. Kelvin units were the only units that resulted in *k* being constant.

Analysis Questions

1. Are pressure and temperature directly or indirectly proportional? How do you know?

Pressure and temperature are directly proportional. The data fits a straight line better than a curve. As the temperature increases, the pressure increases at a constant rate.

2. Explain how you can convert between Celsius and Kelvin temperature scales.

Kelvin temperature is equal to the temperature in degrees Celsius plus 273.15.

K = °C + 273.15

3. What is the Kelvin temperature scale based on? What is special about 0 K?

The Kelvin temperature scale is a based on the motion of particles (kinetic energy). Zero Kelvin is the temperature at which particles stop moving (have no kinetic energy) and is referred to as "absolute zero."

4. Calculate the pressure of the air inside the test tube if it were warmed to 100.0 °C.

Gay-Lussac's law $\frac{P}{T} = k$

Rearrange Gay-Lussac's law to solve for pressure $P = kT$

The value for the constant, *k*, can be taken from the Data Analysis section above (calculation for *k* using Kevin temperatures).

$k = 0.35\ \text{kPa/K}$

Convert the temperature to Kelvin units

$\text{K} = °\text{C} + 273.15$
$\text{K} = 100.0\ °\text{C} + 273.15$
$\text{K} = 373.2$

Solve for pressure $\text{P} = \left(\frac{0.35\ \text{kPa}}{\text{K}}\right)(373.2\ \text{K})$

$\text{P} = 130\ \text{kPa}$

Synthesis Questions

Use available resources to help you answer the following questions.

1. Another lab group did the same experiment using a syringe instead of a test tube. Could this cause a problem? Explain.

The volume of a syringe could change. Pressure is affected by both temperature and volume. This would mean the recorded pressures were affected by a change in volume and not just temperature alone. This would lead to incorrect results for Gay-Lussac's law.

2. Explain why an over-inflated tire may pop when it is driven fast on a hot day.

Both the outside temperature and the friction of the road will increase the temperature of the air inside the tire. This increase in temperature will cause the air particles inside the tire to move faster and exert more pressure on the walls of the tire. If the pressure exerted by the air inside the tire exceeds the strength of the tire, the tire will pop.

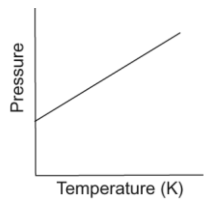
3. Would a gas cooled nearly to absolute zero remain a gas or would it have changed to a liquid or a solid?

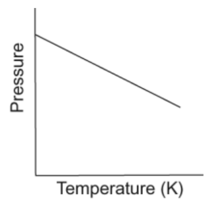
It would have changed into a solid, because the particles are moving so slowly they would be held together by weak forces of attraction.

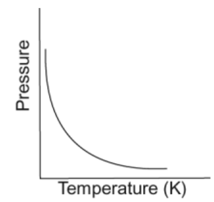
Multiple Choice Questions

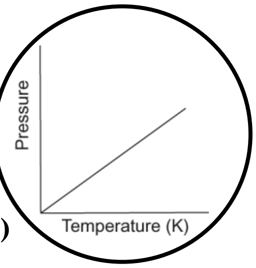
Select the best answer or completion to each of the questions or incomplete statements below.

1. Which of the following graphs best represents the relationship between the pressure of a gas and Kelvin temperature?

A) 

B) 

C) 

D) 

2. The reason that the pressure becomes zero at a temperature of absolute zero is:

- **A. At absolute zero, all molecular motion stops**
- **B.** At absolute zero, there is a complete vacuum
- **C.** At absolute zero, the volume of the gas is very small
- **D.** At absolute zero, all the energy of the gas is given off as light

3. What is the equivalent of 423 K in degrees Celsius?

- **A.** 273 ˚C
- **B.** 0 ˚C
- **C. 150 ˚C**
- **D.** 696 ˚C

4. If a container of gas is at a temperature of 27 °C and a pressure of 107 kPa, what would the pressure of the gas become if the temperature were doubled to 54 °C?

- **A.** 214 kPa
- **B.** 53.5 kPa
- **C. 117 kPa**
- **D.** 97 kPa

5. If the temperature of a gas in a closed container decreases, the pressure inside the container will

- **A.** Increase
- **B. Decrease**
- **C.** Stay the same
- **D.** Either A or B depending on the type of gas

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The **kinetic** energy of gas particles is related to the speed of the gas particles. **Temperature** is the average energy of motion of these particles. **Pressure** in a gas is due to the collisions of particles with the walls of the container the gas is contained in. If the temperature increases, the gas particles move faster, so they hit the wall more often. This **increases** the gas pressure. The temperature at which all particle motion stops is known as **absolute zero**. This temperature is **–273.15 °C**. The **Kelvin** temperature scale places zero as the temperature at which all motion stops and then numbers forward creating only positive temperature values.

Extended Inquiry Suggestions

Repeat the experiment using three different volumes of air. Determine the effects of changing the volume of air (by changing the size of the test tube) on the slope of the best fit line as well as the value of absolute zero calculated.

Repeat the experiment using a different gas in the test tube. Compare the results with air.

Determine the effect of different units of pressure on Gay-Lussac’s Law.

10. Phase Change

Objectives

Determine how to add heat to a substance without the temperature of the substance increasing. Through this investigation, students:

- Determine the effect of a phase change on the temperature of a substance
- Explain the difference between heat and temperature
- Determine the melting point and boiling point of pure water

Procedural Overview

Students conduct the following procedures:

- Collect temperature data as they freeze water by inserting it into a salt/ice bath until the water freezes and the ices temperature decreases to –6.0 °C
- Collect temperature data as they add a constant amount of heat to ice until the ice melts and the temperature of the water rises to 8 °C
- Collect temperature data as they add a constant amount of heat to water until it boils for 6 to 8 minutes

Time requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	20 minutes
♦ Lab activity	90 minutes (45 minutes for each part)[1]

[1] Refer to the Lab Preparation section for tips on ways to make this lab fit into one 45 minute lab period.

Materials and Equipment

For each student or group:

- Data collection system
- Stainless steel temperature sensor[1]
- Hot plate
- Beaker (2),150-mL or larger
- Graduated cylinder, 10-mL
- Test tube, 10-mm x 100-mm
- Test tube rack
- Ring stand
- Utility clamp
- Stir rod
- Tablespoon
- Distilled (deionized) water, 103 mL
- Crushed ice to fill the beaker
- Rock salt, 200 g

[1]A fast response temperature sensor is not appropriate for Part 2 of this investigation.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- States of matter
- Kinetic molecular theory
- Energy

Related Labs in this Manual

Labs conceptually related to this one include:

- Heat of Fusion
- Intermolecular Forces

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to your data collection system �(2.1)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Naming a data run �(8.2)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

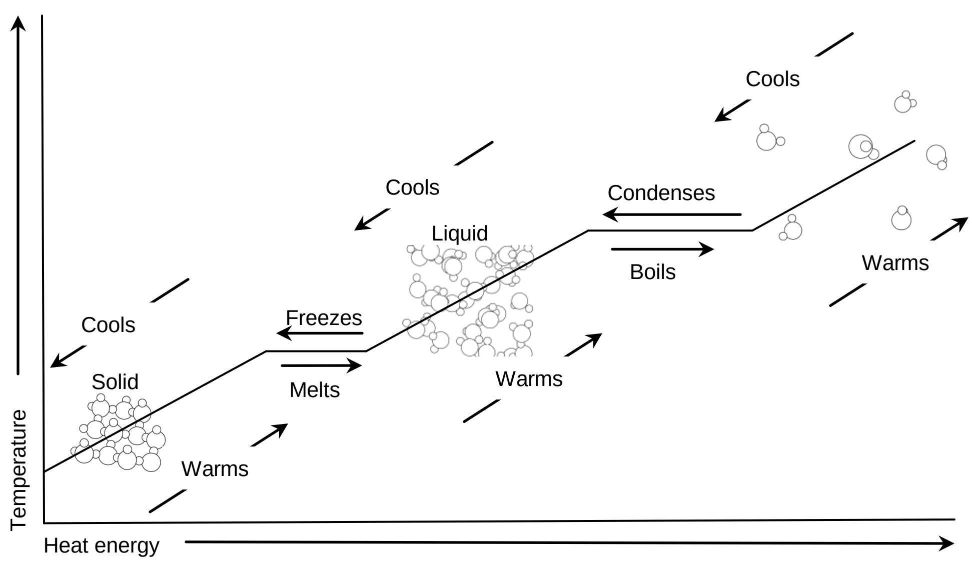
The kinetic molecular theory explains that all matter consists of particles in constant motion. Solid substances contain vibrating particles closely locked into position by electrostatic forces of attraction. Particles that make up a liquid are very close to each other, but with weaker electrostatic attractions between them than those in a solid, enabling the particles to move past one another. As a gas, the particles are far apart from each other, do not exhibit any attraction for each other, and move freely.

All pure substances can exist as a solid, liquid, or a gas, depending on temperature and pressure. (We will assume constant pressure throughout this experiment.) In order for a substance to change phases, energy must be added or removed. During warming, heat energy is absorbed and either makes the molecules move faster or breaks the attractions between particles. During cooling, heat energy is removed by either causing the molecules to move slower or by forming attractions amongst neighboring particles.

Adding heat to a substance will generally cause the temperature of the substance to increase, but not always. Heating a substance in a specific phase causes the temperature of the substance to increase. However, heating a substance as it undergoes a phase change does not result in a temperature change. Understanding the difference between heat and temperature explains this. Heat is generally defined as the flow or transfer of energy due to a difference in temperature. More specifically, heat is a measure of the total change in internal energy of a substance. Total internal energy refers to both the potential energy (attractions between molecules) and the kinetic energy (motion of the molecules). Temperature is a measurement related to the average kinetic energy of molecules only. During a phase change, the added heat only breaks the attractions between the molecules (an increase in potential energy). There is no change in the kinetic energy of the particles and therefore no change in temperature.

For a pure substance, the temperature at which a phase change occurs can identify the substance. The four main phase changes are melting point (solid to liquid), freezing point (liquid to solid), boiling point (liquid to gas), and condensing point (gas to liquid).

Phase Changes



Pre-Lab Discussion and Activity

Boiling Different Volumes of Water

Hold up two small beakers of water. Beaker A should be a quarter of the way filled with distilled water, and beaker B should be filled halfway with distilled water. Ask your students to predict which beaker of water will boil at a higher temperature and have them explain their reasoning. Place the beakers on an already heated hot plate and project the collected data. During data collection, discuss what is happening at the molecular level using the magnetic molecules activity below. Remember to check the beakers of boiling water. You may have to remove beaker A if all the water evaporates.

1. Will the water in beaker A boil at a lower temperature, higher temperature, or the same temperature as the water in beaker B?

Water boils at the same temperature regardless of the amount of water being heated.

2. What are some similarities and differences between what occurs in beaker A and beaker B?

Similarities: Both contain water, receive the same amount of heat from the hot plate, and are heated for the same amount of time.

Differences: The amount of water being heated, and the rate at which the water temperature increases (refer to the live data collection).

Magnetic Molecules

Use magnets to represent the water molecules. Explain to the students that attractive forces between water molecules hold them together similar to the way magnets attract each other. The water molecules in the beaker obtain energy from the heat produced by the hot plate. Similarly, muscle power can transfer energy to the magnets.

With the magnets stuck to each other, shake the group gently to represent energy transfer to the particles motion. The same thing happens to the water molecules in the beaker. As the heat transfers, they are "shaking" faster. Temperature measures the change in molecular motion.

Write the definition of temperature on the board: Temperature is related to the average kinetic energy (motion) of the particles.

Next, pull apart the magnets to represent energy being used to cause a phase change. Guide the students toward understanding that energy is required to pull the particles apart (break the attractive forces). The energy used to pull the particles apart is called potential energy. Write the definition of heat on the board: Heat is the change in the total internal energy which includes both the kinetic energy (motion) *and* the potential energy (breaking attractions) of a substance. Tell the students that unlike temperature, heat cannot be directly measured.

Finally, shake the separated magnets more vigorously to represent another increase in temperature. Have the students help you graph a "shaking" versus "Time" curve to model the expected data the students will collect in this lab.

3. Does energy make the molecules/magnets shake faster? What type of energy does this represent?

Yes. Energy from the teacher's arm transfers to the magnets causing them to move (shake). Similarly, energy from the hot plate transfers to the water molecules causing them to move faster. This energy of motion is called kinetic energy.

4. What is the relationship between particle motion and temperature?

There is a direct relationship between particle motion and temperature. As the temperature of the particles increase, the molecules move faster.

5. What needs to happen to water molecules in the liquid state for them to turn into water vapor? Why does this require energy?

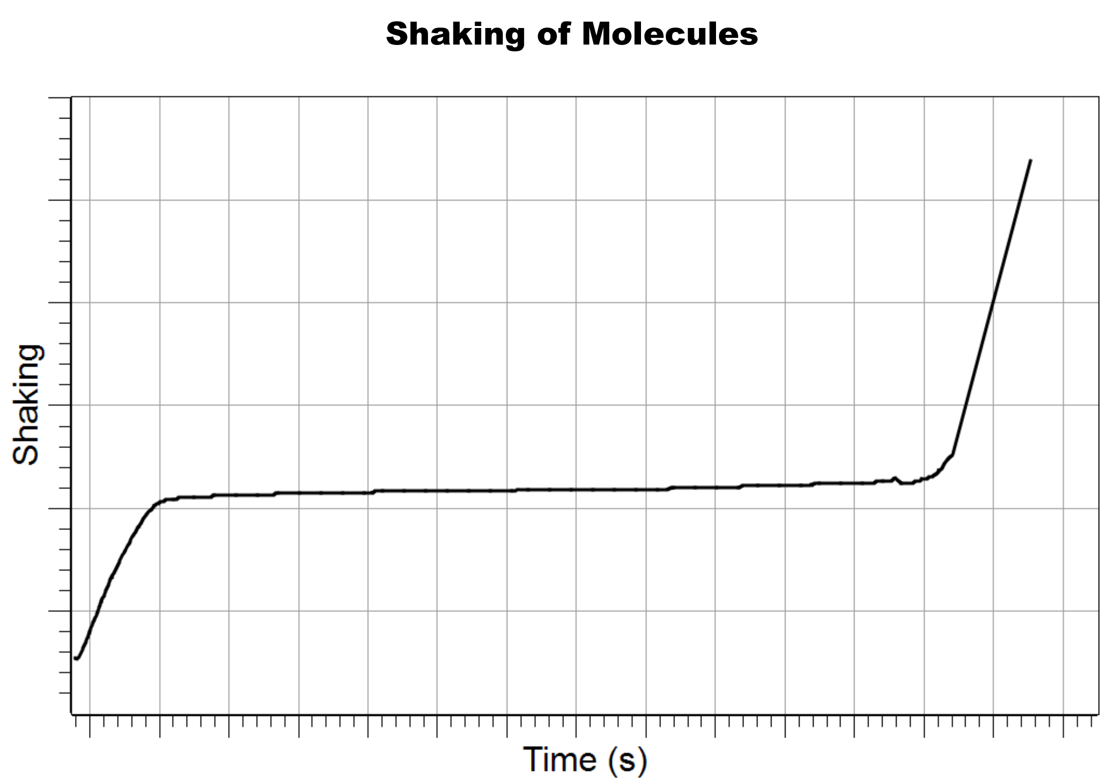
Water molecules in the liquid state must be pulled apart (separated) from each other to turn into water vapor. This requires energy to break apart the attractions that hold the particles together.

6. What are the two types of energy that make up heat? How are they related to individual molecules/magnets?

Heat involves both kinetic energy and potential energy. Kinetic energy makes the molecules move faster and potential energy separates the molecules apart.

7. What would a graph of "shaking" versus "time" look like? Why?

The energy from the heat will be used to make the molecules move faster. When the particles are shaking with enough energy to overcome the attractive forces holding them together, the added energy will be used to pull the particles apart instead of making them "shake" faster. This creates the horizontal line. Once all the particles are separated, the added energy will be used to increase the "shaking" of the individual molecules again.



Boiling Different Volumes of Water – Conclusions

Use the data collected in "Boiling Different Volumes of Water" above to engage the students in a discussion about what the data collected means. Guide the students to understand that no matter how much or little water we have, at a given pressure, pure water always boils at the same temperature. The temperature at which a substance boils is called its boiling point.

8. What happened in our boiling different volumes of water demonstration? Did beaker A boil at a lower temperature, higher temperature, or the same temperature as the water in beaker B? Explain the results.

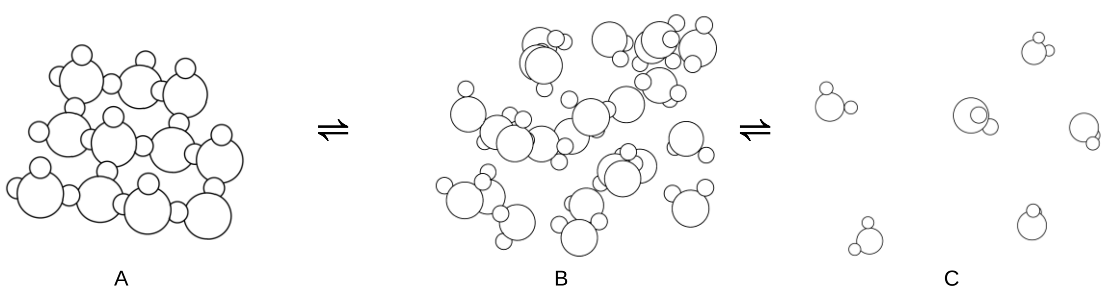
They boiled at the same temperature, but the rate of heating, indicated by the slope of the line, was greater in beaker A than in beaker B. Therefore, the water in beaker B took longer to reach the boiling point. The difference in the quantity of water molecules explains this. Beaker B contained more water, which means there are more water molecules that need to speed up and move apart. It therefore takes more heat, and thus more time, for the water to reach its boiling point.

9. Describe the shape of the graph during boiling? Why does this happen?

During boiling, the graph formed a horizontal line showing no temperature change because the added heat energy only breaks the attractions between particles, which only changes the potential energy. Temperature only measures changes in kinetic energy.

Phases changes at the Molecular Level

Show the students a video or picture of water molecules in its three phases.



10. Name the different phase changes and whether you need to add or remove heat to make these changes take place.

Solid ice to liquid water is melting and requires adding heat. Liquid water to solid ice is freezing and requires removing heat. Liquid water to gaseous water vapor is boiling and requires adding heat. Gaseous water vapor to liquid water is condensing and requires removing heat.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Teacher Tip: To save time, there are several ways you can have the students perform this lab:

- Do either Part 1 or Part 2 as a demonstration (such as in the Pre-Lab Discussion and Activity section), and have the students perform the other part.
- Have students start on Part 2 as they are waiting for their ice to melt. This will require two temperature sensors per group.

- ♦ Have some groups do Part 1 and other groups do Part 2, and then share data.
- ♦ Have the students prepare Parts 1 and 2 at the same time, and then collect the data simultaneously. This will require two temperature sensors per group.
- ♦ Freeze the temperature sensors in test tubes filled with 3 mL of distilled water the night before. This allows the students to skip the Set Up section of Part 1.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Do not touch the hot plate or hot glassware.
- ♦ Allow all glassware and equipment to cool thoroughly before handling.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

Part 1 – Freezing Water and Melting Ice

3	4	1	2
Place the test tube (with the temperature sensor frozen inside) in a test tube rack.	Record the temperature data and stir constantly as the ice melts and the temperature of the water increases to 8 ºC.	Measure 3 mL of distilled water, and pour it into a test tube. Place the temperature sensor in the test tube containing the water.	Collect temperature data as you freeze the water in the test tube by placing it in a salt and ice mixture until it is –6.0 ºC or cooler.

Part 2 – Boiling Water

2	3	4	1
Assemble the ring stand, utility clamp, and temperature sensor. Fill the beaker with distilled water.	Place the beaker filled with water on the warmed hot plate, and lower the temperature sensor into the water.	Start recording temperature data. Continue recording data until the water has boiled for 8 to 10 minutes.	Turn on the hot plate, and allow it to warm to its highest setting.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Part 1 – Freezing Water and Melting Ice

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a stainless steel temperature sensor to the data collection system. �(2.1)

3. ☐ Create a graph of Temperature (°C) versus Time (s). �(7.1.1)

4. ☐ Predict what a temperature versus time graph will look like for freezing water.

The temperature will decrease until it reaches 0 °C and then it will level off and remain horizontal until all the water is frozen. After the water has all turned into ice, the temperature will steadily decrease again.

5. ☐ Predict what a temperature versus time graph will look like for melting ice.

The temperature will increase until it reaches 0 °C and then it will level off and remain horizontal until all the ice is gone. After the ice is gone, the temperature will steadily increase.

6. ☐ Use a graduated cylinder to measure 3 mL of distilled water and pour the water into the test tube.

7. ☐ Place the temperature sensor into the test tube.

Collect Data

8. ☐ Collect temperature versus time data as you freeze the water in the test tube using a salt/ice mixture. To do this:

a. Fill a beaker halfway full of ice.

b. Add two spoonfuls of rock salt.

c. Use a stir rod to mix the rock salt and ice together.

d. Place the test tube containing the temperature sensor into the ice and salt mixture.

e. Start recording temperature data. �(6.2)

Note: You may need to adjust the scale the axes to observe the changes taking place. �(7.1.2)

f. Carefully add more ice around the test tube and two more spoonfuls of salt.

g. Gently stir the mixture with the test tube. Keep the temperature sensor positioned so the tip freezes in the ice.

Note: Sometimes water in the test tube will super cool (remain liquid at lower than 0 ºC). If this happens, quickly remove the test tube from the salt and ice mixture, and then place it back into the mixture

h. Stop recording temperature data when the temperature of the ice falls to –6.0 ºC or cooler. �(6.2)

9. ☐ Name the data run "freezing ice". �(8.2)

10. ☐ Start recording another run of data. �(6.2)

11. ☐ Remove the test tube from the salt and ice mixture, and place it in a test tube rack.

12. ☐ Allow the ice to melt. Once the ice has melted enough, twist the temperature sensor so that the ice on the temperature sensor is constantly mixing with the water that has melted.

13. ☐ Continue recording data and constantly stirring until all the ice melts and the water temperature rises to between 8 and 10 ºC.

14. ☐ What are the independent and dependent variables in this experiment?

The independent variable is the time that constant heat is added to the ice.

The dependent variable is the temperature.

15. ☐ Record the temperature at which the water froze and the ice melted.

Ice melts and freezes at (or near) 0 ºC.

16. ☐ From where does the heat causing the ice to melt come?

It comes from the surroundings, in this case, the room temperature air.

17. ☐ Stop recording data when the temperature of the water is between 8 and 10 ºC. �(6.2)

18. ☐ Name the data run "melting ice". �(8.2)

Part 2 – Boiling Water

Set Up

19. ☐ Ensure that the data collection system is on, that a stainless steel temperature sensor is connected, and that a graph of Temperature (°C) versus Time (s) is displayed. �(7.1.1)

20. ☐ Turn on your hot plate to its highest setting, and allow it to warm completely. This generally takes about 5 minutes.

CAUTION: Ensure that papers, wires, fingers, etcetera do not touch the hot plate.

21. ☐ Allowing the hot plate to warm up ensures that a constant amount of heat is added to the beaker of water. If you did not allow the hot plate to warm up, how would the heat released change over time?

The heat released would slowly increase until meeting the maximum temperature and then it would remain constant.

22. ☐ Why should the hot plate be set to its highest setting?

Any temperature setting could be used as long as it remains at the same setting during the entire experiment (constant amount of heat). Using the highest setting saves time.

23. ☐ Attach a utility clamp onto a ring stand, and securely tighten a stainless steel temperature sensor to the utility clamp.

24. ☐ Fill the beaker with about 100 mL of distilled water.

25. ☐ What do you predict a temperature versus time graph will look like for boiling water?

The temperature will increase until 100 °C and then it will level off. The water will remain at 100 °C until all the water evaporates.

Collect Data

26. ☐ Place the beaker with water in it on the warmed hot plate.

27. ☐ Lower the stainless steel temperature sensor into the water, positioning the sensor in the center of the water. Make sure is the sensor does not touch the bottom or the sides of the beaker.

28. ☐ Begin recording data. �(6.2)

Note: You may need to adjust the scale the axes to observe the changes taking place. �(7.1.2)

29. ☐ Continue recording data until the water has vigorously boiled (large bubbles) for 8 to 10 minutes.

30. ☐ Why is it important for the temperature sensor to remain in the center of the water?

This will ensure that we measure the temperature of the water. If the temperature sensor touches the bottom of the beaker, it will be measuring the temperature of the hot plate, not the water.

31. ☐ What are the independent and dependent variables in this experiment?

The independent variable is the time that constant heat is added to the water.

The dependent variable is the temperature of the water.

32. ☐ Record the temperature when you observe that the water is at a vigorous boil.

Water is boiling at (or near) 100 ºC.

33. ☐ Why is it necessary to boil the water for so long?

To determine how the temperature changes while the water changes from a liquid to a gas.

34. ☐ Stop recording data when the water has been at a vigorous boil for 8 to 10 minutes. �(6.2)

35. ☐ Name the data run "boiling water". �(8.2)

36. ☐ Turn off the hot plate and allow the equipment to cool for at least 20 minutes.

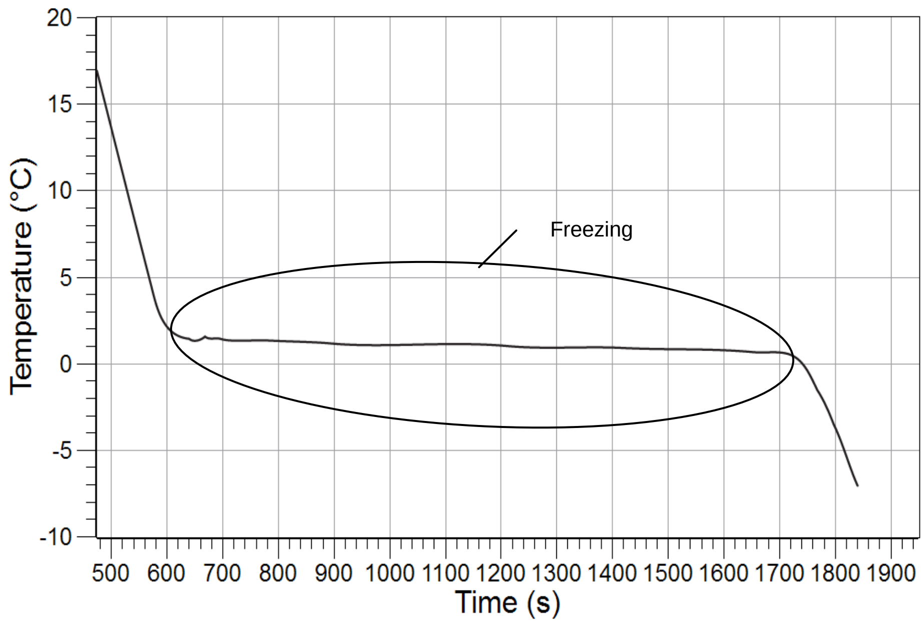
37. ☐ Save your experiment and clean up your lab station according to the teacher's instructions. �(11.1)

Data Analysis

1. ☐ Print or sketch a graph of Temperature (°C) versus Time (s) where heat is removed for freezing water. Label where freezing was occurring. Also label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

Freezing – horizontal line around 0 °C

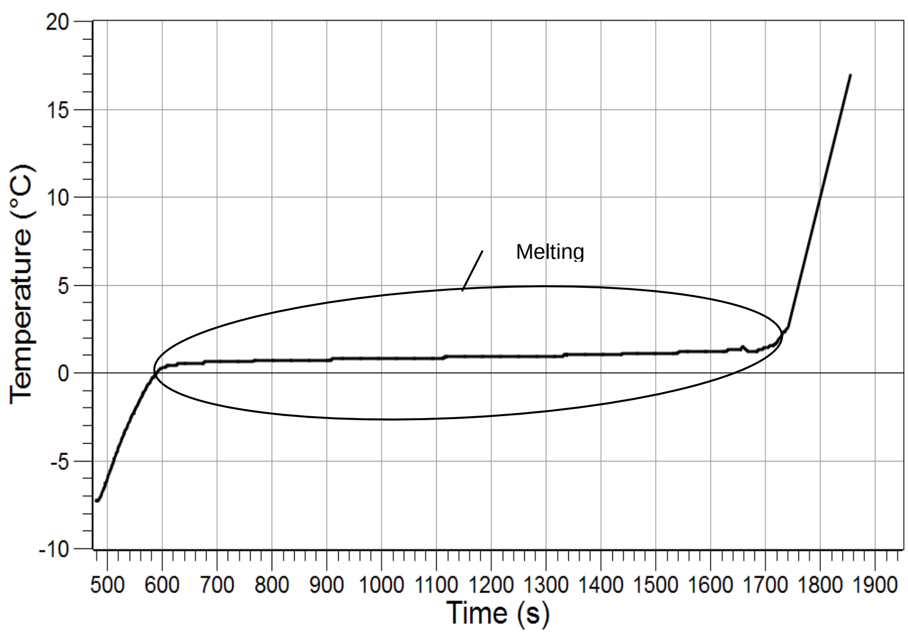
Freezing Water



2. ☐ Print or sketch a graph of Temperature (°C) versus Time (s) where heat is added for melting ice. Label where melting was occurring. Also label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

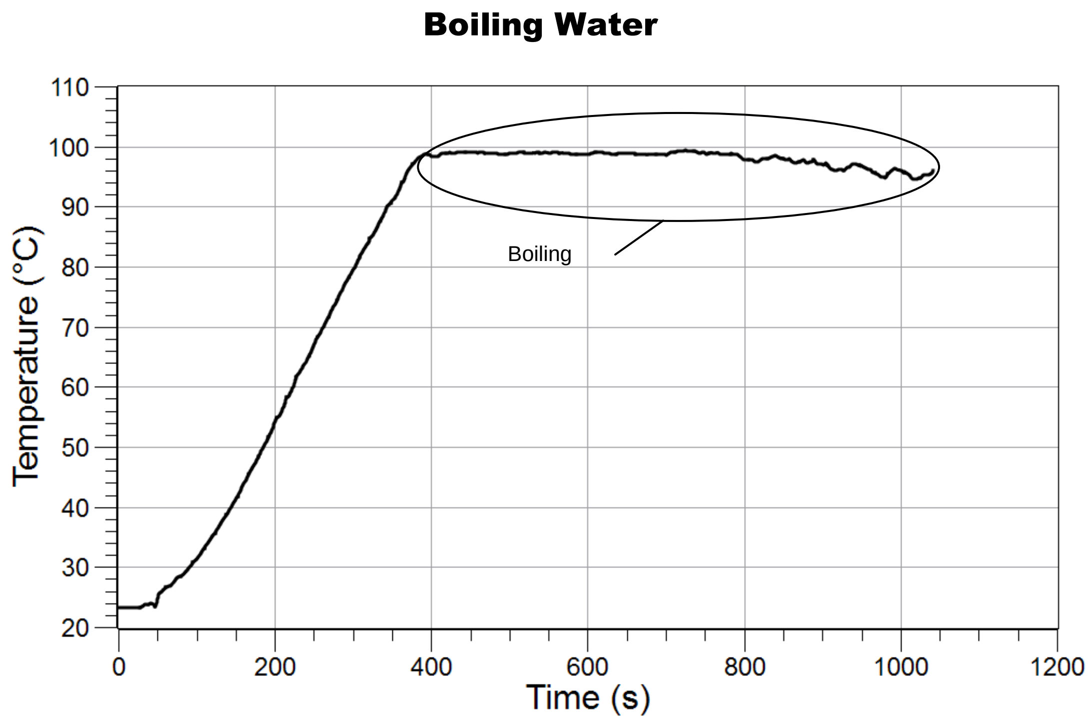
Melting – horizontal line around 0 °C

Melting Ice



3. ☐ Print or sketch a graph of Temperature (°C) versus Time (s) where heat is added for boiling water. Label where boiling was occurring. Also label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

Boiling – horizontal line around 100 °C



Analysis Questions

1. Relate the shape of your graphs above to the behavior of the water molecules. Hint: Explain whether the heat added caused the molecules to move faster or break attractions.

The upward sloping lines on the graphs indicate that the water molecules are moving faster, which causes the temperature to increase.

The horizontal section on the graphs is where the attractions between molecules are breaking. This causes the phase to change (solid to liquid in the first graph, liquid to gas in the second graph). This only changes the potential energy, and therefore does not change the temperature.

2. Explain how it is possible to add heat to a substance without the temperature of the substance increasing.

If heat is added to a substance during a phase change, then the temperature will not increase. This is possible because the added heat is used to break the attractions between particles instead of increasing the average speed of particles.

3. According to your lab results, what is the melting point and boiling point of distilled water? How do your results compare to your classmates?

The melting point is near 0 ºC, and the boiling point is near 100 ºC. The values should all be within a degree or two of their classmates.

4. Explain how heat and temperature are different.

Temperature is directly related to the movement of the molecules (kinetic energy).

Heat is the total change in the internal energy of a substance. The internal energy comes from both the movement of the molecules (kinetic energy) and the electrostatic forces of attraction between the molecules (potential energy).

Synthesis Questions

Use available resources to help you answer the following questions.

1. When drinking a cold glass of ice water, you notice drops of water forming on the outside of the glass. Explain where the newly formed water came from and what phase change was involved.

The water droplets came from water vapor in air. The phase change was condensation. The water vapor in the air changed into liquid water on the outside of the glass.

2. What phase change occurs when humans sweat? Why does this cool our bodies?

Water changes from a liquid to a gas (evaporation). Sweating cools our body by using body heat to convert water molecules (sweat) into water vapor. The newly formed vapor molecules leave our body and in doing so remove the body heat they have absorbed. Removing heat from something makes it cooler.

3. Is it possible to heat a metal without its temperature increasing?

Yes, if you heat it when it is melting. When heat is added, the temperature of the metal will increase until it reaches its melting point. Once the melting point is reached additional heating will not cause a rise in temperature until the phase change is complete.

4. Is it possible to remove heat from a substance without making the temperature decrease? Use an example to explain your answer.

Yes. If heat is removed from a substance undergoing a phase change, then the temperature will not decrease. For example, if heat was removed from liquid water as it was freezing into ice, then there would be no change in the temperature. Condensation could also be used as an example.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Water at a vigorous boil releases bubbles of gas. What gas is in these bubbles?

- **A.** Air
- **B.** Oxygen gas
- **C.** Nitrogen gas
- **D.** **Water vapor**

2. When heat is added to a substance undergoing a phase change, what happens to the temperature of the substance?

- **A.** It increases
- **B.** It decreases
- **C.** **It remains constant**
- **D.** It decreases slightly and then increases

3. What will happen when heat is added to water that is at room temperature?

- **A.** **The temperature of the water will increase**
- **B.** The temperature of the water will decrease
- **C.** The temperature of the water will remain constant
- **D.** The water will immediately change into a gas

4. At the molecular level, what happens to water molecules as they are heated?

- **A.** The molecules move faster
- **B.** The water molecules are broken apart from other water molecules
- **C.** The hydrogen atoms are broken apart from the oxygen atoms
- **D.** **Either A or B will occur**

5. What is the temperature at which a substance changes from a liquid to a gas known as?

- **A.** **Boiling point**
- **B.** Melting point
- **C.** Freezing point
- **D.** Condensation point

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The **kinetic molecular theory** explains that all matter is made up of particles that are in constant motion. Adding **heat** to matter generally causes the particles to move faster. Removing heat, on the other hand, causes particles to move **slower**. **Temperature** is related to the average motion (kinetic energy) of the particles.

2. All matter exists as a solid, liquid, or a gas, depending on the temperature (and to some extent on the pressure). The process of changing from one of these phases to another is called a **phase change**. The temperature at which a phase change occurs can be used to identify pure substances. There are four main phase changes. The temperature at which a solid changes into a liquid is called the **melting point**. The temperature at which a liquid changes into a gas is called the **boiling point**. Melting and boiling both require the **addition** of heat. The temperature at which a liquid changes into a solid is called the **freezing point**. The temperature at which a gas changes into a liquid is called the **condensing point**. Freezing and condensing both require the **removal** of heat.

3. During a phase change, the temperature of the substance remains **constant**. The temperature remains constant because the heat being added or removed is used to break or form attractions between particles instead of causing the particles to change speed. The energy stored in the attractions between particles is **potential energy**. Therefore, heat and temperature are different. Heat is both kinetic and potential energy of a substance, while temperature is related only to the kinetic energy of the particles.

Extended Inquiry Suggestions

Determine whether different amounts of water affect the melting point or boiling point of water.

Determine the melting point of lauric acid.

Identify unknowns using melting points or boiling points.

Determine the effect of adding solutes on the melting point or boiling point of water.

Design a method to either increase or decrease the temperature at which water boils.

Make ice cream to explore freezing point depression.

Determine the effects of volume on the temperature of water.

Does water melt at a higher temperature than water freezes?

11. Specific Heat

Objectives

Experimentally determine the identity of an unknown metal. Through this investigation, students:

- Differentiate the concepts of temperature and heat
- Understand the fundamental components of specific heat

Procedural Overview

Students conduct the following procedures:

- Record physical observations of the unknown substance
- Use a calorimeter to measure the temperature change inside
- Calculate specific heat of an unknown metal
- Identify the metal from a list of options

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	35 minutes
♦ Lab activity	60 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Fast response temperature sensor[1]
- Beaker, 250-mL
- Beaker, 400-mL
- Graduated cylinder, 100-mL
- Balance, centigram
- Thermometer (or stainless steel temperature sensor)
- Hot plate
- Tongs
- Polystyrene cup (2)
- Lid for the polystyrene cup
- Paper towels
- Tap water, 250 mL
- Distilled (deionized) water, 200 mL
- Metal sample, unknown, up to 4 x 4 x 4 cm[2]

[1] The fast response temperature sensor should not be used in boiling water.

[2] Choices include aluminum, copper, lead, tin, and zinc, among others. Avoid alloys. See Table 1 in the Background section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Energy
- Kinetic molecular theory
- Mole
- Molar mass

Related Labs in This Guide

Labs conceptually related to this one include:

- Phase change
- Heat of Fusion
- Heats of Reaction and Solution
- Hess's Law

Using Your Data Collection System

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- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Showing and hiding data runs in a graph. �(7.1.7)
- Naming a data run �(8.2)

- Measuring the difference between two points in a graph. �(9.2)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

The kinetic molecular theory explains that all molecules are in constant motion. Furthermore, the rate at which the particles travel or vibrate (kinetic energy) is directly proportional to their temperature. The molecules in a hot substance travel or vibrate faster on average than those of a similar sample at a cooler temperature.

Though closely related, temperature is different from heat. Heat energy is used to increase the temperature of a substance. Temperature is a measure of the average energy a particle has within a sample. Heat is the change in the total amount of energy of all the particles in the system summed together.

Temperature is measured in Kelvin or degrees Celsius. Energy is measured in calories or joules (1 calorie = 4.184 J). Do not confuse calorie (lower case "c") with the Calorie (upper case "C") commonly used by nutritionists to describe the energy content of food. Food Calories are actually kilocalories (1 Calorie = 1000 calories = 1 kcal).

The amount of heat required to change the temperature of a substance depends on two factors: the amount of substance present (expressed in grams or moles) and the chemical identity of the substance. The amount of energy (in calories or joules) required to raise the temperature of 1 gram of substance 1 ºC is known as the specific heat *c* for that substance. By definition, the specific heat of water is 1 cal/(g·°C) or 4.184 J/(g·°C). That is, one gram of water requires one calorie or 4.184 J of energy to raise its temperature 1 ºC.

The amount of energy (in joules) required to raise the temperature of 1 mole of substance 1 ºC is known as the molar heat capacity for that substance. One mole of water (18.02 g) requires 75.3 J to raise the temperature 1 ºC. Thus, the molar heat capacity of water is 75.3 J/mol · °C. Table 1 at the end of this section lists several substances with specific heat in J/(g·°C) and molar heat capacity in J/(g·°C).

The first law of thermodynamics is the law of conservation of energy. It states that energy can neither be created nor destroyed, although it can be changed. Because of this, as an object cools, the heat released can be tracked and quantified.

The purpose of a calorimeter is to separate the system being studied from its surroundings. A calorimeter traps heat energy inside the calorimeter so that it can be measured. This experiment uses a simple calorimeter: two polystyrene cups (one nested inside the other) to insulate the contents from the surrounding environment, a known quantity of water to absorb the heat energy, and a temperature sensor to measure the change in the water's temperature.

By knowing the exact amount of water inside the calorimeter and measuring its temperature change, the number of calories (or joules) transferred to the water can be determined using the specific heat of water.

For example, a 25.0 g sample of water inside a calorimeter was heated from 20.0 ºC to 32.3 ºC by a hot object placed inside the water. Each gram of water increased 12.3 °C requiring 12.3 cal/g or 51.5 J/g. These values were calculated using the specific heat of water (1 cal/(g·°C) and

4.18 J/(g·°C)). Since there were 25 g of water, a total of 308 calories (25 g × 12.3 cal/g) or 1290 J (25 g × 51.4 J/g) were transferred from the hot object to the water.

By knowing the temperature change of the object and its mass, the specific heat and molar heat capacity of the object can then be calculated from the 308 cal or 1290 J of energy transferred in a similar fashion, using the general equations below.

$$\text{specific heat} = \frac{\text{energy transferred in calories or joules}}{(\text{g of substance})(\text{change in temperature})}$$

$$\text{molar heat capacity} = \frac{\text{energy transferred in joules}}{(\text{moles of substance})(\text{change in temperature})}$$

Table 1: Specific heat and molar heat capacity of water and select metals

Substance	**Specific Heat (J/(g·°C))**	**Molar Heat Capacity (J/(mol·°C))**
Aluminum	0.902	24.3
Copper	0.385	24.5
Gold	0.129	25.4
Iron	0.450	25.1
Lead	0.129	26.7
Nickel	0.444	26.1
Silver	0.235	25.3
Tin	0.228	27.1
Water	4.184	75.3
Zinc	0.386	25.2

Pre-Lab Discussion and Activity

Temperature and Heat

In order to convey the difference between temperature and heat, present the class with three different containers that hot coffee could be poured into: a pot, a cup, and a spoon. If the coffee poured into each container comes from the same source, then the coffee will initially have the same temperature, but will cool at different rates. Use this example to explain the concept of heat as a change in the total quantity of internal energy. Also, introduce the first law of thermodynamics (conservation of energy).

1. Immediately after pouring coffee from the same source into a pot, a cup, and a spoon, in which container will the coffee have the highest temperature?

They all have the same temperature. Temperature is a measure of the average kinetic energy (movement) of molecules. In the moment immediately after pouring the coffee, the molecules in the different containers are moving at the same speed, so they have the same temperature.

2. The coffee in which container will cool to a lower temperature the fastest? Explain your reasoning.

The coffee in the spoon will cool the fastest. The smallest container contains the least amount of coffee and therefore the smallest amount of internal energy. Loosing even a small amount of energy from the spoonful of coffee represents a greater proportion compared to the other containers.

3. Where did the heat energy transfer to?

The energy was transferred to the surrounding air (and the container—especially if it is a good conductor). The air molecules absorbed the heat. This is the first law of thermodynamics, the law of conservation of energy. This law says that energy cannot be created or destroyed, only changed. Here, the energy did not disappear, it moved from the coffee to the air. As the coffee cooled (lost energy) the air surrounding the coffee warmed (gaining the same amount of energy).

Calorimetry

Using two small blocks of aluminum, a simple calorimeter is presented as a way of trapping the escaping heat inside the container allowing it to warm water instead of air. First, place both aluminum blocks in a beaker of boiling water. After the blocks reach equilibrium with the water, place one block of aluminum on the bench and the other in a beaker of water at room temperature.

4. What is the temperature of the blocks in the boiling water?

The blocks are in equilibrium with the water, so they should both have a temperature equal to the boiling water (100 ºC).

5. After the blocks are removed from the boiling water, what happens to them? What happens to the air surrounding one block and the water surrounding the other?

The blocks cool as they lose heat energy to their surroundings. The air and water warm as the heat transfers from the blocks to their surroundings.

6. In which system is it easier to measure the amount of heat transferred: the beaker of water or the open bench? Why?

The beaker of water traps the heat energy into a small space more efficiently and can be measured using a thermometer. Heat transferred into the air moves away too quickly (dissipates). This heat cannot be measured accurately.

7. Will the water in the beaker stay warm indefinitely? Where is its heat energy going?

The water in the beaker eventually cools to room temperature as heat is lost through the glass and out the top of the beaker. Eventually, the heat transfers to the surrounding air, as did the block of aluminum on the bench.

8. Is there a way to prevent or minimize the loss of heat from the water to the surroundings? Have you ever used something similar to keep food or drinks hot?

Instead of a glass beaker, use a material that insulates to prevent or slow the transfer of heat energy. A polystyrene cup (or two nested together) with a plastic lid minimizes the loss of heat to the surroundings. Hot drinks served in polystyrene cups keep the drinks hot for a longer period of time. Glass vacuum flasks are also very efficient at keeping soup or drinks hot for long periods.

Specific Heat

Explain that different objects require different amounts of energy to increase their temperature by the same amount. This property of matter is called specific heat. Specific heat is defined as the amount of energy required to raise the temperature of 1 gram of a substance 1 ºC. Demonstrate the concept of specific heat by showing that objects at the same initial temperature transfer different amounts of heat to water in a calorimeter. To do this, boil a block of aluminum and a block of iron for 10 minutes. Remove each block from the boiling water, pat it dry with a paper towel and place it in the calorimeter. Display the temperature change in the calorimeter for the students to compare. Discuss the results with the students.

9. Immediately after removing the blocks from the boiling water, which block had the higher temperature?

The blocks should each have the same temperature, equal to that of the boiling water (~100 °C).

10. Why do you need to dry the blocks before placing them in the calorimeters?

The water that comes out on the blocks is hot and will add to the temperature increase in the calorimeter, but will not have come from the block. The added water will also add to the mass of the water in the calorimeter, which is important in calculating the heat transfer.

11. Did the water in each calorimeter reach the same temperature? Explain.

No the water did not reach the same temperature. This means that a different amount of heat was transferred from each block. This shows that the different metals at the same initial temperature are able to transfer different amounts of energy.

The actual results depend on the mass of the blocks used. If the masses of the two blocks are equal, then the temperature of the calorimeter with aluminum reaches a higher temperature because the specific heat of aluminum is greater than that of iron. This means that more energy needed to be transferred to the aluminum in order for it to raise its temperature to that of boiling water (100 °C).

The specific heat of aluminum is 0.902 J/(g·°C) and that of iron is 0.450 J/(g·°C).

12. If both blocks were put into a fire, which would take longer to increase its temperature by 10 °C?

It depends on the mass of the blocks; if equal masses were placed in the fire then the aluminum takes longer to increase its temperature because it requires a greater input of energy in order to raise its temperature 1 °C.

Lab Preparation

Although this activity requires no specific lab preparation, allow 10 minutes to assemble the equipment needed to conduct the lab.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Be aware that boiling water and steam can cause painful burns.
- Avoid coming in contact with hot equipment and glassware.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

5	4	3	1	2
Transfer the metal to the calorimeter and continue to collect data until the temperature stabilizes.	Place a temperature sensor in the colorimeter and start recording temperature data.	Add 100 mL of distilled water to the calorimeter and find the mass of calorimeter and water together.	Heat the metal in boiling water.	As the metal is being heated in the boiling water, construct the calorimeter and find the mass of the empty calorimeter.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a fast response temperature sensor to the data collection system. �(2.1)

3. ☐ Create a graph display of Temperature (°C) versus Time (s). �(7.1.1)

4. ☐ Heat 250 mL of tap water in a 400-mL beaker on a hot plate to boiling.

5. ☐ Construct a simple calorimeter by placing one polystyrene cup into another polystyrene cup.

Collect Data

6. ☐ Describe the metal sample (color, luster, density).

Answers will vary depending on the metal sample provided. Some possibilities (arranged by metal identity) are:

- gray, dull, very light for its size (aluminum)
- reddish-gold, dull, heavy for its size (copper)
- silvery-gray, shiny, heavy for its size (iron)
- bluish-gray, dull, very heavy for its size (lead)
- silvery, shiny, heavy for its size (nickel)
- gray, dull, light for its size (tin)
- bluish-gray, dull, heavy for its size (zinc)

7. ☐ Measure the mass of the metal sample and record the value in grams in Table 2 below.

Table 2: Measurements of metal sample, calorimeter and water

Measurement	Metal Trial 1	Metal Trial 2
Mass of metal sample (g)	21.42	21.42
Mass of 250-mL beaker and calorimeter, empty (g)	110.25	110.25
Mass of 250-mL beaker and calorimeter, with water (g)	210.68	210.91
Temperature of the boiling water (°C)	97.5	97.5

8. ☐ Add the metal to the boiling water and allow it to heat for at least 5 minutes.

9. ☐ While the metal is heating, measure the mass of an empty calorimeter in a 250-mL beaker and record the value in grams in Table 2 above.

10. ☐ Using a 100-mL graduated cylinder, add 100.0 mL distilled water to the calorimeter.

11. ☐ Measure the mass of the beaker and calorimeter with the water and record the value in grams in Table 2 above.

12. ☐ Why is the mass of the calorimeter determined both with and without water?

By subtracting the weights, the exact amount of water in the calorimeter can be determined.

13. ☐ Place the fast response temperature sensor into the calorimeter, making sure it is submerged in the water.

14. ☐ Using a thermometer (or stainless steel temperature sensor), record the temperature of the boiling water and record the value in degrees Celsius in Table 2 above.

CAUTION: The fast response temperature sensor is not designed for temperatures above 70 °C.

15. ☐ What is the temperature of the metal sample in the boiling water?

By allowing the metal to heat for five minutes, equilibrium is reached, and the metal sample is the same as the temperature of the boiling water.

16. ☐ While viewing the graph display, start recording data. �(6.2)

17. ☐ Adjust the scale of the graph. �(7.1.2)

18. ☐ Allow the temperature to stabilize (remain constant for at least 30 seconds).

19. ☐ With a pair of tongs, remove the metal sample from the boiling water and quickly dry the metal and the tongs with a paper towel.

CAUTION: The metal will be hot!

20. ☐ Why is it necessary to dry the metal sample before adding it to the calorimeter?

Excess water on the metal sample adds to the volume of water in the calorimeter. Excess water also carries additional energy that reduces the accuracy of the results.

21. ☐ Carefully, without splashing any of the water from inside the calorimeter, use the tongs to add the metal to the calorimeter.

22. ☐ Cover the calorimeter with a lid.

23. ☐ Why must you avoid splashing water out of the calorimeter?

Heat energy leaving the metal sample is absorbed by the water in the calorimeter. If there is less water in the calorimeter, the temperature of the system measures higher than expected. This leads to the erroneous impression that amount of energy released and the specific heat value are higher.

24. ☐ After the temperature has reached a maximum and remains constant for at least two minutes, stop recording data. �(6.2)

25. ☐ What is the temperature of the metal sample in the calorimeter at the end of the experiment?

It is the same as the temperature of the water in the calorimeter.

26. ☐ Name the data run as “metal trial 1”. �(8.2)

27. ☐ Use the tongs to remove the metal from the calorimeter and dispose of the water.

28. ☐ Repeat the steps in the Collect Data section, recording the values in the second trial column in Table 2 above.

29. ☐ Name the data run as “metal trial 2”. �(8.2)

30. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

1. ☐ Use the graph of Temperature (°C) versus Time (s) to determine the final temperature, initial temperature, and change in temperature (ΔT) for the water in the calorimeter. Follow the steps below to complete this on your data collection system:

a. Display the run of data you want to analyze. �(7.1.7)

b. Measure the difference between the final temperature and the initial temperature. When this step is complete, the final temperature, initial temperature, and change in temperature will be displayed on the screen. �(9.2)

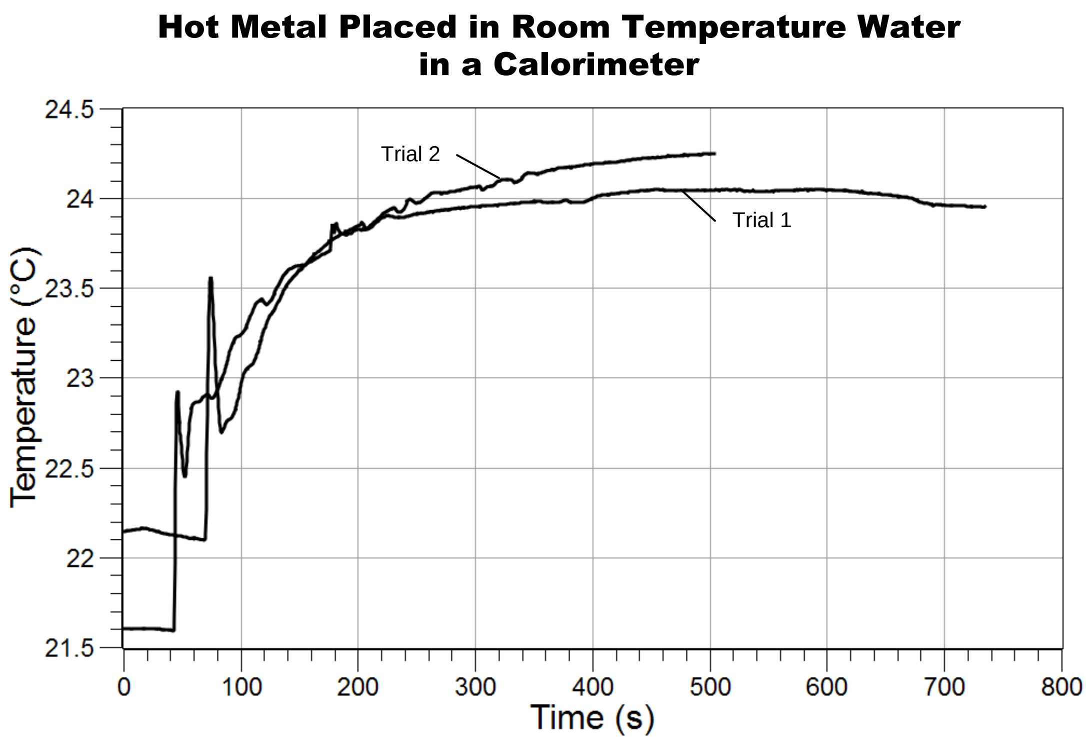
c. Record the values for each trial in Table 3 below.

Table 3: Temperature changes in the calorimeter

Parameter	Trial 1	Trial 2
T_{final} (˚C) of water in the calorimeter	24.0	24.2
T_{initial} (˚C) of water in the calorimeter	21.6	21.9
$\Delta T = T_{\text{final}} - T_{\text{initial}}$ (˚C)	2.4	2.3

2. ☐ Create a graph with both runs of data displayed on your data collection system. �(7.1.3)

3. ☐ Sketch or print a copy of the graph of Temperature (°C) versus Time (s) for both trials of data. Label each trial as well as the overall graph, the x-axis, the y-axis, and include numbers on the axes. �(11.2)



4. ☐ Calculate the mass of water in the calorimeter for each trial. Record your answers in Table 4 below.

Table 4: Mass of water in the calorimeter

Parameter	Trial 1	Trial 2
Mass of 250-mL beaker and calorimeter, with water (g)	210.68	210.91
Mass of 250-mL beaker and calorimeter, empty (g)	110.25	110.25
Mass of the water (g)	100.43	100.66

5. ☐ Determine the amount of energy in Joules transferred from the metal block to the water using the information provided below. Show your work in Table 5 below.

$q = m \times c \times \Delta T$

q = energy transferred

m = mass of water in the calorimeter

c = specific heat of water = 4.184 J/(g·°C)

ΔT = change in temperature in the colorimeter

Table 5: Calculating energy in joules transferred

Parameter	Trial 1	Trial 2
Show your work here	$(100.43\ \text{g})\left(\frac{4.184\ \text{J}}{\text{g}\cdot{}^{\circ}\text{C}}\right)(2.4\ {}^{\circ}\text{C}) = 1008.5\ \text{J}$	$(100.66\ \text{g})\left(\frac{4.184\ \text{J}}{\text{g}\cdot{}^{\circ}\text{C}}\right)(2.3\ {}^{\circ}\text{C}) = 968.7\ \text{J}$
Joules transferred (J)	1.0×10^3	9.7×10^2

6. ☐ Calculate the temperature change for the metal sample. Record the answer in Table 6 below.

Table 6: Change in temperature of the unknown metal

Parameter	Trial 1	Trial 2
T_{initial} (°C) of the metal (in the boiling water)	97.5	97.5
T_{final} (°C) of the metal (in the calorimeter at the end of the experiment)	24.0	24.2
$\Delta T = T_{\text{initial}} - T_{\text{final}}$ (°C)	73.5	73.3

7. ☐ Using the mass of the metal sample, its temperature change, and the number of joules of energy transferred to the water, calculate the specific heat of the metal sample for each trial and then find the average of the two trials. Show your work in Table 7 below.

$$c = \frac{q}{m \cdot \Delta T}$$

c = specific heat
q = energy transferred
m = mass of the unknown metal
ΔT = change in temperature of the metal

Table 7: Calculating specific heat

Parameter	**Trial 1**	**Trial 2**
Show your work here:	$\frac{1.0 \times 10^3 \text{ J}}{(21.42 \text{ g})(73.5\,^\circ\text{C})} = 0.64 \text{ J}/(\text{g} \cdot {}^\circ\text{C})$	$\frac{9.7 \times 10^2 \text{ J}}{(21.42 \text{ g})(73.3\,^\circ\text{C})} = 0.62 \text{ J}/(\text{g} \cdot {}^\circ\text{C})$
Specific heat (J/(g·°C))	0.64	0.62
Average specific heat (J/(g·°C))	$\frac{(0.64 \text{ J}/(\text{g} \cdot {}^\circ\text{C}) + 0.62 \text{ J}/(\text{g} \cdot {}^\circ\text{C}))}{2} = 0.63$	

Analysis Questions

1. What is the identity of the metal you used in this experiment? (Use Table 1 in the Background section and the color, luster, and density you described from the question at the beginning of the Collect Data section of the Procedure.)

Answers will vary depending on the identity of the metal sample and the accuracy of the results. The sample data was collected using an aluminum block.

2. What is the percent error for the experimentally determined specific heat?

$$\textbf{percent error} = \left| \frac{(\text{accepted value} - \text{experimental value})}{\text{accepted value}} \right| \times 100$$

$$\text{percent error} = \left| \frac{(0.90 - 0.63)}{0.90} \right| \times 100 = 30\%$$

3. Based on the identity of the metal sample and the average specific heat value you determined, what is its molar heat capacity?

$$\left(\text{average specific heat} \frac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}\right)\left(\frac{\text{g atomic weight}}{1 \text{ mole}}\right) = \text{molar heat capacity} \frac{\text{J}}{\text{mol} \cdot {}^\circ\text{C}}$$

$$\left(\frac{0.63 \text{ J}}{\text{g} \cdot {}^\circ\text{C}}\right)\left(\frac{26.98 \text{ g}}{1 \text{ mol}}\right) = 17 \text{ J}/(\text{mol} \cdot {}^\circ\text{C})$$

Synthesis Questions

Use available resources to help you answer the following questions.

1. Can other liquids besides water be used in a calorimeter?

Yes, other liquids can be used in a calorimeter, because all substances can transfer energy.

2. What would you need to change in the procedure or calculations when using some other liquid?

A liquid other than water might require a different amount of energy to realize the change in temperatures. This means that a different specific heat applies in the calculations. Also, the amount of heat energy being transferred and the specific heat of the liquid may require a different volume of liquid in the calorimeter.

3. In what situation would water be unsuitable for calorimetry?

If the water would freeze or boil because of the material added to the calorimeter, a liquid with a different freezing point or boiling point or specific heat would be needed.

4. How can the concept of molar heat capacity help explain why large bodies of water can dramatically affect local climate?

Because water has a high molar heat capacity, it can store large amounts of energy from the sun. When this heat transfers to colder air, it can maintain milder climate. For example, the Gulf Stream current in the Atlantic Ocean carries warm water from the tropics northward towards the United Kingdom. This causes milder temperatures in that area compared with other locations at the same latitude.

Another example involves on-shore and offshore breezes. These result from air being heated over the land or water. The convection currents created in the air produce wind that flows from colder areas to warmer areas. (Warm air rises, drawing in cooler air to replace it).

5. Brass is an alloy of copper and zinc. What is the specific heat of a sample of brass weighing 11.8 g if it requires 197.1 J to increase its temperature by 44.3 ºC?

$$\frac{197.1\ \text{J}}{(11.8\ \text{g})(44.3\ ^\circ\text{C})} = 0.377\ \text{J}/(\text{g}\cdot{}^\circ\text{C})$$

6. What is the final temperature if 100.0 mL of water at 10.0 ºC absorbed 1000 J?

$$4.184\ \text{J}/(\text{g}\cdot{}^\circ\text{C}) = \frac{1000\ \text{J}}{(100.0\ \text{g})(T_\text{f} - 10.0\ ^\circ\text{C})}$$

$$\left(\frac{4.184\ \text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(100.0\ \text{g})(T_\text{f} - 10.0\ ^\circ\text{C}) = 1000\ \text{J}$$

$$(T_\text{f} - 10.0\ ^\circ\text{C}) = \frac{1000\ \text{J}}{(100.0\ \text{g})\left(\frac{4.184\ \text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)} = 2.39\ ^\circ\text{C}$$

$$T_\text{f} = 2.39\ ^\circ\text{C} + 10.0\ ^\circ\text{C} = 12.4\ ^\circ\text{C}$$

7. When a 15.0 °C block of an unknown plastic weighing 42.5 g is added to a calorimeter with 50.0 mL of water at 25.0 °C, the temperature of the water decreases to 24.3 °C. What is the specific heat of the plastic?

Note: A negative number of joules indicates energy transferred from the water. A positive number of joules indicates energy absorbed by the plastic. The amount of energy is the same.

$$\left(\frac{4.184\ \text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(50.0\ \text{g})(24.3\ {}^\circ\text{C} - 25.0\ {}^\circ\text{C}) = -146\ \text{J}$$

$$-146\ \text{J} = +146\ \text{J}$$

$$\frac{+146\ \text{J}}{(42.5\ \text{g})(24.3\ {}^\circ\text{C} - 15.0\ {}^\circ\text{C})} = 0.370\ \text{J}/(\text{g}\cdot{}^\circ\text{C})$$

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. How many joules are required to increase the temperature of a 30.0 g block of gold from 25.0 °C to 50.0 °C? The specific heat of gold is 0.129 J/(g·°C).

- **A.** 4.184 J
- **B.** **96.8 J**
- **C.** 193.5 J
- **D.** 3138 J

2. Which of the following describes substances with lower values for specific heat?

- **A.** **Better conductors of heat**
- **B.** Better insulators of heat
- **C.** Worse conductors of heat
- **D.** None of the above

3. Finding a specific heat of 0.233 J/(g·°C) probably indicates a sample of what metal?

- **A.** Aluminum
- **B.** Nickel
- **C.** **Silver**
- **D.** Tin

4. What calculation converts specific heat to molar heat capacity?

- **A.** Multiply by the mass of the sample
- **B.** Divide by the mass of the sample
- **C.** **Multiply by the molar mass**
- **D.** Divide by the molar mass

5. What is the specific heat of iron in cal/(g·°C)?

- **A. 0.108**
- **B.** 0.450
- **C.** 1.88
- **D.** None of the above

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Energy is classified in one of two ways: **potential** (stored energy) and **kinetic** (energy in motion). **Heat** is a form of energy in motion. It is measured in calories or joules. Heat is a measure of the change in the **total** amount internal energy in a sample. Temperature is different from heat. It is measured in units of kelvin or degrees Celsius. Temperature is the **average** amount of molecular motion in a sample.

2. Heat energy is used to increase the temperature of a substance. The amount of heat required to raise the temperature of **1 g** of substance by **1 °C** is called the **specific heat**. For example, one gram of lead requires 0.129 J of energy to increase its temperature from 24.0 to 25.0 °C. The amount of heat required to raise the temperature of one mole of substance by one degree Celsius is called the **molar heat capacity**.

3. The amount of heat transferred can be trapped inside a **calorimeter**. This device insulates the system being studied from its surroundings. Most often, water is inside the calorimeter. Because this substance has a specific heat of **4.184 J/(g·°C)**, it can absorb a lot of energy with a **small** change in temperature. By measuring the increase in the temperature of the water, the amount of energy transferred inside the calorimeter can be determined.

Extended Inquiry Suggestions

Repeat the experiment using metal samples that have been cooled in a refrigerator instead of heated in boiling water.

Repeat the experiment with a variety of other solid substances, including various plastics, glass, and metal alloys.

Compare the specific heat values of various coins. Include coins from different countries. Consider comparing one-cent coins from the United States produced before and after 1982.

Using calorimetry, design and test an apparatus to determine the caloric content of a variety of nuts.

Use the Dulong-Petit law (specific heat = 3R/M, where M is molar mass and R is the universal gas constant in J/(mol·K)) to determine the molar masses of pure metal samples.

12. Heat of Fusion

Objectives

Determine the heat of fusion for water. Through this investigation, students:

- Use calorimetry to determine the energy exchanged
- Determine the molar heat of fusion of ice
- Observe and describe how temperature changes before, during, and after a phase change

Procedural Overview

Students conduct the following procedures:

- Record the temperature change of heated water resulting from the addition of ice cubes
- Use the data collected to calculate the heat of fusion of ice
- Evaluate the results by calculating the percent error from an accepted value

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	15 minutes
♦ Lab activity	40 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Fast response temperature sensor
- Graduated cylinder, 100-mL
- Beaker (2), 250-mL
- Hot plate
- Polystyrene cup (2)
- Lid for polystyrene cup
- Paper towels
- Water, 200 mL
- Ice cubes (2)

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Calorimetry
- ♦ Molecular arrangement of solids, liquids, and gases

Related Labs in This Guide

Labs conceptually related to this one include:

- ♦ Specific Heat
- ♦ Phase Change
- ♦ Heats of Reaction and Solution
- ♦ Hess's Law

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- ♦ Starting a new experiment on the data collection system �(1.2)
- ♦ Connecting a sensor to the data collection system �(2.1)
- ♦ Recording a run of data �(6.2)
- ♦ Displaying data in a graph �(7.1.1)
- ♦ Adjusting the scale of a graph �(7.1.2)
- ♦ Displaying multiple data runs on a graph �(7.1.3)
- ♦ Showing and hiding data runs in a graph �(7.1.7)
- ♦ Naming a data run �(8.2)
- ♦ Measuring the difference between two points in a graph �(9.2)
- ♦ Saving your experiment �(11.1)
- ♦ Printing �(11.2)

Background

Intermolecular forces (attractions between molecules) are responsible for holding molecules together to form solids, liquids, and gases. In solids, the attractions are strong and lock the molecules in place, preventing them from moving away from each other. In liquids, the attractions are strong enough to hold the molecules together, but not strong enough to lock them in a rigid structure. The weaker attractions allow the liquid molecules to flow around one another. In gases, the attractions are too weak to hold the particles together. Gas particles move independently from one another with space between the gas particles.

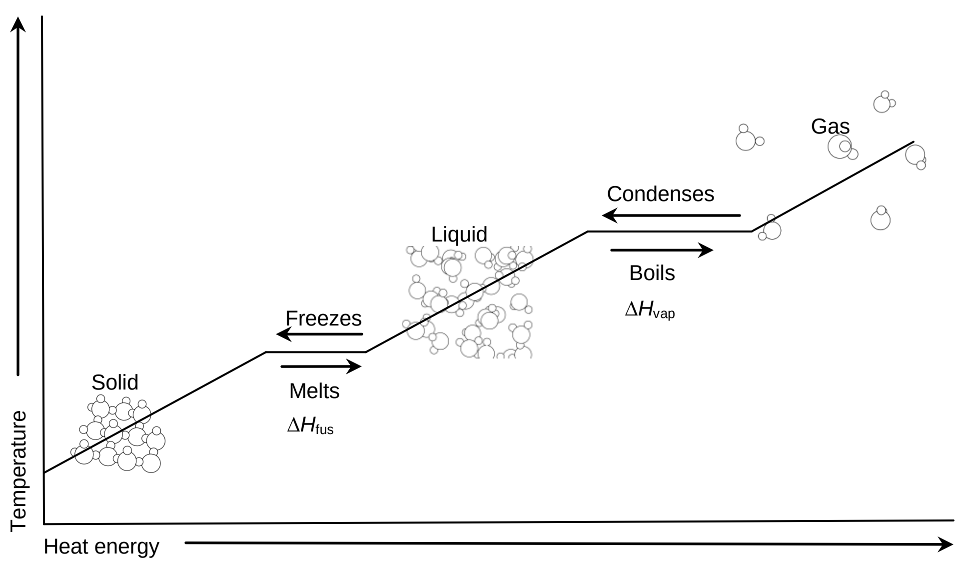
The relative strengths of these attractions depend on the kinetic energy of the system and the chemical structure of the molecule. For example, substances containing strong hydrogen bonds require greater amounts of energy to change their physical state.

The kinetic energy of molecules relates directly to temperature. Adding heat energy to a system either warms the substance (by increasing the vibrations or movement of the molecules) or overcomes the attractions between the molecules, allowing the substance to change its physical state.

Adding heat energy first increases the temperature of a substance until it reaches a point where a phase change can occur (melting point or boiling point). Once this temperature is reached, adding more energy breaks the attractions between the molecules.

The amount of energy required to break or form the attractions in a solid is called the heat of fusion (ΔH_{fus}). For the evaporation or condensation of a liquid, it is the heat of vaporization (ΔH_{vap}). Because the heat energy is being used to break attractions and not to increase the kinetic energy of the molecules, no temperature change is observed during a phase change.

Phase Changes



The first law of thermodynamics (conservation of energy), states that energy cannot be created or destroyed only changed. This means that the energy used to increase the temperature is given back when the substance cools. It also means that energy added to break the attractions (an endothermic process) is given back when the attractions form again (an exothermic process). This allows for the measurement of the heat of fusion through calorimetry.

A calorimeter prevents heat energy from escaping the system by insulating the substance being studied from its surroundings. The heat absorbed by the substance during fusion (melting) is taken from water inside a calorimeter. By knowing the amount of water and its temperature change, the amount of energy transferred to break the attractions during melting can be determined. From this, the heat of fusion for the substance inside the calorimeter can be calculated.

In this lab, students calculate the heat of fusion ΔH_{fus} of ice. To do this, students place an ice cube in warm water (50 °C) and measure the change in temperature of the water as the ice absorbs the heat and melts. In addition to measuring the change in temperature, students also need to know the exact amount of warm water (in grams) that is placed in the calorimeter as well as the exact amount of ice (in grams). The amount of energy transferred q_{trans} from the warm water to the ice is calculated using the equation below.

$$q_{\text{trans}} = m_{\text{H}_2\text{O}} \times c \times \Delta T$$

In this equation, m is the mass of warm water, c is the heat capacity of water (4.184 J/(g·°C)), and ΔT is the change in temperature of the warm water as the ice melted.

Because of the first law of thermodynamics (conservation of energy), the energy transferred that is lost by one substance in the calorimeter is gained by the other. In this case, the energy transferred was lost by the water q_{lost} and gained by the ice q_{gain}.

$$q_{\text{lost}} = -q_{\text{gain}}$$

In the equation above, the value of q_{lost} is always negative. This indicates a flow of energy *out* of the system (the water). What we are interested in, is the energy gained q_{gain} by the ice, which has the same value as q_{trans}, but the opposite sign:

$q_{\text{lost}} = q_{\text{trans}}$ (negative value)
$q_{\text{gain}} = -q_{\text{trans}}$ (positive value because it is the negative of a negative value).

The energy gained by the ice $-q_{trans}$ is used to melt the ice q_{melt} and to bring the temperature of the melted water up to the final temperature $q_{0\,°\text{C}\to T_f}$. This is summarized in the equation below.

$$-q_{\text{trans}} = q_{\text{melt}} + q_{0\,°\text{C}\to T_f}$$

Since the heat of fusion of ice is only the energy required to melt the ice q_{melt}, the energy used to heat the melted ice up to the final temperature needs to be determined. Once determined, this value must be subtracted from the total heat transferred (gained by the ice). To determine the energy absorbed by the freshly melted ice to warm from 0 °C to the final temperature the equation below is used.

$$q_{0\,°\text{C}\to T_f} = m_{\text{ice}} \times c \times \left(T_f - 0\,°C\right)$$

In this equation, m is the mass of ice, c is the heat capacity of water (4.184 J/(g·°C)), and ΔT is the change in temperature of the freshly melted ice.

The energy required to melt the ice (solid at 0 °C to a liquid at 0 °C) can be calculated by subtracting the energy absorbed from the freshly melted ice water from the total amount of energy transferred to the ice.

$$q_{\text{melt}} = -q_{\text{trans}} - q_{0\,°\text{C}\to T_f}$$

Finally, the heat of fusion can be calculated by dividing the energy required to melt the ice (q_{melt}) by the number of moles of ice. Students must convert the grams of ice measured to moles.

$$\Delta H_{\text{fus}} = \frac{q_{\text{melt}}}{\text{moles}}$$

Pre-Lab Discussion and Activity

Phase Change

Review the arrangement of molecules in solids, liquids, and gases and the intermolecular forces holding them together. This is initially presented in the Phase Change experiment. Remind students that molecules in solids are packed closely and held rigidly in place. Molecules in liquids are still held together, but the attractions are not strong enough to hold the molecules in one place. Molecules in a liquid are able to flow. Gases have little or no attractions between the molecules, allowing them to move freely with space between them.

1. When the forces of attraction between molecules in a solid are broken, to what physical state does the substance change? What is this process called?

Liquids form when solids melt. The process is called fusion.

2. Does breaking attractions require or release energy? Is it an endothermic or exothermic process?

Breaking attractions always requires energy. It is an endothermic process.

3. When a gas becomes a liquid are attractions broken or formed? Is this endothermic or exothermic?

When condensation occurs, attractions are forming between the gas molecules to produce a liquid. This releases energy as an exothermic process.

With a stainless steel temperature sensor, display the temperature of a small beaker of water as it is heated from room temperature to boiling. As the water warms, explain that heat is being added to the water causing the molecules to move faster. Temperature is the measurement of the average kinetic energy (molecular motion) in the sample. The water remains liquid because energy is being used to heat the water, not to break the attractions between the molecules. Once the boiling point is reached, the energy is no longer being used to increase the kinetic energy (the temperature stops increasing). Instead it is being used to break the attractions between the molecules. Consequently, water vapor is formed.

Teacher Tip: If a gas burner is available, the boiling of water can be done in an unwaxed paper cup for dramatic effect! Because water has a high specific heat and heat of vaporization, the water absorbs the energy preventing the cup from igniting.

4. As the temperature of the water increases where does the energy from the hot plate go?

It is being absorbed by the water molecules and used to increase their speed.

5. When the water is boiling where does the energy from the hot plate go?

It is being used to break the attractions between the water molecules.

6. Why is the temperature of the water not increasing while the water is boiling?

Before boiling, the energy is used to heat the water and increase the speed of the molecules but not break any attractions between them. When the boiling point is reached (100 ºC), the energy is used to break the attractions between the molecules but not to increase their speed.

The amount of energy required to break the attractions between molecules is known as latent heat. The latent heat required to change a solid into a liquid is called the heat of fusion (ΔH_{fus}). The latent heat required to change a liquid into a gas is called the heat of vaporization (ΔH_{vap}). The amount of heat required depends on the strength of the attractions between the molecules. This depends on the molecular structure of the substance and the amount of substance present. Latent heat is often expressed in kilojoules per mole (kJ/mol).

The reverse process is equal in energy, but opposite in direction. For example, melting ice requires an input of 6.01 kJ/mol to break the attractions between molecules, however, to change liquid water into ice requires removing 6.01 kJ/mol as the attractions are formed. This is the reason steam burns hurt so much—to turn liquid water into a vapor requires 40.65 kJ/mol; when it changes from a gas back to a liquid the energy is released. If it condenses on your hand, all that energy goes into your skin!

7. If 250 kJ is required to melt a block of ice, how much energy is removed from the water when the block of ice is formed?

Because the processes are opposites, the same amount of energy is transferred from the ice to freeze it as is absorbed when it melts. In this case, the amount is 250 kJ.

Calorimetry

Show a simple calorimeter with two nested polystyrene cups, a lid, and the fast response temperature sensor. Remind students that the calorimeter insulates and prevents the flow of heat between the system being studied and its surroundings. The water inside the calorimeter stores heat. By measuring its temperature change, the specific heat of the water can be used to determine the amount of energy transferred inside the calorimeter. The applicable equation is $q = m \times c \times \Delta T$, where q is the energy transferred, m is the mass of the water, c is specific heat of water (4.184 J/(g·°C)), and ΔT is the temperature change of the water ($\Delta T = T_{final} - T_{initial}$).

Because of the first law of thermodynamics (conservation of energy), the energy lost by one substance in the calorimeter is gained by the other. In the equation $q_{lost} = -q_{gain}$, the value of q_{lost} is always negative. This indicates a flow of energy *out* of the system. Since energy is required to break the attractions in a solid when it melts, the energy lost from the water inside the calorimeter is used to overcome the molecular attractions in the solid. This can be measured, leading to the value for the heat of fusion.

8. Would a paper cup or a metal container work as a calorimeter?

No, heat energy would escape from the calorimeter easily and the water would not trap all of the heat. This makes the system inefficient and introduces substantial error into the measurements.

9. If 100.0 mL of water inside a calorimeter decreases from 50.0 °C to 25.0 °C, how much energy is lost from the water?

Use the density of water (ρ = 1.000 g/mL) to find the mass of the water in the calorimeter:

$$\text{mass} = \rho \times V$$

$$\text{mass} = \left(\frac{1.000\ \text{g}}{\text{mL}}\right)(100.0\ \text{mL}) = 100.0\ \text{g}$$

Then, use the mass of the water, the specific heat of water, and the temperature change to calculate the energy lost:

$$q = m \times c \times \Delta T$$

$$q = (100.0\ \text{g})\left(\frac{4.184\ \text{J}}{\text{g} \cdot {}^\circ\text{C}}\right)(25.0\ {}^\circ\text{C} - 50.0\ {}^\circ\text{C}) = -10{,}460\ \text{J} = -10.5\ \text{kJ}$$

10. How much energy would be gained by the system in the calorimeter described above?

$$q_{\text{lost}} = -q_{\text{gain}}$$

$$-q_{\text{lost}} = q_{\text{gain}}$$

$$-(-10.5\ \text{kJ}) = q_{\text{gain}} = +10.5\ \text{kJ}$$

Lab Preparation

These are the materials and equipment to set up prior to the lab.

The day before the experiment, prepare enough ice cubes to provide each group with two. Alternatively, acquire a bag of ice cubes on the day of the experiment.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Avoid being burned from contact with the hot plate. Allow it to cool thoroughly before storing.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	3	4	1	5
Record the mass of the calorimeter and the 50 ºC water.	Place a temperature sensor in the 50 °C water and start recording the temperature of the calorimeter water.	Add an ice cube to the 50 °C water and then cover the calorimeter. Continue to record data until the ice melts.	Heat 150 mL of water to 50 ºC and pour it into a calorimeter.	Remove the lid and sensor then record the mass of the calorimeter, water, and melted ice.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a fast response temperature sensor to the data collection system. �(2.1)

3. ☐ Display a graph of Temperature (°C) versus Time (s). �(7.1.1)

4. ☐ Using a graduated cylinder, add 100 mL of water to a 250-mL beaker.

5. ☐ Use a hot plate to warm the water to approximately 50 ºC.

6. ☐ Place one polystyrene cup in the other one to create a simple calorimeter and then place the calorimeter inside an empty 250-mL beaker for stability.

7. ☐ What happens to the temperature of the warm water when the ice is added?

The temperature of the water decreases.

Collect Data

8. ☐ Measure and the mass of the beaker plus the empty calorimeter to two decimal places. Record the mass below.

Mass of the beaker plus the empty calorimeter (g): ____________________ 110.25 g

9. ☐ Pour the heated water into the calorimeter.

10. ☐ Measure the mass of the beaker plus the calorimeter with the water to two decimal places. Record the mass below.

Mass of the beaker plus calorimeter with water (g): ____________________ 207.01 g

11. ☐ How can the mass of water in the calorimeter be calculated?

The mass of the water can be calculated by subtracting the mass of the empty calorimeter from the mass of the calorimeter with water.

12. ☐ Insert the temperature sensor into the calorimeter making sure it is submerged in the water.

13. ☐ Start recording data. �(6.2)

14. ☐ Adjust the scale of the graph. �(7.1.2)

15. ☐ Allow the temperature to stabilize.

16. ☐ Use a paper towel to dry an ice cube.

17. ☐ Why is it necessary to dry the ice cube with a paper towel before placing it in the calorimeter?

This removes any water that has already melted.

18. ☐ Without splashing any water from the calorimeter, add the ice cube and then cover the calorimeter with the lid.

19. ☐ What happens to your results if some of the water in the calorimeter splashes out when you add the ice?

The mass of the water in the calorimeter becomes greater than it actually is, which produces error in the total amount of heat absorbed by the ice.

20. ☐ Gently swirl the calorimeter without splashing any water.

21. ☐ As soon as the ice finishes melting, stop recording data. �(6.2)

22. ☐ Name the data run "ice trial 1". �(8.2)

23. ☐ Remove the lid and the temperature sensor.

24. ☐ Measure the mass of the beaker plus the calorimeter with the water and melted ice to two decimal places. Record the mass below.

Mass of the beaker plus calorimeter with water and melted ice (g): 218.94 g

25. ☐ Repeat the process. Record the values in Table 1 below.

Table 1: Collected masses for trial 2

Mass	**Ice Trial 2**
Beaker plus empty calorimeter (g)	110.26
Beaker plus calorimeter with water (g)	210.38
Beaker plus calorimeter with water and melted ice (g)	225.56

26. ☐ Name the data run "ice trial 2".

27. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

1. ☐ Calculate the amount of water in the calorimeter before adding the ice. Record it in Table 2 below.

Table 2: Mass of water in the colorimeter

Mass	**Ice Trial 1**	**Ice Trial 2**
Beaker plus calorimeter with water (g)	207.01	210.38
Beaker plus empty calorimeter (g)	110.25	110.26
Water (calorimeter with water – calorimeter empty) (g)	96.76	100.12

2. ☐ Calculate the amount of ice added to the calorimeter. Record it in Table 3 below.

Table 3: Mass of ice in the colorimeter

Mass	**Ice Trial 1**	**Ice Trial 2**
Beaker plus calorimeter with water and melted ice (g)	218.94	225.56
Beaker plus calorimeter with water (g)	207.01	210.38
Ice (calorimeter with melted ice – calorimeter with water) (g)	11.93	15.18

3. ☐ Determine the number of moles of ice added to the calorimeter. Use the following equation, and record your work in Table 4 below.

$$(\text{g ice})\left(\frac{1\ \text{mol}}{18.02\ \text{g}}\right) = \text{moles ice}$$

Table 4: Moles of ice

	Ice Trial 1	**Ice Trial 2**
Moles of ice (show your work)	$(11.93\ \text{g})\left(\frac{1\ \text{mol}}{18.02\ \text{g}}\right) = 0.6620\ \text{mol}$	$(15.18\ \text{g})\left(\frac{1\ \text{mol}}{18.02\ \text{g}}\right) = 0.8424\ \text{mol}$

4. ☐ Use the graph of Temperature (°C) versus Time (s) to determine the final temperature, initial temperature, and change in temperature for the calorimeter water. Follow the steps below to complete this on your data collection system:

a. Display the run of data you want to analyze. �(7.1.7)

b. Measure the difference between the final temperature and the initial temperature. When this step is complete, the final temperature, initial temperature, and change in temperature will be displayed on the screen. �(9.2)

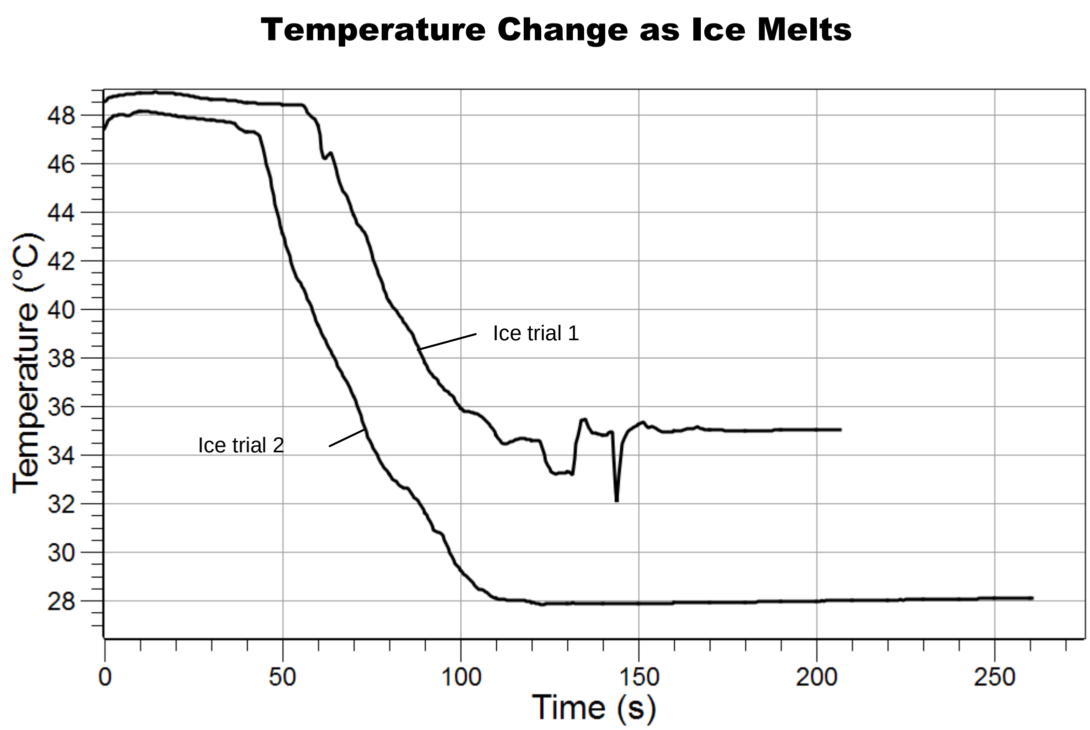
c. Record the values for each trial in Table 5 below.

Table 5: Collected and calculated temperature

Temperature	**Ice Trial 1**	**Ice Trial 2**
Final temperature (˚C)	31.3	27.9
Initial temperature (˚C)	48.2	47.7
Change in temperature, $\Delta T = T_{\text{final}} - T_{\text{initial}}$ (˚C)	–16.9	–19.8

5. ☐ Create a graph with both runs of data displayed on your data collection system. �(7.1.3)

6. ☐ Sketch or print a graph of Temperature (°C) versus Time (s). Make sure that each trial is labeled as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



7. ☐ Calculate the energy transferred from the water in the calorimeter to the ice q_{trans} and the energy gained by the ice from the water $-q_{trans}$. Use the following equation, and show your work in Table 6 below.

$$q_{trans} = m_{H_2O} \times c \times \Delta T$$

Table 6: Energy transferred from water to ice

	Ice Trial 1	**Ice Trial 2**
Energy transferred from the water (J) q_{trans}	$96.76\text{ g} \times \frac{4.184\text{ J}}{\text{g} \cdot {}^{\circ}\text{C}} \times -16.9\ {}^{\circ}\text{C} = -6840$	$100.10\text{ g} \times \frac{4.184\text{ J}}{\text{g} \cdot {}^{\circ}\text{C}} \times -19.8\ {}^{\circ}\text{C} = -8290$
Energy absorbed by the ice (J) $-q_{trans}$	$-(-6840\text{ J}) = 6840$	$-(-8290\text{ J}) = 8290$

8. ☐ The energy gained by the ice $-q_{\text{trans}}$ is used to melt the ice q_{melt} and to bring the temperature of the melted water up to the final temperature $q_{0\,^\circ\text{C}\rightarrow T_f}$.

$$-q_{\text{trans}} = q_{\text{melt}} + q_{0\,^\circ\text{C}\rightarrow T_f}$$

Determine the energy absorbed by the freshly melted ice to warm from 0 ºC to the final temperature. Use the following equation, and record your work in Table 7 below.

$$q_{0\,^\circ\text{C}\rightarrow T_f} = m_{\text{ice}} \times c \times \left(T_f - 0\,^\circ\text{C}\right)$$

Table 7: Energy used to warm the melted ice water

	Ice Trial 1	**Ice Trial 2**
Show your work here $q_{0\,^\circ\text{C}\rightarrow T_f}$	$11.93\text{ g} \times \dfrac{4.184\text{ J}}{\text{g}\cdot{}^\circ\text{C}} \times \left(31.3\,^\circ\text{C} - 0\,^\circ\text{C}\right)$	$15.18\text{ g} \times \dfrac{4.184\text{ J}}{\text{g}\cdot{}^\circ\text{C}} \times \left(27.9\,^\circ\text{C} - 0\,^\circ\text{C}\right)$
Energy used to warm the ice water (J)	1560	1770

9. ☐ Determine the energy required to melt the ice (solid at 0 ºC to a liquid at 0 ºC). Use the following equation, and record it in Table 8 below.

$$q_{\text{melt}} = -q_{\text{trans}} - q_{0\,^\circ\text{C}\rightarrow T_f}$$

Table 8: Energy required to melt the ice

	Ice Trial 1	**Ice Trial 2**
Show your work here	q_{melt} = 6840 J – 1560 J	q_{melt} = 8290 J – 1770 J
Energy used to melt the ice q_{melt} (J)	5280	6520

10. ☐ Calculate the heat of fusion for ice (in kJ/mol) using the amount of energy absorbed by the ice and the number of moles of ice in the calorimeter. Use the following equation and record it in Table 9 below.

$$\Delta H_{\text{fus}} = \frac{q_{\text{melt}}}{\text{moles}}$$

Table 9: Heat of fusion

	Ice Trial 1	Ice Trial 2
Show your work	$(5280\text{ J})\left(\frac{1\text{ kJ}}{1000\text{ J}}\right) = 5.28\text{ kJ}$ $\Delta H_{\text{fus}} = \frac{5.28\text{ kJ}}{0.6620\text{ mol}} = 7.98\text{ kJ/mol}$	$(6520\text{ J})\left(\frac{1\text{ kJ}}{1000\text{ J}}\right) = 6.52\text{ kJ}$ $\Delta H_{\text{fus}} = \frac{6.52\text{ J}}{0.8424\text{ mol}} = 7.74\text{ kJ/mol}$
Heat of fusion (kJ/mol)	7.98	7.74

11. ☐ Calculate the average value for the heat of fusion of ice from the two trials.

$$\frac{(7.98\text{ kJ/mol} + 7.74\text{ kJ/mol})}{2} = 7.86\text{ kJ/mol}$$

Analysis Questions

1. The accepted value for the heat of fusion for ice is 6.01 kJ/mol. What is the percent error for the experimentally determined heat of fusion?

$$\textbf{percent error} = \left|\frac{(\textbf{accepted value} - \textbf{experimental value})}{\textbf{accepted value}}\right| \times 100$$

$$\text{percent error} = \left|\frac{(6.01\text{ kJ/mol} - 7.86\text{ kJ/mol})}{6.01\text{ kJ/mol}}\right| \times 100 = 30.8\%$$

2. What are some possible sources of error and how might you correct or at least account for them?

The calorimeter is not completely efficient at trapping heat. Some heat energy escapes through the walls and the lid. Using polystyrene cups with thicker walls or a better lid may improve heat retention but will not entirely eliminate the exchange of heat with the surroundings. Mathematically adding lost energy back into the value for energy can account for the loss. An experimentally determined "calorimeter constant" is a way to account for lost energy.

The temperature of ice removed from a freezer and immediately put into the calorimeter is below 0 ºC. Inside the calorimeter, heat energy would be consumed as the ice warms to its freezing point before melting. Making sure that the starting temperature of the ice is correctly represented in the calorimetry calculations can improve accuracy.

Melt water on the surface of the ice cube adds mass to the calorimeter. The ice cube should be completely dried with a paper towel before immediately adding it to the calorimeter.

Synthesis Questions

Use available resources to help you answer the following questions.

1. How much energy would be required to melt a 50.0 kg block of ice at 0 ºC? The heat of fusion for ice is 6.01 kJ/mol.

$$(50.0\ \text{kg})\left(\frac{1000\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol}}{18.02\ \text{g}}\right) = 2770\ \text{mol}$$

$$(2770\ \text{mol})\left(\frac{6.01\ \text{kJ}}{1\ \text{mol}}\right) = 16700\ \text{kJ}$$

2. How are heat of fusion and heat of vaporization similar? How are they different?

Both the heat of fusion and the heat of vaporization are similar in that they are a measure of the amount of energy required to overcome intermolecular attractions in a particular physical state (heat of fusion for solids, heat of vaporization for liquids). This is measured in terms of energy per unit mass, most often as kilojoules per mole of substance.

They are different from one another in that the heat of fusion is the amount of energy required to break the rigid attractions in a solid, allowing the molecules to flow around one another as a liquid, while the heat of vaporization is the amount of energy required to separate the molecules of a liquid from one another so they are able to move freely in space as a gas.

Use the following information to answer Synthesis Questions 3 and 4 below:

An unknown substance that has a mass of 250.0 g reaches its melting point in a coffee cup calorimeter. As the substance changes from solid to liquid its temperature remains unchanged. The mass of the water in the calorimeter is 500.0 g. The water has an initial temperature of 90.0 ˚C and drops to 19.0 ˚C as the unknown solid melts.

3. How many kJ of energy the water release while the solid was melting:

$$q = m \times c \times \Delta T$$

$$q = 500.0\ \text{g} \times \frac{4.184\ \text{J}}{\text{g} \cdot {}^\circ\text{C}} \times (19.0\ {}^\circ\text{C} - 90.0\ {}^\circ\text{C})$$

$$q = -149000\ \text{J} = -149\ \text{kJ}$$

4. If an unknown substance has a formula weight of 77.3 g/mol, what is the heat of fusion for the substance?

$$(250.0\text{ g})\left(\frac{1\text{ mol}}{77.3\text{ g}}\right) = 3.23\text{ mol}$$

$q_{lost} = -q_{gain}$

$-q_{lost} = q_{gain}$

$-(-149\text{ kJ}) = q_{gain} = +149\text{ kJ}$

$$\Delta H_{fus} = \frac{q}{\text{moles}}$$

$$\Delta H_{fus} = \frac{149\text{ kJ}}{3.23\text{ mol}} = 46.1\text{ kJ/mol}$$

5. How is heat of fusion different from specific heat? For which is there no observed temperature change?

Specific heat is the amount of energy required to increase only the temperature (the average kinetic energy) of the system, while the heat of fusion is the amount of energy required to break the intermolecular attractions. There is no observed temperature change for heat of fusion. As heat is added to break the intermolecular forces, there is no change in temperature (no change in kinetic energy of the molecules.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. What happens to the average kinetic energy of the molecules as ice changes to liquid water at 0 °C?

- **A.** Increases
- **B.** Decreases
- **C.** **Remains the same**
- **D.** Unable to determine with this information

2. The heat of fusion for a substance is the energy measured during what?

- **A.** **Phase change from solid to liquid**
- **B.** Phase change from a liquid to a gas
- **C.** Temperature change
- **D.** Change in kinetic energy

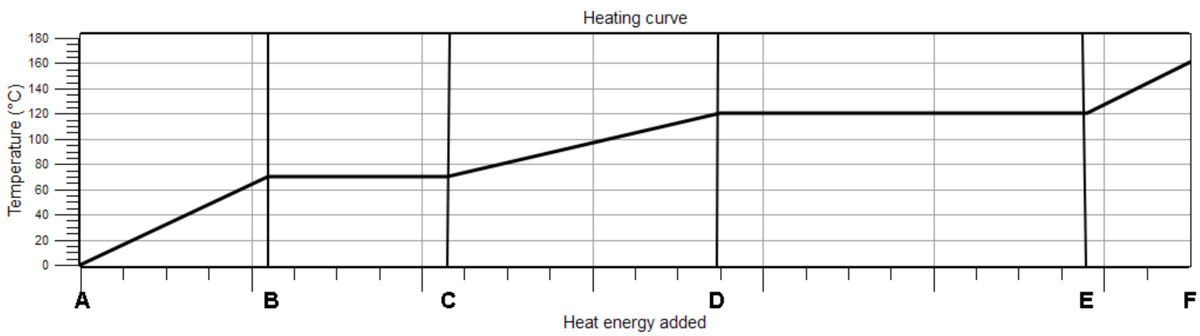
3. When 20.0 g of a substance is completely melted at its melting point, 820 joules of energy are absorbed. What is the heat of fusion for this substance?

- **A.** 0.024 J/g
- **B.** 16.4 J/g
- **C.** 24 J/g
- **D.** **41 J/g**

4. The heat of fusion for ethanol is 5.02 kJ/mol. How much energy is required to melt 24.5 g of ethanol at its freezing point? The molar mass of ethanol is 46.07 g/mol.

- **A.** 0.205 kJ
- **B.** **2.67 kJ**
- **C.** 5.67 kJ
- **D.** 123 kJ

Use the graph below to answer Multiple Choice Questions 5 and 6. The graph represents the relationship between temperature and energy added to a substance. At point "A", the substance is a solid.



5. What is the freezing point of the substance depicted in the graph above?

- **A.** 20 ºC
- **B.** **70 ºC**
- **C.** 120 ºC
- **D.** 160 ºC

6. During which interval does the heat absorbed by the substance represent the heat of vaporization?

- **A.** B to C
- **B.** C to D
- **C.** **D to E**
- **D.** E to F

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. When a substance changes physical state, it requires **energy**. The molecules must overcome and break free from their **attraction** to each other. These forces are called **intermolecular forces**. In water, they are hydrogen bonds. The energy necessary to overcome these forces is called the **latent heat**. Specifically, the energy required to change a substance from a solid to a liquid is the heat of **fusion**, and the energy required to change from a liquid to a gas is the heat of **vaporization**. Fusion is another term for **melting**.

2. The stronger the intermolecular forces within a substance, the **greater** the latent heat is. Water has particularly high latent heats because its hydrogen bonds are among the **strongest** types of intermolecular forces. The stronger, or more tightly, they "hang on" to one another, the **more** energy that is necessary to pull them apart. The **same** amount of energy that water absorbs when it melts is released when it freezes. An input of energy, however, can be used either to overcome the attractions or increase the kinetic energy of the molecules. Increasing the kinetic energy can be seen as an increase in **temperature**. During an **endothermic** phase change requiring an input of energy, such as fusion or vaporization, the temperatures **remain constant**. After the phase change is complete, further addition of energy causes the temperature to **increase**.

Extended Inquiry Suggestions

Determine the heat of fusion for other substances with melting points less than 70 °C (or 95 °C, if using a stainless steel temperature sensor).

Compare and contrast the heat of fusion values for various waxes, fatty acids, and oils, especially saturated and unsaturated oils.

Compare the melting points and heats of fusion for different concentrations of frozen salt water mixtures.

Design and test a method to determine the heat of vaporization for water.

13. Intermolecular Forces

Objectives

Students determine the effects of molecular size and shape on the strength of intermolecular forces for different alcohols within the same homologous series and between isomeric pairs. Through this investigation, students:

- Explain why evaporation causes a decrease in temperature
- Describe the relationship between evaporation rate and the strength of intermolecular forces of attraction among molecules
- Describe the relationship between the size of molecules in a homologous series and the strength of intermolecular forces of attraction between them
- Describe the influence the shape of a molecule has on its vapor pressure by comparing the evaporation rates of two isomeric alcohol pairs

Procedural Overview

Students conduct the following procedures:

- Graph temperature versus time data for five alcohols from the same homologous series and two pairs of isomeric alcohols as they evaporate from a stainless steel temperature sensor
- Compare the rate of evaporation for the seven alcohols to determine the relative strengths of the intermolecular forces of attraction

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	50 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Stainless steel temperature sensor[1]
- Graduated cylinder, 10-mL
- Test tube (7), 15-mm x 100-mm
- Test tube rack
- Stopper (7), to fit the test tubes
- Wash bottle and waste container
- Masking tape (2), 6 cm strips
- Methanol (CH_3OH), 5 mL[2]
- Ethanol (C_2H_5OH), 5 mL[2]
- Propanol (C_3H_7OH), 5 mL[2]
- Butanol (C_4H_9OH), 5 mL[2]
- Pentanol (C_3H_7OH), 5 mL[2]
- 2-Propanol (C_3H_7OH), 5 mL[3]
- 2-Butanol (C_4H_9OH), 5 mL[3]

[1] A fast-response temperature sensor is not appropriate for this experiment.

[2] The homologous series portion of the experiment can be performed using as few as three of the alcohols listed, provided that one of them is either butanol or propanol when used as a comparison in the molecular shape portion of the experiment.

[3] The molecular shape portion of the experiment can be performed using either 2-butanol or 2-propanol, provided that an isomeric alcohol is also investigated as part of the homologous series (2-butanol paired with butanol and 2-propanol paired with propanol).

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Kinetic molecular theory
- Bonding
- Electronegativity
- Polar and non-polar molecules
- Average kinetic energy
- Pressure

Related Labs in This Guide

Labs conceptually related to this one include:

- Phase Change
- Properties of Ionic and Covalent Compounds

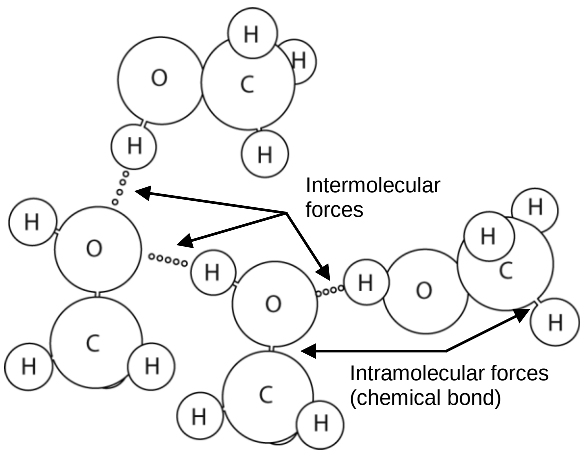
Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

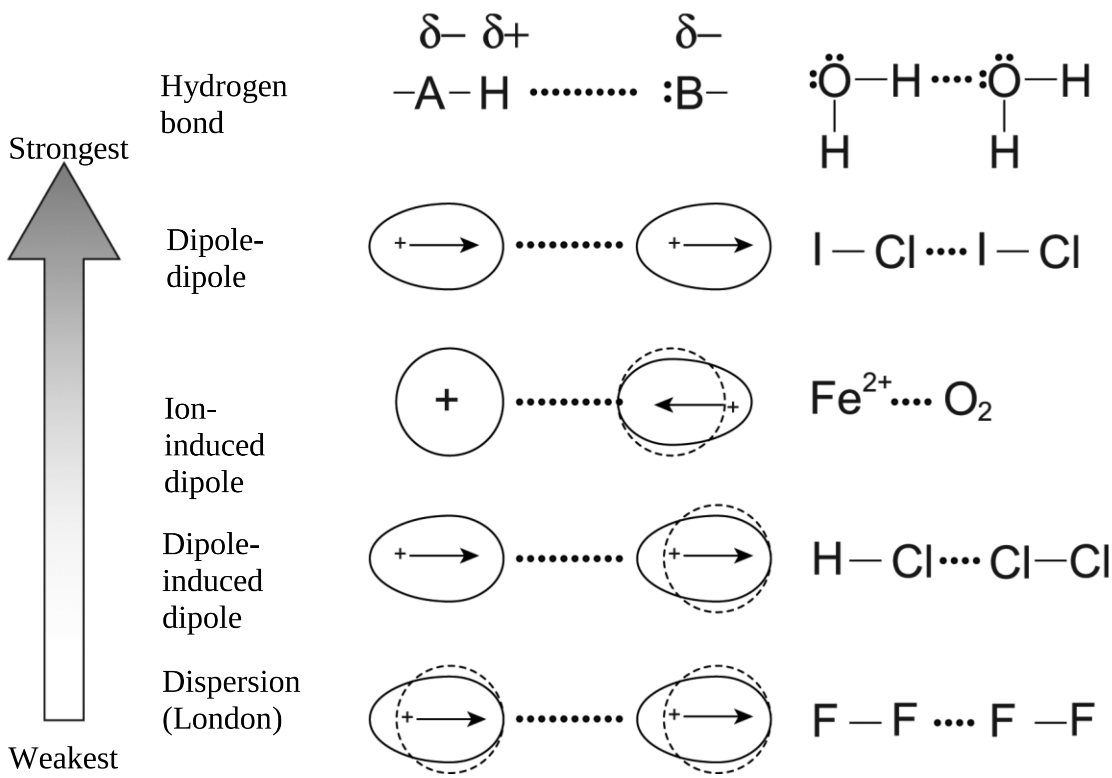
- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to your data collection system �(2.1)
- Recording a run of data �(6.2)
- Display data a in graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Selecting data points in a graph �(7.1.4)
- Showing and hiding data runs in a graph. �(7.1.7)
- Naming a data run �(8.2)
- Measuring the difference between two points in a graph �(9.2)
- Finding the slope and intercept of a best-fit line �(9.6)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

Matter is made up of atoms, which are held together by electrostatic forces of attraction. The two main classifications of these attractions are intramolecular forces and intermolecular forces. Intramolecular forces are chemical bonds holding atoms together. These form when the positive charge of a nucleus strongly attracts the negative charge of the electron clouds of neighboring atoms. Covalent bonding, ionic bonding, and metallic bonding are the three types of intramolecular forces. Intermolecular forces are the much weaker attractions between neighboring molecules. These are not chemical bonds. Breaking them only serves to separate the molecules from one another and not change them into new substances. The picture below illustrates the difference between the intramolecular and intermolecular forces found in methanol.



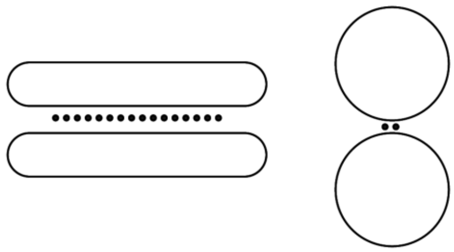
Intermolecular forces arise from a number of different interactions between molecules. The first type of intermolecular force is dipole-dipole attractions, which occur between polar molecules. Dipole-dipole attractions result when the positive end of a polar molecule is attracted to the negative end of another polar molecule. Another type of intermolecular force is charge-induced dipole interactions (either ion-induced dipole or dipole-induced dipole). These are the result of a charged species causing a non-polar molecule's electrons to shift within the electron cloud, creating a temporary dipole in the non-polar molecule. The third type of intermolecular forces is dispersion (London) forces, which occur between non-polar molecules. Even though these molecules are non-polar, there exists a temporary dipole induced from a nearby molecule whose electrons are momentarily not uniform. These attractions are very weak, but tend to increase as the number of electrons and molecular mass (size) of the molecules increase. The final intermolecular force is hydrogen bonding. Hydrogen bonding is a special dipole-dipole interaction that occurs only between molecules containing hydrogen directly bonded to either fluorine, oxygen, or nitrogen. Hydrogen bonds are significantly stronger than other dipole-dipole attractions because hydrogen is such a small atom and fluorine, oxygen, and nitrogen are highly electronegative.



Intermolecular forces can be classified in order of strength as dispersion (London) (weakest), dipole-induced dipole, ion-induced dipole, dipole-dipole, and hydrogen bonding (strongest). The strength of intermolecular forces within a substance can explain many of its physical properties, including its physical state at room temperature, its melting and boiling points, and its vapor pressure. At room temperature, substances with strong intermolecular forces exist as solids, while substances with very weak intermolecular forces exist as gases. When heat is added to a substance, the particles in the substance increase their average kinetic energy. With higher kinetic energies, more particles can overcome the intermolecular attractions holding them together and escape as individual gas particles.

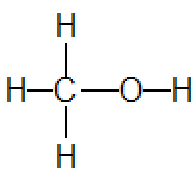
Evaporation occurs when a substance in a liquid phase changes to vapor. The rate at which a substance evaporates depends on several variables including surface area, temperature, and the vapor pressure of the liquid. The larger the exposed surface area, the greater the rate of evaporation because there are more molecules in position to escape. The evaporation rate will also increase when the temperature of the liquid increases. An increase in average kinetic energy means that more particles have high enough energies to overcome the intermolecular attractions. Lastly, liquids with the weakest intermolecular forces have the highest vapor pressures. Vapor pressure is a measure of the tendency for a liquid to convert to a gas. When the vapor pressure is equal to the overlying atmospheric pressure, the liquid will boil. Thus, liquids with higher vapor pressures boil at lower temperatures. It follows, then, that liquids with higher vapor pressures will also evaporate faster. They evaporate faster because less energy is required to overcome the weaker intermolecular attractions.

In this investigation, students will use measured evaporation rates to determine how the size and shape of molecules making up different liquids affects the strength of the intermolecular forces. They will use the homologous series of primary alcohols to investigate the influence of size. As the size of the molecule increases, the rate of evaporation decreases. From this, students can conclude that larger molecules have lower vapor pressures because of stronger intermolecular forces. They will use the isomeric alcohols to investigate the influence of molecular shape. As the shape of the molecule becomes more spherical, the rate of evaporation increases. This is a result of reduced surface contact area: long molecules have many interactions along their sides, but spheres can only make contact at one point. From this, students can conclude that more spherical molecules have higher vapor pressures, and thus, weaker intermolecular interactions.

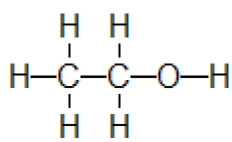


Surface contact area of molecules with different shapes.

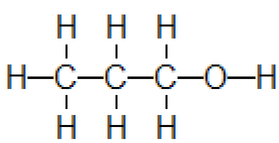
The structural formulas of the alcohols used in this investigation are given below.



Methanol, CH_4O

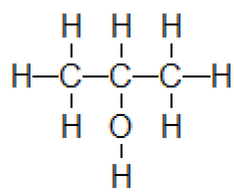


Ethanol, C_2H_6O

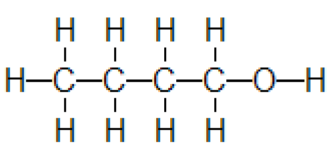


Propanol, C_3H_8O

C_3H_8O ← Isomeres →

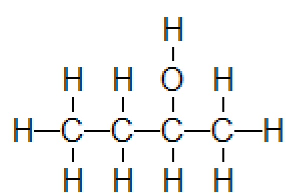


2-Propanol, C_3H_8O

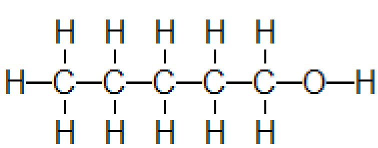


Butanol, $C_4H_{10}O$

$C_4H_{10}O$ ← Isomeres →



2-Butanol, $C_4H_{10}O$



Pentanol, $C_5H_{12}O$

Pre-Lab Discussion and Activity

Observations of Water and Nail Polish Remover

Working in pairs, have the students place a drop of nail polish remover on the back of one hand and a drop of water on the back of the other hand. Have the students observe the shape (how much each drop spreads out) and feel (heat transfer) of the drops. Have the students wave their hands and compare the feel of the nail polish remover and water drops. Engage the students in a discussion about the similarities and differences between nail polish remover (acetone or ethyl acetate) and water. Explain that the liquids feel "cool" because the heat required to evaporate the molecules is being drawn from their hands.

1. Compare the shape of the drop of water to the shape of the drop of nail polish remover.

The water drop forms a bubble, and the drop of fingernail polish remover spreads out into a thin layer.

2. Which drop feels colder when you wave your hands?

The nail polish remover feels colder than the water.

3. How are nail polish remover and water similar?

They are both clear liquids. They both feel cold when applied to the skin.

4. Are there any differences between nail polish remover and water?

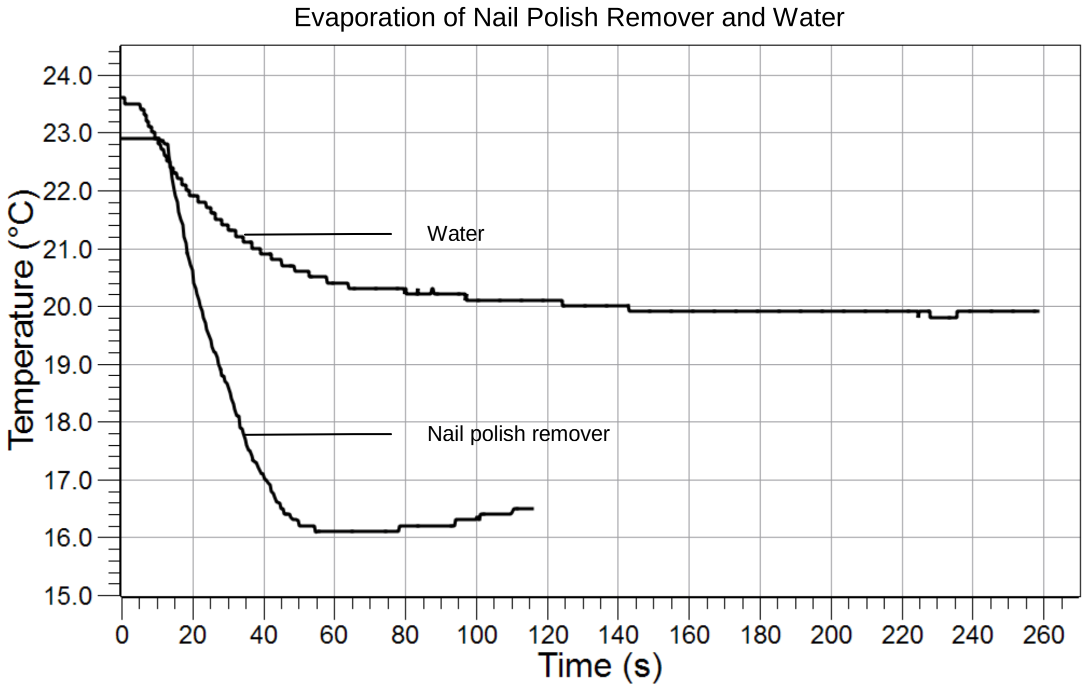
Nail polish remover has an odor and water does not. Nail polish remover may also be colored and water is not.

5. Why do the liquids make your skin feel "cool?" What is evaporation?

The liquid is evaporating. Evaporation is the process of changing from a liquid to a vapor.

Collect Temperature Data of Nail Polish Remover and Water

Have the students predict what will happen to the temperature when nail polish remover and water are evaporated from the stainless steel temperature sensor. Place a stainless steel temperature sensor in a test tube containing approximately 5 mL of water. Project a Temperature (°C) versus Time (s) graph for the students to see. Start recording data, and then remove the temperature sensor from the liquid. Tape it so that it is hanging down from a table. Stop recording data when the temperature stabilizes (2 to 4 minutes). Repeat the procedure with nail polish remover. Have the students compare the rate of evaporation of water with that of nail polish remover, and have them suggest explanations for the difference.



6. What do you think will happen to the temperature when water evaporates from a stainless steel temperature sensor?

The temperature will decrease as the water evaporates.

7. What do you think will happen to the temperature when nail polish remover evaporates from a stainless steel temperature sensor? How will it compare to water?

The temperature will decrease as the nail polish remover evaporates. The nail polish remover should evaporate at a faster rate and reach a lower temperature than the water.

8. How does the change in the temperature of water compare to that of nail polish remover?

The evaporation of water caused the temperature to decrease from 23.6 °C to 20.2 °C for a difference of 3.4 °C. The evaporation of nail polish remover caused the temperature to decrease from 22.8 °C to 16.2 °C for a difference of 6.6 °C. The evaporation of nail polish remover decreased the temperature more than the evaporation of water.

9. How does the rate of evaporation for water compare to that of nail polish remover?

Water evaporated at a slower rate than the nail polish remover. Rate of evaporation is the temperature difference divided by the time difference and is equal to the slope of the line. The slope can be determined by creating a linear fit to the first 20 seconds of data.

10. Why does the nail polish remover evaporate at a faster rate than water?

Water molecules are attracted to each other stronger than the molecules making up nail polish remover (acetone or ethyl acetate) are attracted to each other. This will be explained in the next part using magnets.

Magnets and Intermolecular Forces

Find sets of magnets with varying strengths to represent intermolecular forces with varying strengths. Have the students pull the magnets apart and compare the ease with which different magnets come apart relative to the strength with which the sets of magnets are held together. Engage the students in a discussion about how the varying strengths of the magnets compare to the differences in evaporation rates between water and nail polish remover.

In the discussion, explain to the students that molecules are held together by attractive forces between them similar to the way magnets are attracted to each other. The strength of the attractive forces holding molecules together vary as well; some molecules are held together very tightly and have strong intermolecular forces of attraction, while other molecules are held together by weaker forces. During evaporation, the liquid molecules obtain energy from their environment (such as the heat in the stainless steel temperature sensor). Liquids held together with weaker forces have higher vapor pressures and evaporate quicker than those held together with stronger forces (with lower vapor pressures).

11. Describe how the different sets of magnets compare to each other?

Some sets of magnets can be pulled apart easily, and others are much more difficult to pull apart.

12. How do the sets of magnets relate to our observations about the varying evaporation rates between water and nail polish remover?

The magnets demonstrate attractive forces. Just as the magnets are held together by attractive forces, the water molecules and nail polish remover molecules are held together in the liquid state by attractive forces. Energy (from our muscle power) is needed to separate the magnets just like energy (from heat) is needed to separate molecules.

13. Explain why a drop of water formed a bubble, but the nail polish remover spread evenly over the skin.

The molecules in water are strongly attracted to each other (cohesion). The nail polish remover “molecules” are not as strongly attracted to each other and, therefore, spread over the skin.

14. Explain why water evaporates at a slower rate than nail polish remover.

Water is held together by stronger intermolecular forces of attraction. The water molecules are harder to separate from each other and require more energy to separate, resulting in a slower rate of evaporation.

15. How can you use temperature changes and evaporation rates to compare the intermolecular forces holding molecules together in a variety of liquids?

Larger changes in temperature and faster evaporation rates mean higher vapor pressures because of weaker intermolecular forces. Smaller changes in temperature and slower evaporation rates mean lower vapor pressures because of stronger intermolecular forces.

Homologous Series of Primary Alcohols

Hold up a sample of each alcohol for the students to see. Tell the students the goal of this lab is to determine the relative strength of intermolecular forces between the different alcohols and to determine if there is a relationship between the size of the molecules and the strength of the intermolecular forces holding them together. Engage the students in a review of the types of intermolecular forces and the molecular structures in the homologous series of primary alcohols. Use pictures similar to those in the Background section to help the students visualize the intermolecular forces involved.

16. Give a physical description of the primary alcohols.

The alcohols are clear, colorless liquids that have an odor.

17. Do the primary alcohols have weak, moderate, or strong intermolecular forces? How do you know?

The primary alcohols have moderate intermolecular forces because they are liquids at room temperature. Solids have strong intermolecular forces and gases have weak intermolecular forces.

18. What is a homologous series? Give examples.

A homologous series is one in which all the members have the same general formula, differing only by a CH_2 group. Substances that make up a homologous series have similar physical and chemical properties. Straight chain hydrocarbons are an example of a homologous series, the first four of which are methane, ethane, propane, and butane. All follow the general formula, C_nH_{2n+2}.

19. Given the chemical formula of methanol (CH_3OH) and ethanol (CH_3CH_2OH), predict the chemical formula of propanol (with three carbon atoms), butanol (with four carbon atoms), and pentanol (with five carbon atoms).

The chemical formula for propanol is $CH_3CH_2CH_2OH$.

The chemical formula for butanol is $CH_3CH_2CH_2CH_2OH$.

The chemical formula for pentanol is $CH_3CH_2CH_2CH_2CH_2OH$.

20. What are the different types of intermolecular forces? What types of molecules involve each type of intermolecular force?

Dispersion (London) forces exist between non-polar molecules. Dipole-induced dipole interactions exist between either an ion or a polar molecule and a non-polar molecule. Dipole-dipole interactions exist between polar molecules. Hydrogen bonding takes place between molecules possessing a hydrogen atom directly bonded to either a nitrogen, oxygen, or fluorine atom; the hydrogen on the nitrogen, oxygen or fluorine of one molecule is attracted to the nitrogen, oxygen, or fluorine of another molecule.

21. What are the types of intermolecular forces involved in primary alcohols? How do you know?

All alcohols have a hydrogen attached directly to an oxygen atom. For primary alcohols, the alcohol end of one molecule can form a hydrogen bond with the alcohol end of a neighboring molecule.

22. What happens to the size of the molecules in a homologous series?

The size increases.

23. How can you determine if there is a relationship between the size of a molecule and the strength of its intermolecular forces?

Evaporate each type of alcohol. The alcohols that evaporate quicker and have a greater change in temperature have higher vapor pressures caused by weaker intermolecular forces.

Isomeric Alcohol Pairs

Provide the students with a magnet, a flat steel bar, and a steel sphere to investigate the effect surface contact area (shape) has on the strength of the interactions. As was done with the magnet sets in the Magnets and Intermolecular Forces section, have the students pull the bar and sphere from the magnet. Even though the forces of attraction between the magnet and steel are identical, the shape of the object plays an important role in the amount of energy needed to overcome the attractions.

Engage the students in a discussion about the differences in contact area affecting the amount of effort required to pull the steel objects from the magnet. It is the same with molecules and intermolecular forces: long, straight molecules have a lot of surface contact area to form attractions, while spherical molecules can only interact at a single point of contact. Using isomers (molecules with the same chemical formula, but a different arrangement of atoms), the students will see that primary alcohols will have lower vapor pressures because of stronger interactions between molecules resulting in lower evaporation rates than secondary and tertiary alcohols. Use pictures similar to those in the background section to help the students visualize the arrangement of atoms in the isomers used in this experiment compared to their straight chain counterparts.

24. Which steel object was easier to separate from the magnet? Explain.

The sphere was easier to separate, because it only came in contact with a very small area of the magnet.

25. Which would require more energy to evaporate: a liquid made up of long, flat molecules or one made up of round, spherical molecules? Explain.

The liquid made up of long, flat molecule will require more energy to evaporate because they have more contact area between molecules.

Lab Preparation

Although this activity requires no specific lab preparation, allow 10 minutes to gather the equipment needed to conduct the lab.

Teacher Tips:

- You do not need to use all 5 alcohols to see the trend in the homologous series. The students will get good results using as few as three alcohols.
- When performing this lab, make certain to have good ventilation in the room. Butanol and pentanol have strong odors.
- If you would like to minimize waste of the alcohols used in this experiment, you can have the first group of students set up and label the test tubes and then re-use the test tubes for each of the remaining class periods. Alternatively, pour small amounts into either 50-mL or 100-mL beakers for the students to measure from. When they are done with evaporating each alcohol they can then pour the unused alcohol back into the beaker for the next group or class.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Use a wafting motion when smelling chemicals.
- ♦ Alcohols are flammable. Smother pentanol and 2-butanol fires with sand or a Type B fire extinguisher.
- ♦ Ensure that there is good ventilation in the room. Butanol and pentanol have strong odors.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	5	3	4
Place the temperature sensor in the test tube containing methanol.	Measure 5 mL of each alcohol into separately labeled test tubes.	Repeat the process for each of the other alcohols.	Start recording data, and then remove the temperature sensor from the methanol.	Stop data recording when the temperature begins to increase.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Prepare test tubes with methanol, ethanol, propanol, butanol, pentanol, 2-propanol, and 2-butanol using the following procedure:

a. Use a graduated cylinder to measure approximately 5 mL of an alcohol.

b. Transfer the alcohol into a test tube, seal it with a stopper, and label the test tube.

c. Clean the graduated cylinder by rinsing it several times with water.

2. ☐ Why is it necessary to stopper a test tube containing alcohol?

The stopper will prevent the alcohol from evaporating.

3. ☐ Predict what will happen to the temperature as each alcohol evaporates from the temperature sensor.

The temperature will decrease.

4. ☐ Predict how the temperature changes will compare among the five different alcohols of the homologous series (methanol, ethanol, propanol, butanol, and pentanol). Explain your prediction.

The difference between the starting temperature and the ending temperature should get smaller as the molecule gets larger. This is because the intermolecular attractions are stronger as the molecules increase in size.

5. ☐ Predict how the temperature changes will compare between the isomeric forms of the alcohols (2-propanol versus propanol, and 2-butanol versus butanol). Explain your prediction.

The difference between the starting temperature and the ending temperature should be larger for the more spherical molecules (2-propanol and 2-butanol). This is because the intermolecular attractions are weaker as the contact area between molecules decrease.

6. ☐ Start a new experiment on the data collection system. �(1.2)

7. ☐ Connect a stainless steel temperature sensor to the data collection system. �(2.1)

8. ☐ Display Temperature (°C) versus Time (s) in a graph. �(7.1.1)

Collect Data

9. ☐ Remove the stopper from the test tube containing methanol, and place the temperature sensor in the methanol.

10. ☐ While viewing the graph display, start recording data. �(6.2)

11. ☐ Remove the temperature sensor from the methanol and tape it such that the metal section is hanging over the edge of the lab table. Replace the stopper on the test tube of methanol.

12. ☐ Adjust the scale of the axes as necessary to see the temperature change. �(7.1.2)

13. ☐ What is happening to the liquid on the temperature sensor? How is the temperature affected?

The liquid is evaporating which causes a decrease in temperature.

14. ☐ How can the evaporation rate be determined from the graph?

The evaporation rate is the change in temperature over the change in time, which is given by the slope of the graph.

15. ☐ Stop recording data when the temperature begins to increase. �(6.2)

16. ☐ Name the data run “methanol”. �(8.2)

17. ☐ Rinse the temperature sensor several times with clean water, and then dry it completely.

18. ☐ Why is it necessary to clean and dry the temperature sensor after each trial?

You must clean the temperature sensor to remove any alcohol remaining on the temperature sensor. The temperature sensor must be completely dry because any remaining water will contaminate the next sample and cause a change in the temperature readings that is not due to the pure alcohol being tested.

19. ☐ Repeat the "Collect Data" steps for each of the remaining alcohols. Name each data run according to the alcohol being tested. �(8.2)

20. ☐ Save your experiment and clean up the lab station according to the teacher’s instructions, especially those concerning your excess alcohols. �(11.1)

Data Analysis

1. ☐ Determine the initial temperature, final lowest temperature, and change in temperature for each alcohol in the homologous series. Record the values in Table 1 below. �(9.2)

Table 1: Temperature changes in the homologous series of alcohols

Alcohol	Initial Temperature (°C)	Final Lowest Temperature (°C)	Change in Temperature (°C)
Methanol	24.1	11.8	–12.3
Ethanol	24.4	15.1	–9.3
Propanol	22.8	18.3	–4.5
Butanol	22.6	20.7	–1.9
Pentanol	23.8	22.1	–1.7

2. ☐ Determine the initial temperature, final lowest temperature, and change in temperature for each isomeric alcohol pair. Record the values in Table 2 below. �(9.2)

Table 2: Temperature changes for isomeric pairs of alcohols

Alcohol	Initial Temperature (°C)	Final Lowest Temperature (°C)	Change in Temperature (°C)
Propanol	22.8	18.3	–4.5
2-Propanol	20.3	14.7	–5.6
Butanol	22.6	20.7	–1.9
2-Butanol	24.1	20.3	–3.8

3. ☐ Determine the evaporation rate for the first 20 seconds of each trial in the homologous series and record the values in Table 3 below.

a. Display the run of data you want to analyze. �(7.1.7)

b. Select the first 20 seconds of decreasing temperature data. �(7.1.4)

c. Apply a linear fit and determine the slope of the linear fit line. �(9.6)

Table 3: Evaporation rate for alcohols in a homologous series

Alcohol	Equation of Linear Fit For the First 20 Seconds	Evaporation Rate (°C/s)
Methanol	y = –0.560x + 23.7	–0.560
Ethanol	y = –0.358x + 24.3	–0.358
Propanol	y = –0.085x + 23.0	–0.085
Butanol	y = –0.014x + 22.6	–0.014
Pentanol	y = –0.008x + 23.8	–0.008

4. ☐ Determine the evaporation rate for the first 20 seconds of each trial for the isomeric alcohol pairs. Record the values in Table 4 below.

a. Display the run of data you want to analyze. �(7.1.7)

b. Select the first 20 seconds of decreasing temperature data. �(7.1.4)

c. Apply a linear fit and determine the slope of the linear fit line. �(9.6)

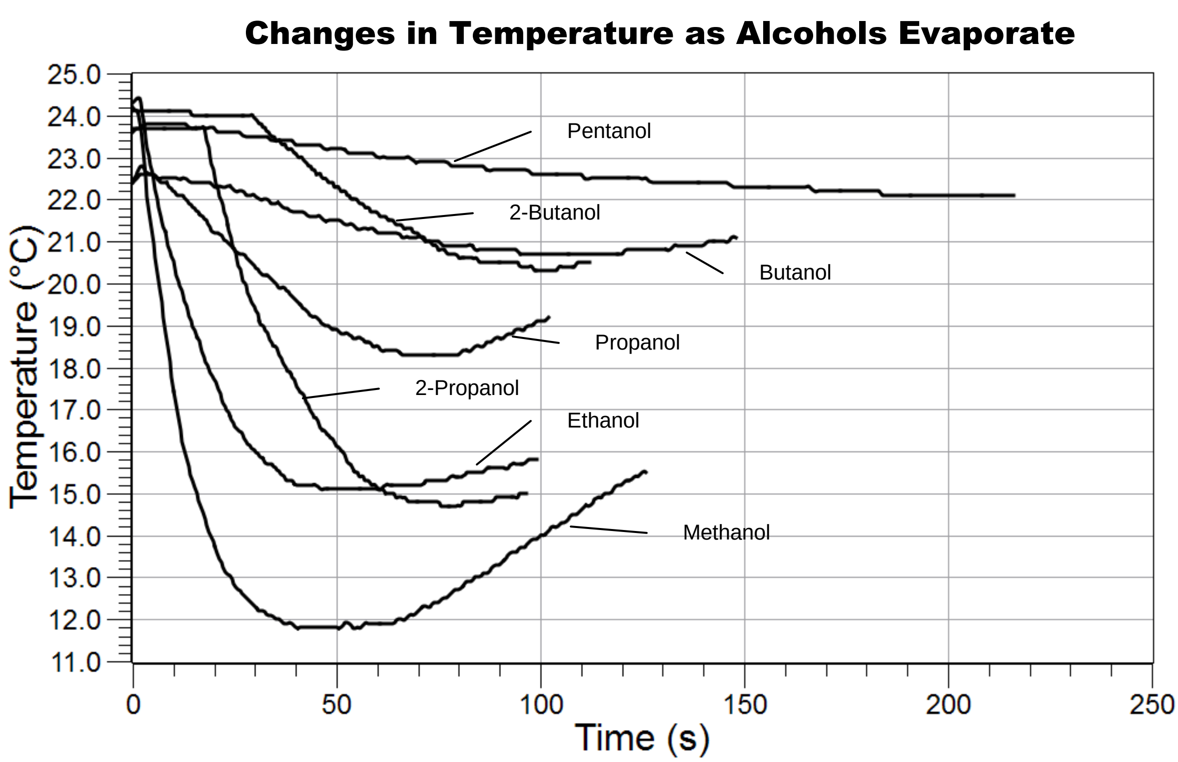
Table 4: Evaporation rates for isomeric pairs of alcohols

Alcohol	Equation of Linear Fit For the First 20 Seconds	Evaporation Rate (°C/s)
Propanol	y = –0.085x + 23.0	–0.085
2-Propanol	y = –0.279x + 29.4	–0.279
Butanol	y = –0.014x + 22.6	–0.014
2-Butanol	y = –0.116x + 23.5	–0.116

5. ☐ Create a graph with all seven runs of data displayed on your data collection system. �(7.1.3)

Note: Not all data collection systems will display all seven runs of data on one set of axes. If this is not possible, you may decide to display all the homologous alcohols on one set of axes and the isomeric pairs on a different set.

6. ☐ Sketch or print a graph of Temperature (°C) versus Time (s) for all seven alcohols on one set of axes. Be sure to label each alcohol. Also label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Analysis Questions

1. How does evaporation affect temperature? Explain.

Evaporation causes the temperature of the substance to decrease. This is because the molecules in the liquid with the highest kinetic energy are able to escape from the intermolecular attractions and become a vapor. When the particles with the highest energy are removed, the average kinetic energy of the remaining molecules decreases.

2. Explain how the magnitude of the evaporation rate changed with the size of the molecules in the homologous series?

The magnitude of the evaporation rate decreased as the size of the molecules increased.

3. Explain how the magnitude of the evaporation rate changed with the shape of the molecules in the isomeric alcohol pairs?

The magnitude of the evaporation rate increased as the molecule became more spherical in shape.

4. Which evaporated liquid from the experiment has the strongest intermolecular forces? How does your data support your answer?

Pentanol has the strongest intermolecular forces. Pentanol had the smallest change in temperature and the smallest evaporation rate, which means that the molecules are held together more tightly than in the other alcohols.

5. Which evaporated liquid from the experiment has the weakest intermolecular forces? How does your data support your answer?

Methanol is the liquid with the weakest intermolecular forces. Methanol had the largest change in temperature and the highest evaporation rate, which means that the molecules are weakly held together.

6. Explain the effects of molecular size on the strength of intermolecular forces for different alcohols from the same homologous series.

Larger molecules have stronger intermolecular forces of attraction.

7. Explain the effects of molecular shape on the strength of intermolecular forces for different isomeric alcohol pairs.

Spherical molecules have weaker intermolecular forces of attraction.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Vigorous exercise causes people to sweat. How does perspiration regulate body temperature?

Perspiration cools our body down. When we sweat, water evaporates from our skin. The water molecules with the most kinetic energy escape and become gas molecules. The energy used to overcome the intermolecular forces between the liquid water molecules comes from body heat, thus drawing heat and cooling the internal temperature.

2. If you enclosed methanol in one container and butanol in a second container, how would the pressures in those containers compare? Why?

The pressure in the container containing methanol would be greater, because methanol evaporates more easily (has a higher vapor pressure) than butanol.

3. Which do you expect to have the higher boiling point, butanol or 2-butanol? Why?

Butanol will have the higher boiling point, because it has more surface contact area resulting in stronger attractions between molecules than 2-butanol. This means more heat energy (a higher temperature) is required to separate the butanol molecules.

4. Would you expect water (H_2O) or hydrogen sulfide (H_2S) to have stronger intermolecular attractions? Explain your reasoning.

You would expect water to have stronger intermolecular attractions. Both water and hydrogen sulfide are polar molecules of similar size and shape. However, the water molecule is capable of forming very strong hydrogen bonds, whereas hydrogen sulfide only experiences the weaker dipole-dipole interactions. This can be seen by comparing the boiling points for the two liquids (water = 100 °C; hydrogen sulfide = –60.3 °C).

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Which alcohol will evaporate the fastest?

- **A. Methanol**
- **B.** Ethanol
- **C.** Propanol
- **D.** Butanol

2. Which of the following are the strongest intermolecular forces found in propanol?

- **A.** Dipole-dipole
- **B.** Dispersion (London)
- **C.** Ionic bonding
- **D. Hydrogen bonding**

3. How does the size of an alcohol affect the strength of its intermolecular forces?

- **A.** As the size of the alcohol decreases, the strength of intermolecular forces increases
- **B. As the size of the alcohol decreases, the strength of intermolecular forces decreases**
- **C.** As the size of the alcohol increases, the strength of its intermolecular forces decreases
- **D.** The size of the alcohol does not have an effect on its intermolecular forces

4. As a liquid evaporates, the temperature of the remaining liquid will

- **A. Decrease**
- **B.** Increase
- **C.** Stay the same
- **D.** Increase or decrease depending on the liquid

5. Which of the following substances has the weakest intermolecular forces of attraction?

- **A.** H_2O
- **B.** $\mathbf{Cl_2}$
- **C.** $C_4H_{10}O$
- **D.** NH_3

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Physical properties, such as the state of matter, evaporation rate, and boiling points, can all be explained by intermolecular forces. The attraction that holds molecules together in the liquid and solid states are called **intermolecular** forces. There are several types of intermolecular forces. **Dispersion (London)** forces occur between non-polar molecules and are the **weakest** type of intermolecular forces. **Dipole-dipole** interactions occur between polar molecules, because the positive end of one molecule is attracted to the negative end of another. **Hydrogen bonding** is the strongest type of intermolecular force and occurs between molecules that contain hydrogen and either fluorine, **oxygen**, or nitrogen. This attraction is strong because of the large difference in **electronegativity** between these atoms.

2. The process of changing from a liquid to a gas is called **evaporation**. The rate at which evaporation occurs depends on the **strength** of the intermolecular attractions holding the particles together. Liquids with strong intermolecular attractions evaporate **slowly**, while liquids with weak intermolecular attractions evaporate **quickly**. In homologous series, such as primary alcohols, the evaporation rate **decreases** as the size of the molecule increases. **Methanol** evaporates quickly, because it has weak intermolecular forces. **Propanol**, on the other hand, evaporates at a much slower rate, because the intermolecular forces are stronger. In all cases, evaporation causes a decrease in the **temperature** of the remaining liquid.

Extended Inquiry Suggestions

Determine the effects of initial temperature on the rate of evaporation.

Investigate the relationship between evaporation rate and boiling point.

Determine the effect salt concentration has on the evaporation rate of water.

14. Concentration of a Solution: Beer's Law

Objectives

Determine the concentration of a copper(II) sulfate solution using a colorimeter. Through this investigation, students:

- Understand the difference between absorbance and transmittance of light
- Become aware of the effects the three variables (chemical substance, path length, and concentration) have on the absorption of light
- Determine the relationship between absorbance of light through a solution and the concentration of the solution (Beer's law)

Procedural Overview

Students conduct the following procedures:

- Dilute a copper(II) sulfate solution of known concentration to create five calibration standards
- Graph the absorbance of orange (610 nm) light against concentration
- Use linear regression to determine the line of best fit and, subsequently, determine the concentration of an unknown copper(II) sulfate solution

Time requirement

♦ Preparation time	20 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	45 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Colorimeter
- Sensor extension cable
- Glass cuvette with cap (7)
- Beaker (2), 100-mL
- Test tube (6), 20-mm x 150-mm
- Test tube rack
- Volumetric pipet with a bulb or a pump (2), 10-mL
- Non-abrasive cleaning tissue
- 0.80 M Copper(II) sulfate ($CuSO_4$), 30 mL[1]
- Unknown Copper(II) sulfate ($CuSO_4$), 10 mL[2]
- Distilled (deionized) water, 30 mL

[1]To formulate using anhydrous copper(II) sulfate ($CuSO_4$) or copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$), refer to the Lab Preparation section.

[2]Refer to the Lab Preparation section for instructions on how to make the unknown solution.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Molarity
- Linear regression line of best fit
- Ionic nomenclature

Related Labs in This Guide

Labs conceptually related to this one include:

- Properties of Ionic and Covalent Compounds
- Electrolyte versus Non-Electrolyte Solutions
- Gay-Lussac's Law and Absolute Zero
- Single Replacement Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

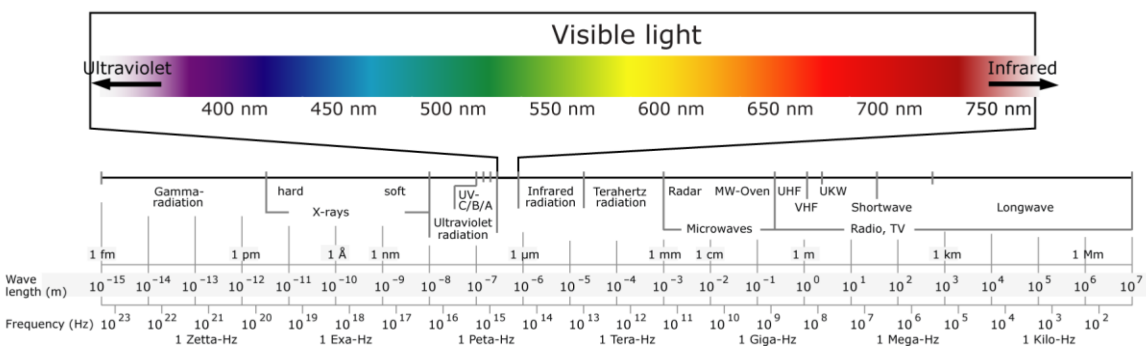
- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Manual sampling mode with manually entered data �(5.2.1)
- Monitor live data in a digits display �(6.1)
- Starting a manually sampled data set �(6.3.1)
- Recording a manually sampled data point �(6.3.2)
- Stopping a manually sampled data set �(6.3.3)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)

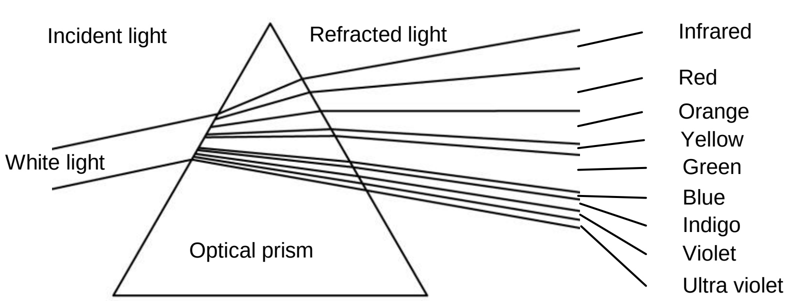
- Showing and hiding connecting lines between data points � (7.1.8)
- Applying a curve fit � (9.5)
- Creating calculated data � (10.3)
- Saving your experiment � (11.1)
- Printing � (11.2)

Background

Visible light is only one small region of the entire electromagnetic spectrum. Visible light is made up of a spectrum of colors ranging approximately between 400 and 700 nm. Combining the entire set of wavelengths in the visible spectrum together produces “white” light. Sunlight is white light; a prism is able to separate and spread the wavelengths so that we may see the various colors that are present. A rainbow is the result of water droplets in the air acting as prisms.

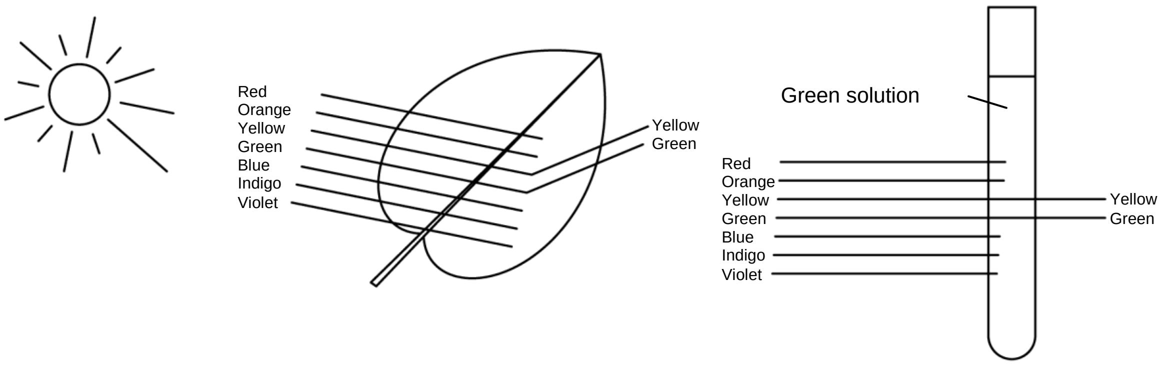
Electromagnetic Spectrum





A prism separating white light into its different colors (wavelengths) of light.

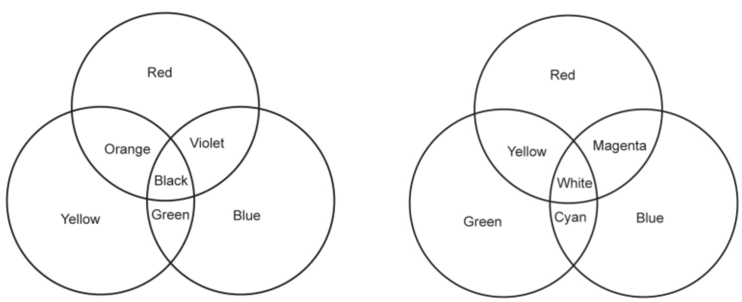
Light is able to interact with matter in three main ways: absorbed, reflected (scattered), or passed through (transmitted) without interaction. The color of an opaque object comes from the light that is reflected from its surface, while the color of a transparent object comes from the wavelengths of light that are transmitted through it. Colored objects always absorb some portion of white light; it is the reflected wavelengths that were not absorbed that produce the observed color of the object.



Yellow and green light are reflected from a green leaf.

Yellow and green light are transmitted through a green solution.

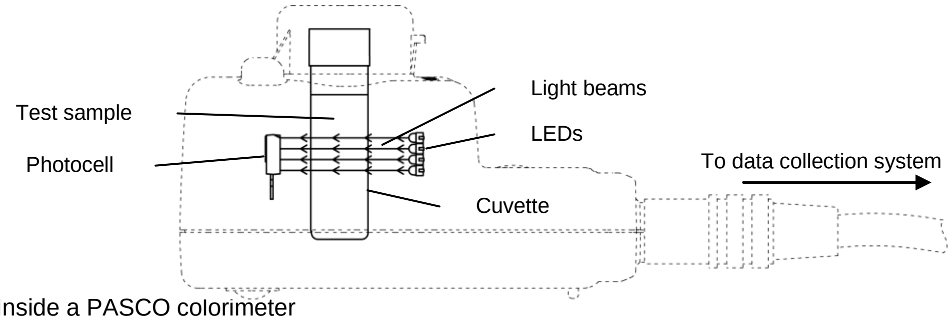
The wavelengths of light mix to produce a single observed color much the same way that different paint pigments mix to produce a new color. While the primary colors in art are red, blue, and yellow (which mix to form black), the primary colors in light are red, blue, and green (which mix to produce white light). Black objects absorb all visible wavelengths of light, and white objects reflect all visible wavelengths of light.



Subtractive (pigment) colors

Additive colors

The human eye is very sensitive to variations in absorbed and transmitted light. A colorimeter is an electronic “eye” that quantitatively measures the absorbance and transmittance of light through solutions. The PASCO colorimeter measures the absorbance and transmittance of four specific wavelengths of light: red (660 nm), orange (610 nm), green (565 nm), and blue (468 nm).



Inside a PASCO colorimeter

The transmittance of light is reported as a percentage of the light that passed through the sample without being absorbed. Transmittance is logarithmic and requires a transformation in order to be useful in the Beer's law equation; a more convenient method is to measure the absorption of light (absorbance). Absorbance is a unitless value usually measured on a scale of 0 to 3 and can either be read directly from the colorimeter or calculated from transmittance data using the formula given below.

$$\text{absorbance} = -\log\left(\frac{\%\ \text{transmittance}}{100}\right)$$

The particular wavelength that is chosen to measure absorbance or transmittance is based on the absorption properties of the solution. Any wavelengths that are either fully transmitted or fully absorbed in the concentration range of the solutions being tested cannot be used. The ideal wavelength is one that shows the strongest absorption while still allowing some light to pass through the samples (transmit) even at the highest concentrations being investigated.

Beer's law (also known as the Beer-Lambert law or the Beer-Lambert-Bouguer law) is an empirical relationship between the amount of light absorbed (the absorbance) and properties of the solution the light passes through. It was independently developed in various forms by Pierre Bouguer (for atmospheric applications) in 1729, Johann Heinrich Lambert (for absorbing media in general) in 1760, and August Beer (for solutions) in 1852.

The relationship that Beer finalized may be expressed by the mathematical formula:

$$A = \varepsilon \times b \times c$$

where A is the absorbance of light, ε (epsilon) is the absorption coefficient (molar absorptivity) of the substance, b is the distance the light travels through the sample (path length), and c is the concentration of the solution. The value of the absorption coefficient varies from one absorbing substance to the next and also with the selected wavelength; it is determined experimentally.

In this experiment, a series of aqueous copper(II) sulfate solutions will be analyzed. Since the same substance is being used, the absorption coefficient will be constant. The path length will also be constant, because the colorimeter uses only one size of cuvette in its sample chamber. By combining these constants into a single constant, Beer's law is simplified.

$$A = \text{constant} \times c$$

A graph of absorbance versus concentration will give a straight line. In this experiment, the students will create a series of copper(II) sulfate solutions with known concentrations. The absorbance of these solutions will be measured and plotted against their known concentrations to create a calibration curve. A best fit line using linear regression will be found, and the equation of this line will be used to determine the concentration of a copper(II) sulfate solution with an unknown concentration.

Pre-Lab Discussion and Activity

Color: Interactions Between Light and Matter

Engage the students in a discussion about color. Use a prism to split white light into a visible spectrum of light. Review that different wavelengths of visible light have different colors and energy. Have the students describe the different ways in which light and matter interact to provide us with a diverse spectrum of color. Hold up different colored objects and ask the students why they are different colors. Differentiate between absorbed, reflected, and transmitted light using opaque and transparent objects.

Teacher Tip: To help students conceptualize the idea of how we perceive color, use three flash lights each covered with a different filter of red, green, or blue. In a darkened room, project various combinations of the beams onto a white, reflective surface, such as a dry erase ink board or a large piece of paper. Place beakers of water with various food colorings added between the flashlights and the reflective surface. Additionally, place larger pieces of the same filters on an overhead projector and hold objects of different colors (various crayons or colored markers work very well) in front of the filtered light. Plastic cling-wraps of various colors are now available in local grocery stores and make suitable, inexpensive filters.

1. Why do different objects have different colors? What is needed for us to see color?

Color is a physical property of matter. Matter and light interact and the resulting light is received by our eyes. Different objects have different colors because they are made up of different substances.

2. What colors make up sunlight (white light)? How can white light be separated into its individual colors?

Sunlight is made up of the entire visible spectrum of light. White light can be separated by using a prism.

3. List the ways light can interact with matter.

Light can be transmitted through, reflected (scattered) by, or absorbed into matter.

4. How does the interaction of light and matter differ between an opaque white object and an opaque black object?

A black object absorbs all the wavelengths of visible light, and a white object reflects all the wavelengths of visible light.

5. How does the interaction of light and matter differ between an opaque green object and an opaque yellow object? How are the interactions similar?

The green object is reflecting a mixture of wavelengths of light that produce the ‘green’ color and absorbing the rest. The yellow object is reflecting a mixture of wavelengths of light that produce the ‘yellow’ color and absorbing the rest. Both green and yellow objects absorb some wavelengths of light and reflect others.

6. How does water (a transparent form of matter) interact with light differently than a blue aqueous solution? How are the interactions similar?

Most wavelengths of light pass (are transmitted) through water. In a blue, aqueous solution, some wavelengths of light are absorbed by the substance and the rest of the colors pass through. The remaining, transmitted wavelengths combine to produce a blue color.

Variables Affecting Color and Beer's Law

Use several simple demonstrations to help the students realize there are three qualities that influence how color is perceived: the chemical substance, its concentration, and the path length of the light through the sample.

Begin by showing the students that various chemical substances cause the color of solutions to be different. Do this by diluting different food coloring in separate beakers of water.

Next, show that concentration is a factor by transferring 2 to 3 mL of a concentrated colored solution in a 1000-mL beaker (labeled "B") to another 1000-mL beaker (labeled "A") containing clean water. Explain that both beakers contain the same chemical substance, but one is more concentrated than the other. Review the terms "dilute" and "concentrated," and have the students evaluate the effect of concentration on the amount of light absorbed.

To demonstrate the affects of path length, use a 1-mL pipet to draw some of the colored solution from the beaker containing the concentrated solution of food coloring used previously. Have the students compare the color of the solution in the pipet with that in the beaker. Point out to the students that both solutions are the same chemical substance and have the same concentration. Finally, combine all three observations of the different variables together to introduce Beer's law.

7. Why can aqueous solutions be different colors?

Different substances can be dissolved in water, which have different physical properties and absorb different wavelengths of light.

8. Is the solution in beaker A or beaker B darker? Why?

Beaker B contains the darker solution. This is because it has more of the chemical substance (is more concentrated) than the solution in beaker A. By having more of the chemical substance in solution, it is able to absorb more of the light. Less light is transmitted to the eyes, so it appears darker.

9. What do the terms "concentrated" and "dilute" mean when they are used to describe solutions?

A concentrated solution has a large amount of solute dissolved in the solvent, while a dilute solution has a very small amount of solute dissolved in the solvent.

10. How does the concentration of a solute affect the amount of light the solution absorbs?

The concentration of solute and the amount of light the solution absorbs are directly proportional. As the concentration of solute increases, the amount of light absorbed also increases.

11. If 1 mL of the colored solution in beaker B is put into a pipet, will the solution in the pipet have the same color or will it be lighter or darker than the solution in beaker B?

The solution in the pipet will have the same color, but it will be lighter.

12. Is the solution in the pipet the same as, darker than, or lighter than the solution in beaker B? Explain.

The solution in the pipet is lighter than the solution in beaker B. This is because the width of the pipet is small and contains less substance for the light to interact with. The beaker, on the other hand, is much wider and the light has more opportunities to interact with the absorbing particles. The distance through which the light passes is called the "path length."

13. What are the three variables that affect the appearance of a solution?

The solute that makes up the solution, its concentration, and the width of the container holding the solution (path length) all affect the way the eyes perceive the solution.

14. State Beer's law.

Beer's law states that absorbance A is equal to the product of the absorption coefficient (molar absorptivity) of the substance ε, the distance light travels through the material (the path length b), and the concentration of the solution c.

$A = \varepsilon \times b \times c$

Colorimeter as a Tool to Determine Concentration

Hold up a colorimeter for the students to see, and explain that a colorimeter is an electronic measuring device that measures the absorbance or transmittance of light. The colorimeter is similar to your eyes in that it can detect subtle differences in the color of a solution; however, unlike your eyes, the colorimeter is also able to assign a quantitative (numerical) and reproducible value to the absorbance and transmittance of light. A cuvette filled with a solution is placed in the colorimeter. A light passes through the solution, and the absorbance or transmittance of four wavelengths of light—red (660 nm), orange (610 nm), green (565 nm), or blue (468 nm)—are measured.

Note: The colorimeter must first be calibrated with a cuvette filled with the solvent being used (usually water), and is called a "blank." Failure to calibrate with a blank will produce a systematic error in the results, because water does absorb some light. This will cause each absorbance reading to be slightly larger than it should be.

Demonstrate how to calibrate the colorimeter. After calibration, hold up a cuvette that contains a 0.8 M copper(II) sulfate solution. Have the students predict which wavelengths of light are absorbed and which colors are transmitted. Project a digits display so the students can see the transmittance and then absorbance of each wavelength of light. Guide the students into understanding why the orange (610 nm) wavelength of light is used in the experiment. End the discussion by helping the students understand how the concentration of an unknown solution can be determined.

15. Which variable affecting the appearance of a solution is held constant by the design of the colorimeter?

The path length of light traveling through the sample is held constant. This is always the same, because only one size of cuvette fits into the colorimeter's sample chamber.

16. If the colorimeter is used with the same chemical substance, how is Beer's law simplified? Explain.

Since the absorption coefficient (molar absorptivity ε) and the path length b are exactly the same for each solution being tested, then $\varepsilon \times b$ will be a constant.

Beer's law can be simplified to, absorbance = constant × concentration

17. What color/wavelength of light do you think is being transmitted? How do you know?

Students should predict that blue light is being transmitted; they may or may not predict that some of the other colors are also transmitted. Blue light is being transmitted the strongest, because it is the predominant color of the solution.

18. Using the proper units, report the transmittance of each of the four colors/wavelengths of light passing through the 0.8 M copper(II) sulfate solution.

Sample data for 0.8 M copper(II) sulfate:

Red (660 nm) = 0.0%

Orange (610 nm) = 6.5%

Green (565 nm) = 37.6%

Blue (468 nm) = 100.0%

19. What color/wavelength of light do you think is being absorbed by the 0.8 M copper(II) sulfate solution? Why?

Red light is being absorbed, because it is not transmitted (0.0%). Green and orange light are partially absorbed, but blue light is not being absorbed at all (100.0%).

20. How is absorbance reported? What are the absorbance values for each of the four colors/wavelengths of light passing through the 0.8 M copper(II) sulfate solution?

Absorbance values are unitless and are reported on a scale of 0 to 3. Sample data for 0.8 M copper(II) sulfate:

Red (660 nm) = 3.000

Orange (610 nm) = 1.190

Green (565 nm) = 0.424

Blue (468 nm) = 0.000

21. The most concentrated solution you will test in this experiment is 0.8 M. Which color/wavelength of light would be best to use? Which other color(s) could be used? Which color(s) cannot be used?

Orange (610 nm) light is the best to use, because it has a strong absorbance, but still allows some light to be transmitted (is not 0.0%); green (565 nm) light could also be used. Blue (468 nm) light cannot be used, because it is not absorbed at all (0.0 absorbance, 100% transmittance). Red light (660 nm) should not be used, because its absorption is too great at the higher concentrations of copper(II) sulfate (3.0 absorbance, 0.0% transmittance); it may be suitable for lower concentrations, but its effective range would need to be determined first.

22. How will you be able to determine the concentration of an unknown copper(II) sulfate solution?

Using orange (610 nm) light, create a calibration curve that gives the absorbance of known concentrations of copper(II) sulfate. Find the equation of the best fit line ($y = mx + b$). Once the equation of the linear fit is known, the absorbance of the unknown copper(II) sulfate solution can be measured and substituted into the equation. The concentration can then be calculated.

23. The equation of a line is $y = mx + b$. Match each of the components of a line with the variables found in Beer's law.

Recall the equation for Beer's law:

absorbance = absorption coefficient × path length × concentration .

This was simplified by using only one chemical substance (constant absorption coefficient) and one size of cuvette (constant path length), allowing these two variables to be combined into one constant. Now:

absorbance = constant × concentration.

Matching this to the equation for a line gives:

y = absorbance, x = concentration

m = constant (ε × b, absorption coefficient × path length), it will be the slope of the linear fit line

b = should be zero, it is the y-intercept

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. Prepare 500 mL of 0.8 M copper(II) sulfate ($CuSO_4$) solution from either anhydrous copper(II) sulfate ($CuSO_4$) or copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$). This is enough solution for *15* lab groups.

Starting with anhydrous copper(II) sulfate ($CuSO_4$)

- **a.** Add approximately 200 mL of distilled water to a 400-mL beaker.
- **b.** Add 63.84 g anhydrous $CuSO_4$ to the water and heat with constant stirring until the $CuSO_4$ is dissolved.
- **c.** Allow the solution to cool.
- **d.** Transfer the solution to a 500-mL volumetric flask and dilute it to the 500 mL mark with distilled water.
- **e.** Cap and invert the flask at least three times to ensure complete mixing.

Starting with copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$)

- **a.** Add approximately 200 mL of distilled water to a 400-mL beaker.
- **b.** Add 99.87 g $CuSO_4 \cdot 5H_2O$ to the water and heat with constant stirring until the $CuSO_4$ is dissolved.
- **c.** Allow the solution to cool.
- **d.** Transfer the solution to a 500-mL volumetric flask and dilute it to the 500 mL mark with distilled water.
- **e.** Cap and invert the flask at least three times to ensure complete mixing.

2. For the unknown solution, you may create any copper(II) sulfate ($CuSO_4$) concentration you wish that is 0.8 M or less. The following equations can be used to determine the number of milliliters of the prepared 0.8 M copper(II) sulfate solution required to create 100 mL of solution at a desired concentration. This is enough for *10* lab groups.

$$\frac{(\text{desired molarity})(100\text{ mL})}{0.8\text{ M}} = \text{mL required}$$

a. Using a graduated pipet, transfer the appropriate volume of 0.8 M copper(II) sulfate solution to a 100-mL volumetric flask.

b. Dilute to the 100 mL mark with distilled water.

c. Cap and invert the flask at least three times to ensure complete mixing.

Teacher Tips:

- You may decide to prepare several unknowns. Students may be asked to find more than one unknown, or you can have different groups work with different unknowns.
- It is extremely important that students be meticulous when performing the dilutions. Careful attention to the preparation of the calibration standards will produce extremely accurate results.
- The copper(II) sulfate waste generated in this experiment must be collected and disposed of according to your local, state, and federal regulations. We suggest that a large container be provided in which all students can dispose of their waste solutions. At the completion of the experiments, the waste should be processed following the procedures outlined by your organization.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Copper(II) sulfate is hazardous to the environment and should not be disposed of down the drain. Make sure you follow your teacher's instruction to properly dispose of the copper(II) sulfate solutions.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

3	1	5	2	4
Create a calibration curve by graphing absorbance versus concentration.	Create five copper(II) sulfate solutions of known concentration.	Measure the absorbance of the unknown and determine its concentration.	Measure the absorbance of each known copper(II) sulfate solution using a calibrated colorimeter.	Determine the linear equation for the calibration curve's line of best fit.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Part 1 – Calibration Curve

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect the colorimeter to the data collection system using a sensor extension cable. �(2.1)

3. ☐ Calibrate the colorimeter.

- **a.** Fill a cuvette with distilled water and cap it.
- **b.** Holding the cuvette by the lid, wipe the outside of the cuvette with a non-abrasive cleaning tissue.
- **c.** Place the cuvette in the colorimeter and close the lid.
- **d.** Push the green calibrate button on the colorimeter.
- **e.** When the green light turns off, the calibration is complete and the cuvette can be removed from the colorimeter.

4. ☐ Why is it necessary to calibrate the colorimeter using distilled water?

Distilled water is the solvent, and it absorbs some light. Failure to calibrate with a blank will cause each absorbance reading to be slightly larger than it should be.

5. ☐ Configure the data collection system to manually collect absorbance of orange (610 nm) light and concentration in a table. Define concentration as a manually entered data set with units of molarity (M). �(5.2.1)

6. ☐ Why is the absorbance of orange (610 nm) light being used instead of another color?

Orange (610 nm) light is used, because it is partially absorbed and partially transmitted.

7. ☐ List the dependent and independent variables used in this experiment. Also, list the units for each.

The dependent variable is the absorbance of orange (610 nm) light. Absorbance values are unitless.

The independent variable is the concentration of the solution and is measured in units of molarity.

8. ☐ Measure approximately 30 mL of 0.80 M copper(II) sulfate ($CuSO_4$) stock solution into a 100-mL beaker.

9. ☐ Measure approximately 30 mL of distilled water into a different 100-mL beaker.

10. ☐ Take six clean, dry test tubes and place them in a test tube rack. Label the test tubes "1", “2”, “3”, "4", "stock", and "unknown”.

11. ☐ Why do the test tubes need to be dry? What error would be caused by wet test tubes?

Wet test tubes would change the concentration (molarity) of the standard solutions by diluting them by various, unknown amounts which would produce an inaccurate calibration curve.

12. ☐ Prepare the five standard copper(II) sulfate ($CuSO_4$) solutions listed in Table 1 below.

Table 1: Copper(II) sulfate solution concentrations

Test Tube	0.80 M $CuSO_4$ (mL)	H_2O (mL)	Concentration (M)
1	2.0	8.0	0.16
2	4.0	6.0	0.32
3	6.0	4.0	0.48
4	8.0	2.0	0.64
Stock	~10	0	0.80

a. Use a 10-mL volumetric pipet to deliver 2.0, 4.0, 6.0, and 8.0 mL of the 0.80 M copper(II) sulfate solution into test tubes 1 through 4, respectively.

b. Use a different 10-mL volumetric pipet to deliver 8.0, 6.0, 4.0, and 2.0 mL of distilled water into test tubes 1 through 4, such that each test tube has a total of 10.0 mL of solution.

c. Thoroughly mix each solution by swirling each test tube.

d. Pour the remaining 0.80 M copper(II) sulfate solution from the beaker into the "stock" test tube to use as the fifth data point in your calibration curve.

13. ☐ Fill one cuvette with the 0.16 M $CuSO_4$ solution and cap it. Label the top of the cuvette lid.

14. ☐ Continue to fill one cuvette at a time until each solution is in a cuvette, the cap is on, and it is clearly labeled (0.32 M, 0.48 M, 0.64 M, and 0.80 M).

Collect Data

15. ☐ While viewing the table display, start a new manually sampled data set. �(6.3.1)

16. ☐ Use a non-abrasive cleaning tissue to wipe the outside of the cuvette containing 0.16 M $CuSO_4$, and then place the cuvette inside the colorimeter. Close the lid of the colorimeter.

17. ☐ Why is it necessary to wipe the outside of the cuvette before you place it in the colorimeter?

It is necessary to wipe the cuvette because any substances on the outside of the cuvette, including finger prints, will absorb (or scatter) light and cause the absorbance values to be lower than the actual values.

18. ☐ Why is it necessary to close the lid of the colorimeter before recording the data values?

The colorimeter measures the absorbance and transmittance of light. If the lid is open, the light from the room will interfere with the reading, causing the absorbance to be lower than the actual value.

19. ☐ Record the absorbance and enter the concentration into the table on the data collection system. �(6.3.2)

20. ☐ Remove the cuvette from the colorimeter.

21. ☐ Repeat this process for each of the other standard solutions. For each solution, make sure to wipe the outside of each cuvette before placing it in the colorimeter, record the absorbance, and enter the concentration of the solution into the data table on the data collection system. �(6.3.2)

22. ☐ When you have collected all of your data, stop the data set. �(6.3.3)

23. ☐ Copy the absorbance and concentration data collected from your data collection system to Table 2 in the Data Analysis section.

Analyze Data

24. ☐ Create a calibration curve using the absorbance and concentration data you just collected.

a. Create a graph of Orange (610 nm) Absorbance versus Concentration (M). �(7.1.1)

Note: To graph a scatter plot of the data points you may choose to hide the connecting lines between data points. �(7.1.8) Adjust the scale of the graph as needed. �(7.1.2)

b. Apply a linear fit to the data. �(9.5)

25. ☐ What is the equation for the line of best fit?

y = (1.45)x + 0.0118

absorbance = (1.45)(concentration) + 0.0118

26. ☐ Solve the equation for concentration.

$$\text{concentration} = \left(\left[\text{Orange}\left(610\text{ nm}\right)\text{Absorbance}\right] - 0.0118\right) / 1.45$$

Part 2 – Determining an Unknown Concentration

Set Up

27. ☐ Enter the equation you determined for concentration above, into your data collection system's calculator. �(10.3)

Note: When you enter your calculation into your data collection system make sure that the measurement is absorbance and that it is entered exactly like this—[Orange (610 nm) Absorbance].

28. ☐ Monitor Orange (610 nm) Absorbance and Calculated Concentration data in a digits display. �(6.1)

29. ☐ Obtain 10 mL of a copper(II) sulfate solution with an unknown concentration from your teacher. Put this solution in your test tube labeled "unknown".

30. ☐ Fill a dry, clean cuvette with your unknown solution, cap it, and label it "unknown".

Collect Data

31. ☐ Wipe the outside of the cuvette with a non-abrasive cleaning tissue and place the cuvette in the colorimeter.

32. ☐ Close the lid of the colorimeter and record the absorbance and concentration of the unknown solution below:

"Unknown" absorbance: 0.689

"Unknown" concentration (M): 0.467 M

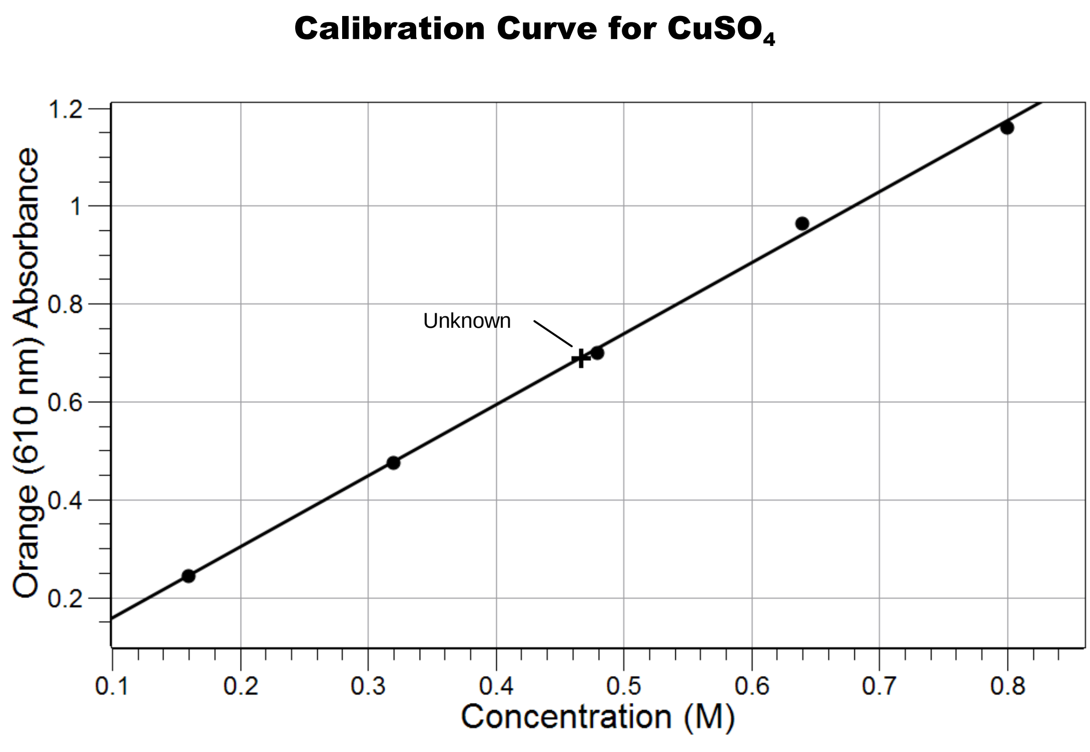
33. ☐ Save your experiment and clean up according to the teacher's instructions, including any special instructions for disposing of the copper(II) sulfate solutions. �(11.1)

Data Analysis

Table 2: Measured orange light absorbance readings

Concentration of $CuSO_4$ Solution (M)	Orange (610 nm) Absorbance
0.16	0.243
0.32	0.474
0.48	0.700
0.64	0.964
0.80	1.158

1. ☐ Plot or print a copy of your calibration curve (Orange (610 nm) Absorbance versus Concentration (M)) including the linear line of best fit. Label the overall graph, x-axis and y-axis, and include units on the axes. �(11.2)



2. ☐ Where does your unknown copper(II) sulfate solution fit on the calibration curve above? Place an "X" on the graph to mark this location and label it "unknown".

Analysis Questions

1. State Beer's law. Does your data support this statement?

Beer's law states that the absorbance of light is directly proportional to the concentration of the solution. Yes, the data supports this because the data fits a straight line (not a curved line) with a constant slope and it goes through the origin (almost).

2. Would an error occur if some of the cuvettes were dirty? Explain.

Yes, the absorbance readings would be higher than the actual values for individual samples because some light would be absorbed (or scattered) by the contaminants.

3. Explain the difference between absorbance and transmittance of light.

Absorbance means the chemical substance interacts with the light and "catches" (absorbs) it. Light that passes through without being trapped is transmitted.

4. Why is $CuSO_4$ a bluish color? Are colors of light other than blue transmitted?

A solution of $CuSO_4$ is blue because of the mixture of different wavelengths of light that are transmitted through it.

Synthesis Questions

Use available resources to help you answer the following questions.

1. A solution of sodium sulfate is clear and colorless, and yet a solution of copper(II) sulfate is blue. Which ions are causing the blue color? Could you use Beer's law to find the concentration of a sodium sulfate solution?

It is the copper(II) ions that are causing the blue color in the solution. Since sodium sulfate is colorless, it will not absorb visible light; because of this, Beer's law cannot be used to find its concentration.

2. What is the minimum number of points needed to create a calibration curve? How many points were used in this experiment? Why?

At least two points are needed to create a line. This experiment used five points to make a calibration curve. The more points that are used, the more accurate the results for the unknowns will be.

3. Other sensors, such as for pH and conductivity, still need to be calibrated, but are able to determine unknown concentrations without creating a calibration curve. Explain how this works.

The calibration curve is stored electronically within the instrument's programming.

Multiple Choice Questions

1. Which of the following variables affects the absorbance of light in a solution?

- **A.** The distance the light has to travel through the solution (path length)
- **B.** The amount of solute in each volume (concentration)
- **C.** The wavelength of the light that is interacting with the solution
- **D. All of the above**

2. A sample 0.10 M copper(II) chloride is placed into a cuvette with a 1.00 cm path length. The solution has a measured absorbance of 2.0. What would you expect the absorbance of a 0.05 M copper(II) chloride solution to be?

- **A. 1.0**
- **B.** 2.0
- **C.** 4.0
- **D.** Not enough information

3. What should be in the 'blank' cuvette when the colorimeter is calibrated?

- **A.** Nothing
- **B.** The solution with the greatest concentration of solute
- **C. The solvent**
- **D.** A 1.0 M sample of the solution

4. What color of light is transmitted through a copper(II) sulfate solution?

- **A.** Red
- **B.** Orange
- **C.** Green
- **D. Blue**

5. How many colors make up white light?

- **A.** 1
- **B.** 3
- **C.** 4
- **D. More than five**

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Visible light is only a small portion of the **electromagnetic** spectrum. When different wavelengths of light combine, they form new colors; when all of the wavelengths in the visible spectrum combine, they produce **white light**. The combined wavelengths can be separated into their individual colors by using a **prism**. Light can be reflected, absorbed, or transmitted when it encounters an object. Black objects **absorb** all visible wavelengths of light, while white objects **reflect** all visible wavelengths of light. Colored objects and solutions absorb only a portion of the wavelengths they encounter. We see the light that is not absorbed by the object or solution. A **colorimeter** measures the absorbance and transmittance of light through solutions. When light is **transmitted**, it passes through the solution; when light interacts with the solute it is **absorbed**.

2. Beer's law is the **directly** proportional relationship between the absorbance of light and the concentration of solute in the solution; increasing the concentration will **increase** the observed absorbance. The relationship depends on the nature of the chemical substance; the concentration of the solution, and the **path length** of the light. Plotting the absorbance values for solutions with **known** concentrations produces a calibration curve. This curve allows for the determination of an **unknown** concentration of the solution.

Extended Inquiry Suggestions

Many water quality studies are done using colorimetric techniques. PASCO offers water quality test kits for use with the water quality colorimeter. The kits allow rapid testing of water for specific ions as the calibration curves are stored in the device. A great extension is to have students analyze a nearby stream or other natural waters taking advantage of the learning from this lab.

Determine the effects of using a different wavelength of light to determine the concentration of copper(II) sulfate.

Repeat the experiment using a different colored solution.

Determine the effect of wavelength on the constant (slope) in the Beer's law equation.

Try using percent transmittance instead of absorbance in Beer's law.

Using packages of different flavored (colored) powdered drink mix, develop and test a method to determine the number of grams of powder in various dilutions of the assorted beverages.

15. pH of Household Chemicals

Objectives

Working with common household chemicals, students develop an understanding of the relationship between pH and the hydronium ion (H_3O^+) concentration. Through this investigation, students:

- Classify solutions as acidic, basic, or neutral based on their pH or hydronium ion (H_3O^+) concentration
- Compare the relative strengths of acids and bases based on their pH or H_3O^+ concentration

Procedural Overview

Students conduct the following procedures:

- Measure the pH of common household chemicals and classify them as acids or bases
- Calculate the hydronium ion (H_3O^+) concentration from the measured pH values and graph pH versus H_3O^+ concentration

Time Requirement

Preparation time	20 minutes
Pre-lab discussion and activity	40 minutes
Lab activity	30 minutes

Materials and Equipment

For each student or group:

- Data collection system
- pH sensor
- Beaker (2), 50-mL
- Graduated cylinder, 50-mL
- Graduated cylinder, 10-mL
- Test tube (10), 15-mm x 100-mm
- Test tube rack
- Wash bottle and waste container
- Buffer solution pH 4, 25 mL
- Buffer solution pH 10, 25 mL
- White vinegar (~5% acetic acid), 5 mL
- Lemon Juice, 5 mL
- Soft drink, 5 mL
- Window cleaner, 5 mL
- Tap water, 5 mL
- Milk, 5 mL
- Coffee, 5 mL
- 0.5 M Sodium bicarbonate (baking soda), 5 mL[1]
- Liquid soap, 5 mL
- Bleach, 5 mL

[1] To formulate using solid sodium bicarbonate (baking soda), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Properties of acids and bases
- Concentration (molarity)
- Calculations involving logarithms

Related Labs in This Guide

Labs conceptually related to this one include:

- An Acid-Base Titration
- Diprotic Titration: Multi-Step Chemical Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Calibrating a pH sensor �(3.6)
- Manual sampling mode with manually entered data �(5.2.1)
- Starting a manually sampled data set �(6.3.1)
- Recording a manually sampled data point �(6.3.2)
- Stopping a manually sampled data set �(6.3.3)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Showing and hiding connecting lines between data points �(7.1.8)
- Adding a measurement to a table �(7.2.2)
- Creating calculated data �(10.3)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

We are surrounded in life by acids and bases. Acids are found in numerous substances including soft drinks, salad dressing, the human body, rain water, and batteries. Acids taste sour, cause indicators to change color (turn litmus red), react with certain metals to form hydrogen gas, and react with bases to form a salt and water.

Bases can be found in soap, cleaning products, baking soda, and beer. Bases taste bitter, feel slippery or soapy, cause indicators to turn color (turn litmus blue), react with oil and grease, and react with acids to form a salt and water.

At the molecular level, an acid is able to donate a hydrogen ion (Brønsted-Lowry acid). The donated hydrogen ion can bond with any available water molecule to form a hydronium ion (H_3O^+). Acids are ranked based on how easily they give up their hydrogen ions. The terms "strong" and "weak" refer not to the concentration of the acid, but the degree to which the acid dissociates (separates into hydrogen ions and the conjugate base).

A strong acid is one that readily and completely dissociates. In a strong acid, every acid unit breaks into a hydrogen ion and the negative counter ion. For example, with hydrochloric acid every unit of HCl splits into H^+ and Cl^-.

In contrast, a weak acid is one that only partially dissociates. In a weak acid only a fraction of the available acid units break into hydrogen ions and the negative counter ions. For example in acetic acid of 100 $HC_2H_3O_2$ units, only one splits into H^+ and $C_2H_3O_2^-$ while the remaining 99 remain combined as $HC_2H_3O_2$.

A base is the complement of an acid. A base is a substance that accepts a hydrogen ion (Brønsted-Lowry base). The more readily a chemical substance bonds with hydrogen ions, the stronger the base.

The pH scale provides a numerical measure of the acid concentration of a chemical substance in solution. The term pH refers to the concentration of H_3O^+ in a solution of the substance. The pH scale is based on the observation that water has a slight tendency to autoionize into H^+ and hydroxide ions (OH^-). Because there is a one-to-one mole ratio, equal amounts of the H_3O^+ ion and the OH^- are formed. At 25 °C, 1.0 x 10^{-7} M of each ion forms.

$$H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$$

$$K_{eq} = K_w = [H_3O^+]\,[OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

The equilibrium constant for this reaction is $K_w = 1.0 \times 10^{-14}$, also called the ion-product constant for water.

The concentration of H_3O^+ ions and OH^- ions is inversely related. If the concentration of H_3O^+ ions increases, then the OH^- ion concentration must decrease and vice versa. The following table shows the two extremes after adding a strong acid or a strong base to water to make a 1.0 M solution.

Table: Concentrations of H_3O^+ and OH^- in 1 M solutions of a strong acid and a strong base

Add a strong acid to create a 1.0 M solution:	**Add strong base to create a 1.0 M solution:**
$\left[H_3O^+\right] = 1.0\ M = 10^0$ $\left[OH^-\right] = \dfrac{K_w}{\left[H_3O^+\right]} = 1 \times 10^{-14}$	$\left[H_3O^+\right] = \dfrac{K_w}{\left[OH^-\right]} = 1 \times 10^{-14}$ $\left[OH^-\right] = 1.0\ M = 10^0$

The range of H_3O^+ ion concentrations will therefore generally be between 1 and 1 x 10^{-14}. Because of the extreme range of values, the pH scale simplifies these numbers by taking the negative logarithm of the H_3O^+ ion concentration to give a range from 0 to 14.

$$pH = -\log [H_3O^+]$$

When there are equal concentrations of H_3O^+ ions and OH^- ions, a solution is said to be neutral and has a pH value of 7. When there are greater amounts of H_3O^+ ions than OH^- ions the solution is acidic and has a pH less than 7. The lower the pH value the more acidic the solution. When there are fewer H_3O^+ ions than OH^- ions, the solution is basic and has a pH greater than 7. The higher the pH value the more alkaline (basic) the solution.

Pre-Lab Discussion and Activity

Discuss acids, bases, and pH with your students using the following activities and questions.

pH and Consumer Products

1. Where have you heard the terms acid, base, or pH used? What type of consumer products use these terms on their labels?

Acids – battery acid, acids in foods, stomach acid, build up of lactic acid during exercise, drugs, acid rain, acidic soil, amino acids for body building.

Bases – soaps, cleaning products, baking soda, antacids, beer.

pH – consumer products such as cosmetics, deodorants, and shampoos are often labeled as “pH balanced.” Also, pH is referred to in swimming pool maintenance, fish tanks, and water quality.

2. What are some properties of acids and bases?

Acids – tart or sour tasting, turn litmus red, react with certain metals to form hydrogen gas, react with bases to form a salt and water.

Bases – bitter tasting, feel slippery or soapy, turn litmus blue, react with oil and grease, react with acids to form a salt and water.

3. What is pH and how is it related to acids and bases?

The pH scale measures the acidity of a substance. Acids have a pH less than 7 and bases have a pH greater than 7.

4. At the molecular level what makes something more or less acidic?

At the molecular level, acidity is a measure of the H_3O^+ ion concentration in a solution. The greater the concentration of H_3O^+ ions, the more acidic the solution.

Conductivity of Acidic, Basic, and Neutral Solutions

Use a conductivity sensor or a light bulb conductivity tester to test the conductivity of a 0.1 M hydrochloric acid (HCl) solution, a 0.1 M sodium hydroxide (NaOH) solution, and distilled water.

Help the students to understand that both acids and bases have a significant amount of ions present, but that distilled water has very few ions. Write the chemical equation for each reaction on the board and help the students analyze the concentration of each ion present.

Write the following table on the board and have students help you fill it in.

Table: Conductivity of 0.1 M HCl, distilled water, and 0.1 M NaOH

	0.1 M HCl	**Distilled H_2O**	**0.1 M NaOH**
Conductivity tester	bright light	no light	bright light
Conductivity sensor (μs/cm)	32,227	15	16,658

5. What do these conductivity readings tell us about hydrochloric acid, sodium hydroxide (base), and distilled water?

Conductivity is a measure of how easily electric current can flow. In order for an electric current to flow, ions (charged particles) must be present in the solution. Good conductors have a greater concentration of ions in the solution. Both hydrochloric acid and sodium hydroxide are good conductors, thus we know their solutions contain many ions. Poor conductors do not contain ions or the concentration of ions is extremely low, therefore distilled water is a poor conductor.

6. What ions are involved in each of the solutions tested? What is the concentration of each ion present?

0.1 M Hydrochloric Acid: $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

Hydrochloric acid completely dissociates to produce 0.1 M of H_3O^+ and 0.1 M Cl^-.

0.1 M Sodium Hydroxide: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

Sodium hydroxide completely dissociates to produce 0.1 M Na^+ and 0.1 M OH^-.

Distilled water: $H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$

The self-ionization of pure water occurs to a very small extent. At 25 °C pure water contains 1.00×10^{-7} M H_3O^+ ions and 1.00×10^{-7} M OH^- ions. Pure water has a very small amount of each ion (0.0000001 M) which explains why it is such a poor conductor.

Hydronium Ion Concentration and the pH Scale

Use a pH sensor to test the pH of 0.1 M HCl, 0.1 M NaOH, and distilled water.

Discuss with students how pH is a measure of the H_3O^+ ion concentration (acidity) of a solution and how a substances ability to donate H^+ ions affects the H_3O^+ ion concentration and pH. Explain that acids are substances that *donate* H^+ ions and bases are substances that *accept* H^+ ions. The addition or removal of H^+ affects the balance H_3O^+ ions and OH^- ions that exist in pure water. The concentrations of H_3O^+ and OH^- are inversely proportional. As the concentration of H_3O^+ increases, the concentration of OH^- decreases. The product of the concentration of H_3O^+ and OH^- is a constant, and is known as the ion-product constant for water.

$$K\text{w} = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

When the concentrations of H_3O^+ and OH^- are equal, the solution is neutral (pH = 7). If the concentration of H_3O^+ is greater than the OH^- concentration, the solution is acidic (pH < 7). If the H_3O^+ concentration is less than the OH^- concentration, the solution is considered alkaline or basic (pH > 7).

Add the following rows to the table you wrote on the board in the previous activity and have the students help you fill it in.

Table: Conductivity, pH, and concentrations of ions in 0.1 M HCl, distilled water, and 0.1 M NaOH

	0.1 M HCl	**Distilled H_2O**	**0.1 M NaOH**
Conductivity tester	bright light	no light	bright light
Conductivity sensor (µs/cm)	32,227	15	16,658
pH	1	7	13
Concentration of H_3O^+ (M)	1×10^{-1}	1×10^{-7}	1×10^{-13}
Concentration of OH^- (M)	1×10^{-13}	1×10^{-7}	1×10^{-1}
Acidic, basic, or neutral? (M)	Acidic	Neutral	Basic

7. How are acids and bases defined in terms of H^+ ions?

Acids are substances that donate H^+ ions and bases are substances that accept H^+ ions.

8. What happens to the concentration of H_3O^+ when an acid is added to water? When a base is added to water?

Acids donate H^+ ions and therefore cause the concentration of H_3O^+ to increase.

Bases accept H^+ ions and therefore cause the concentration of H_3O^+ to decrease.

9. Compare the concentrations of H_3O^+ ions in acids, bases, and neutral solutions to their pH values.

Acidic solutions have larger concentrations of H_3O^+ ions and therefore low pH values. Bases have low concentrations of H_3O^+ ions and therefore high pH values. Neutral solutions have an equal number of H_3O^+ and OH^- ions and therefore have a pH of 7.

Logarithmic Scales and pH

Explain to students that pH is a logarithmic scale and that logarithmic scales are used when there is a wide range of values involved. The logarithmic scale simplifies cumbersome numbers by referring to the exponent to which a base (in this case 10) must be raised to yield the number as opposed to talking about the number itself.

Demonstrate the difference between a logarithmic scale and a linear scale by changing volumes of water as explained below. End by providing the definition of pH as the negative log of the concentration of H_3O^+ ions.

Linear scale demonstration: Start with an empty 100 mL graduated cylinder. Add 10 mL. Record the volume and repeat this 5 times. Explain that linear scales are based on addition and subtraction. Each change involves adding or subtracting the same amount.

Logarithmic scale demonstration: Add 10 mL of water to an empty 100 mL graduated cylinder. This time, multiply 10 mL by 10. Fill the graduated cylinder to the top (100 mL). Multiply 100 mL by 10 and transfer the water to a 1000 mL graduated cylinder and fill it up. Continue the process by having the students suggest larger containers that could be used to hold the increasing amount of water. As you are doing this emphasize that the "log" is the power (exponent) to which 10 is being raised to give the volume of water.

Have the students help you fill out a table as you perform the demonstrations:

Table: Comparison of a linear scale and a logarithmic scale

Number of Changes	Linear Scale	Logarithmic Scale	
	Increase the volume by 10 mL each time	Increase the volume by multiplying by 10.	Log
1	10	10 10^1	1
2	10 + 10 = 20	10 ×10 =100 10^2	2
3	20 + 10 = 30	100 × 10 =1,000 10^3	3
4	30 + 10 = 40	1,000 × 10 =10,000 10^4	4
5	40 + 10 = 50	10,000 × 10 =100,000 10^5	5

10. Why do we have the concept of pH if it is simply the concentration (molarity) of H_3O^+ ions?

The molarity of H_3O^+ ions ranges from 1 x 10^{-1} M to 1 x 10^{-14} M. This is a huge range of numbers and the numbers themselves are cumbersome and difficult to write. The pH scale was implemented to simplify discussions of acids and bases.

11. How is a logarithmic scale different than a typical linear scale? When is logarithmic scale preferred over a linear scale?

A linear scale is based on addition/subtraction. Each step on the scale is the same. A logarithmic scale is based on powers of 10. Each step increases or decreases in size. Logarithmic scales are useful when there is a large range of values being covered.

12. Does a linear scale or a logarithmic scale cover a wider range of values?

A logarithmic scales cover a wider range of values. In the example above the linear scale changed from 10 to 50. The logarithmic scale changed from 10 to 100,000 which is a much larger range.

13. What is pH? Relate the concentration of H_3O^+ ions to the pH.

pH is defined using the following equation: pH = –log [H_3O^+]

The concentration of H_3O^+ ions is typically less than 1 M, therefore the exponents are negative. Since it is easier to work with positive numbers, pH was defined as the –log.

Table: Relating concentration of H_3O^+ to pH

Concentration of H_3O^+ (M)	Exponent Notation (M)	Log of [H_3O^+]	–Log [H_3O^+] = pH
1	10^0	0	0
0.1	10^{-1}	–1	1
0.01	10^{-2}	–2	2
0.0000001	10^{-7}	–7	7
0.0000000000001	10^{-13}	–13	13

14. Does a solution with a pH of 4 have more or less H_3O^+ ions than a solution with a pH of 7? How many times more or less?

A solution with a pH of 4 has 1000 times more H_3O^+ ions than a solution with a pH of 7. One change in pH results in10 times the number of H_3O^+ ions.

Calculating H_3O^+ Ion Concentration from pH

Guide the students through an example of calculating H_3O^+ ion concentration from pH and then let them practice before starting the lab.

15. **Calculate the H_3O^+ ion concentration of an unknown solution that has a pH of 10.6.**

$pH = -\log [H_3O^+]$

The log is the power (or exponent) to which 10 (the logarithmic base) must be raised to give the original number. Therefore $10^{-pH} = [H_3O^+]$

$10^{-10.6} = [H_3O^+]$

$2 \times 10^{-11} = [H_3O^+]$

16. **Calculate the H_3O^+ ion concentration of an unknown solution that has a pH of 2.2.**

$10^{-pH} = [H_3O^+]$

$10^{-2.2} = 6 \times 10^{-3} = [H_3O^+]$

Lab Preparation

These are the materials and equipment to set up prior to the lab.

The following instructions will make 100 mL of baking soda (sodium bicarbonate) solution. This is enough for 20 lab groups.

Prepare 100 mL of 0.5 M sodium bicarbonate ($NaHCO_3$) solution by adding 4.2 grams of baking soda to enough distilled water to create 100 mL of solution.

Teacher Tips:

- The pH sensor calibration is saved in the data file in which it was performed. This requires the pH sensor be recalibrated every time a new file is opened or a different data collection system is used.

- To save time and materials, prepare a class set of labeled test tubes that are already ⅓ full with each of the ten household items. Reuse the same samples for multiple class periods. Be sure to remind the students to not contaminate their solutions.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Many household chemicals are skin, eye, and respiratory irritants, including window cleaner, vinegar, lemon juice, and bleach.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

4	1	3	2
Graph pH versus hydronium ion (H_3O^+) concentration.	Place 5 mL of each household chemical in a clearly labeled test tube.	Calculate the hydronium (H_3O^+) ion concentration for each household chemical using the measured pH values.	Measure the pH of each household chemical.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a pH sensor to your data collection system. �(2.1)

3. ☐ Place 25 mL of pH 4 buffer solution in a 50-mL beaker and 25 mL of pH 10 buffer solution in a second 50-mL beaker. Use these solutions to calibrate the pH sensor.�(3.6)

4. ☐ Using the terms "accuracy" and "precision," explain why is it necessary to calibrate the pH sensor?

A pH sensor needs to be calibrated to ensure accurate results. A pH sensor that is not calibrated will give precise results, but they may not be accurate.

5. ☐ Configure the data collection system to manually collect pH for different household chemicals in a table. Define household chemicals as the manually entered text data. �(5.2.1)

6. ☐ Obtain 10 clean, dry test tubes.

7. ☐ Label each test tube with a household chemical name. The household chemicals are listed in Table 1 in the Data Analysis section below.

8. ☐ Add 5 mL of each household chemical to the appropriately labeled test tube. Each test tube should be approximately one third full.

9. ☐ Does the amount of liquid used in each test tube need to be exact? Explain.

No. The volume required needs to be enough for the pH sensor to be fully submerged in the solution, but not so full that it overflows when the pH sensor is inserted. The amount of solution used does not affect the pH.

Collect Data

10. ☐ Start a new manually sampled data set. �(6.3.1)

11. ☐ Place the pH sensor into the first sample. Make sure the bulb of the pH sensor is fully submerged.

12. ☐ Leave the pH sensor in the solution until the reading stabilizes (about 1 minute) and then record the data point.�(6.3.2)

13. ☐ Remove the sensor from the sample and thoroughly rinse it with clean water.

14. ☐ Repeat the steps above to determine the pH for all the samples.

Note: Remember to thoroughly rinse the pH sensor with clean water after testing each solution.

15. ☐ Why is it necessary to rinse the pH sensor after each sample is tested?

To avoid contaminating the next sample. If any of the samples mix together, the pH values will change.

16. ☐ When you have recorded all of your data, stop the data set.�(6.3.3)

17. ☐ Save your data file and clean up according to the teacher's instructions.�(11.1)

Data Analysis

1. ☐ Calculate the hydronium ion (H_3O^+) for each of the household chemicals using the measured pH value. Follow the steps below to do this on your data collection system.

a. Enter the equation given below into your data collection system's calculator. �(10.3)
concentration = 10^–(pH)

b. Add a column to the table on your data collection system to display the calculated hydronium (H_3O^+) ion concentration. �(7.2.2)

2. ☐ Copy the pH and H_3O^+ ion concentration data from your data collection system to the corresponding columns in Table 1 below.

Table 1: Household chemicals, their pH and H_3O^+ concentrations

	Household Chemical	pH	$[H_3O^+]$ (M)
1	Vinegar	2.6	3×10^{-3}
2	Lemon juice	2.5	3×10^{-3}
3	Soft drink	2.9	1×10^{-3}
4	Window cleaner	8.6	3×10^{-9}
5	Tap water	8.0	1×10^{-8}
6	Milk	6.9	1×10^{-7}
7	Coffee	5.1	8×10^{-6}
8	Baking soda	8.5	3×10^{-9}
9	Liquid soap	10.6	2×10^{-11}
10	Bleach	11.8	1×10^{-12}

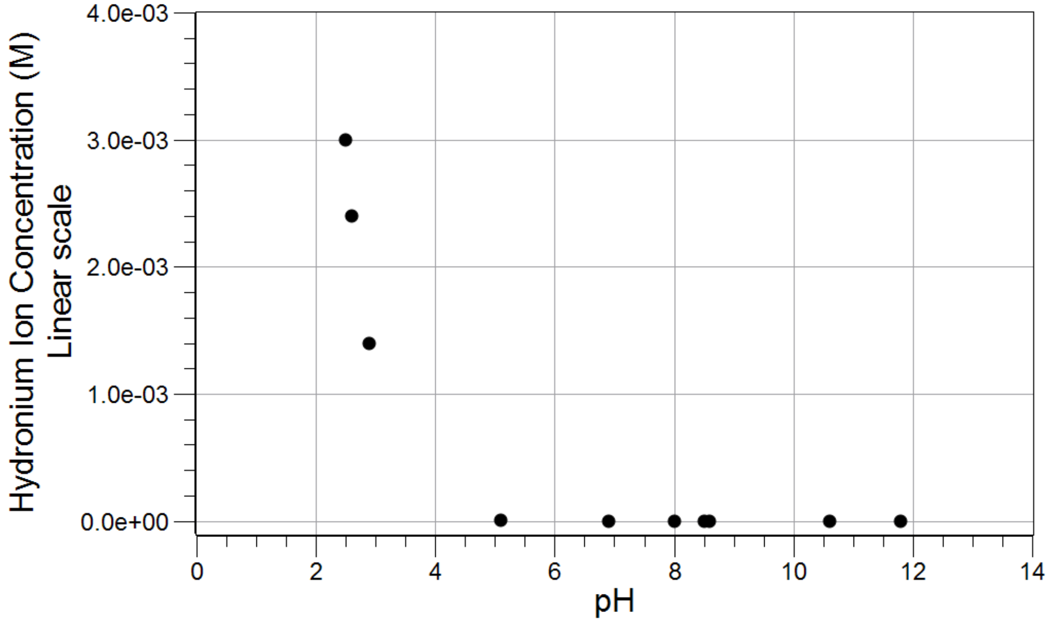
Teacher Tip: Notice that the $[H_3O^+]$ has only one significant figure due to the rules of significant figures with logarithms. The number of decimal places in the pH reading indicates the number of significant figures that the $[H_3O^+]$ has in its coefficient.

3. ☐ Display H_3O^+ concentration versus pH on a graph.�(7.1.1)

Note: To graph a scatter plot of the data points hide the connecting lines between data points feature. �(7.1.8)
If needed, adjust the scale of the graph to show all the data points. �(7.1.2)

4. ☐ Plot or print a copy of the graph of H_3O^+ Concentration (M) versus pH. Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

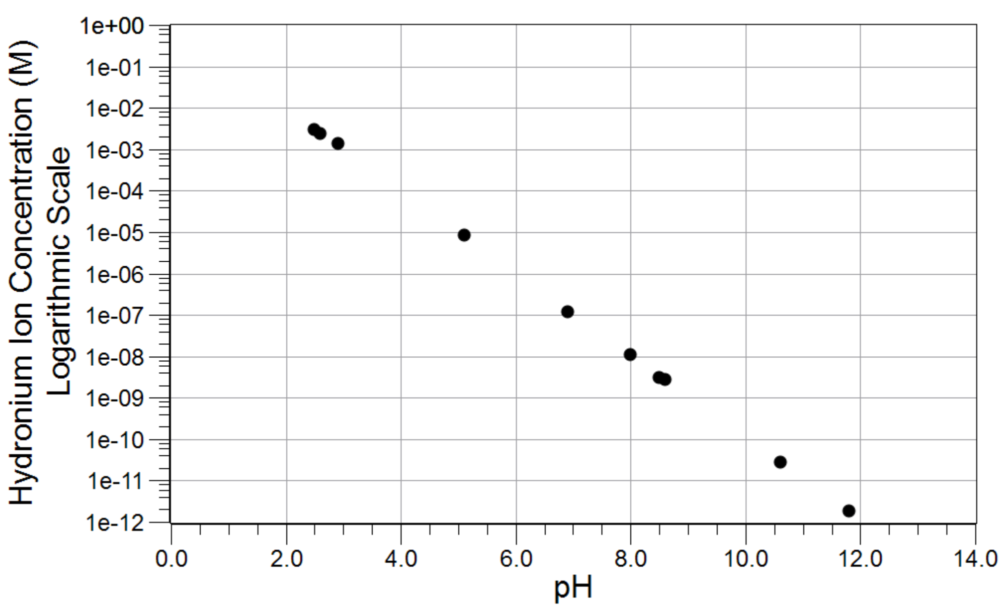
Hydronium Ion Concentration versus pH of Household Chemicals on a Linear Scale



5. ☐ Draw a scatter plot of H_3O^+ ion concentration versus pH on a logarithmic scale.

Teacher note: This graph will need to be drawn because the data collection system cannot create a logarithmic scale on its axes.

Hydronium Ion Concentration versus pH of Household Chemicals on a Logarithmic Scale



Analysis Questions

1. What is pH and why is the pH scale used?

pH is a measure of the concentration of H_3O^+ ions and is mathematically defined as pH = $-\log [H_3O^+]$. The pH scale is used because the numbers 0 to 14 are easier to use than the numbers 1 through 1×10^{-14}.

2. Explain the relationship between pH and H_3O^+ ion concentration.

There is an inverse relationship between pH and H_3O^+ ion concentration. As the pH increases, the H_3O^+ ion concentration decreases.

3. Define the term "acid" and explain the why there are strong and weak acids.

An acid is a substance that donates H^+ ions. The strength of an acid depends on the degree to which the acid dissociates. Strong acids completely dissociate producing a greater number of H_3O^+, whereas weak acids only partially dissociate and produce a smaller amount of H_3O^+ ions.

4. Identify which of the household chemicals tested are acids and list them in order from the lowest to highest pH.

Lemon juice, vinegar, soft drink, coffee, milk

5. Define the term base and explain why there are different strengths of bases.

A base is a substance that accepts H^+ ions. The strength of a base depends on the degree to which the base attracts H^+ ions. Strong bases bond readily with all possible H^+ ions and thus remove the H_3O^+ ions from solution. Weak bases, on the other hand, do not attract H^+ ions as strongly and thus allow a portion of them to remain in solution.

6. Identify which of the household chemicals tested are bases and list them in order from highest to lowest pH.

Bleach, liquid soap, window cleaner, baking soda solution, tap water

Synthesis Questions

Use available resources to help you answer the following questions

1. A nitric acid solution has a pH of 1 and a hydrochloric acid solution has a pH of 3. Which acid solution is more concentrated and by how much?

The nitric acid solution is 100 times more concentrated than the hydrochloric acid solution.

2. What is pOH and how is it related to pH?

pOH is a measure of the hydroxide ion concentration. Mathematically pOH = $-\log [OH^-]$. There is an inverse relationship between pH and pOH. As the pH value of a substance increases, its pOH value decreases.

$[H+][OH^-] = K\text{w}$

$-\log[H+] + -\log[OH^-] = -\log K\text{w} = -\log 1.0 \times 10^{-14}$

pH + pOH = 14

3. If an acid is added to a basic solution, what do you expect to happen to the pH of the basic solution? Why?

The pH of the basic solution should decrease because the added hydrogen ions increase the concentration of H_3O^+ ions in the solution and thus lower the pH.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Why is a 0.1 M hydrochloric acid solution a stronger acid than a 0.1 M acetic acid solution?

- **A. Because more of the H^+ ions dissociate into the solution**
- **B.** Because less of the H^+ ions dissociate into the solution
- **C.** Because there are equal numbers of H^+ ions in the two solutions
- **D.** Because there are equal numbers of OH^- ions in the two solutions

2. How is an aqueous solution of a base different from an aqueous solution of an acid?

- **A.** A basic solution conducts electricity and an acidic solution does not
- **B.** A basic solutions will cause an indicator to change color and an acid will not
- **C.** A basic solution has a greater concentration of H_3O^+ than OH^-
- **D. A basic solution has a lower concentration of H_3O^+ than OH^-**

3. Pure water has a pH of 7 and toothpaste has a pH of 10. The water contains how many times the number of H_3O^+ ions as toothpaste?

- **A.** 1/100
- **B.** 3
- **C.** 10
- **D. 1000**

4. An unknown solution has an H_3O^+ concentration of 6.0×10^{-10} M. This solution is:

- **A.** Acidic
- **B. Basic**
- **C.** Neutral
- **D.** Concentrated

5. An unknown solution has a pH of 4.0. This solution is:

- **A. Acidic**
- **B.** Basic
- **C.** Neutral
- **D.** Concentrated

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. **Acids** are found in numerous substances all around us such as soft drinks, salad dressing, and rain water. Acids taste **sour**, cause indicators to change color, and react with certain **metals** to form hydrogen gas. At the molecular level, an acid is a substance that **donates** a hydrogen ion (H^+) which will form an increased concentration of the hydronium ion (H_3O^+) when the acid is dissolved in water. **Strong** acids dissociate fully in water, whereas **weak** acids only partially dissociate. Acids have pH values **less than** 7. The lower the pH, the **more** acidic the solution.

2. Bases may not be as familiar to you, but they are just as numerous as acids. **Bases** are found in personal hygiene products, cleaning products, and food. Bases taste **bitter**, feel slippery, and react with oil and grease. At the molecular level, a base is a substance that **accepts** or bonds with a hydrogen ion (H^+). By bonding with the H^+ ion, a base causes the concentration of the hydronium ion (H_3O^+) to **decrease**. **Strong** bases have a strong force of attraction and readily bond with H^+ ions which reduces the H_3O^+ ion concentration in solution. Bases have **pH** values greater than 7. The **higher** the pH, the more basic the solution.

Extended Inquiry Suggestions

Have the students bring in labels of different household chemicals and use the ingredients list to predict the pH of the substance. If possible, have the students test their predictions after making their hypotheses.

Investigate the conductivity of strong and weak acids and bases.

Determine the effect of concentration on pH using different concentrations of household chemicals.

Determine the effect of the amount of a substance on its pH.

Determine the effect of mixing acids and bases on the pH of the newly formed solution.

Compare pH measurements taken with a pH sensor against those determined using pH paper or other indicators.

16. Electrochemical Battery: Energy from Electrons

Objectives

Experimentally place metal reactants in their proper order on the table of standard electrode potentials. Through this investigation, students:

- Describe electricity as the flow of electrons
- Learn that some metals form ions easier than others, and that the ease with which ions are formed determines the amount of energy they can produce
- Learn that electrochemical batteries (voltaic cells, batteries) produce useable energy by separating two halves of a spontaneous chemical reaction in which the products have lower potential energy than the reactants
- Learn that an electrochemical battery contains two different metals, a path for ion movement, and an electrolyte solution

Procedural Overview

Students conduct the following procedures:

- Construct electrochemical batteries using electrolyte solutions, wires, and various metal electrodes.
- Measure voltage produced in an electrochemical battery using different metals as the anode and copper as the cathode.

Time Requirement

♦ Preparation time	30 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	40 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Voltage sensor
- Graduated cylinder, 50-mL
- Beaker (2), 50-mL
- Alligator clip (2), 1 black,1 red
- Wash bottle and waste container
- Thick string or yarn, 20-cm
- Knife to cut fruit
- Paper towels
- 0.1 M Sodium chloride (NaCl), 5 to 10 mL[1]
- 0.1 M Hydrochloric acid (HCl), 50 mL[2]
- Copper strip[3]
- Zinc strip[3]
- Magnesium strip[3]
- Nickel strip[3]
- Iron strip[3]
- Lemon[4]
- Tomato[5]

[1] To formulate using table salt (sodium chloride, NaCl), refer to the Lab Preparation section.

[2] To formulate using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

[3] See the Lab Preparation section for ideas on where to obtain the electrode metals.

[4] Any citrus fruit can replace the lemon.

[5] The tomato can be substituted with an apple, kiwi fruit, or potato.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Bohr model of the atom
- Kinetic and potential energy
- Electrolyte solutions
- Law of conservation of matter
- Valence electron shells
- Ion formation

Related Labs in this Guide

Labs conceptually related to this one include:

- Molar Mass of Copper
- Electrolyte and Non-Electrolyte Solutions
- Single Replacement Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

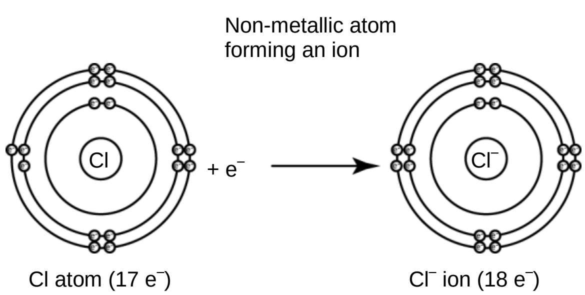
- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Monitoring live data in a digits display �(6.1)

Background

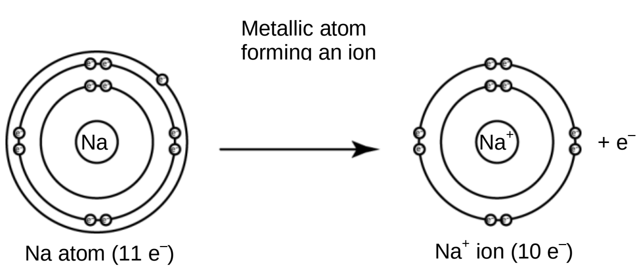
All work requires the input of some form of energy. Energy can be classified as either kinetic or potential. Kinetic energy is energy in motion and includes light, sound, motion, heat, and electricity. Potential energy is stored energy that has yet to be released as kinetic energy.

In chemistry, energy is stored in the bonds between atoms. These bonds form because of the atoms' tendencies to either gain or lose electrons. The only atoms that tend not to bond are those of the inert gases. These atoms do not bond because their electrons are already arranged in a configuration with low potential energy and are thus very stable. The low potential energy, and resulting stability, comes from having a complete outermost energy shell (valence shell).

All other elements can lower their potential energy by filling their valence electron shell. Non-metallic atoms can lower their potential energy and become more stable by gaining electrons to fill their valence shell.



Metallic atoms, on the other hand, need to lose electrons to have a full outermost energy level and thus lower their potential energy to become more stable.



An electrochemical battery (voltaic cell) is designed to take advantage of the relative tendencies of metals to lose electrons. These lost electrons must obey the law of conservation of matter and go somewhere else; the transfer of electrons from one atom to another is known as a “redox” reaction. This is a combination of two separate “half reactions” that always take place together: an “oxidation” half-reaction in which electrons are lost by one reactant, and a “reduction” half-reaction in which the electrons are gained by another. (A number of mnemonics exist including “oil rig” = **o**xidation **i**s a **l**oss of electrons, **r**eduction **i**s a **g**ain of electrons and “Leo the Lion goes ger” = **L**oss of **e**lectrons is **o**xidation, **g**ain of **e**lectrons is **r**eduction.)

$Zn(s)$	→	$Zn^{2+}(aq) + 2e^-$	(oxidation half-reaction)
$Cu^{2+}(aq) + 2e^-$	→	$Cu(s)$	(reduction half-reaction)
$Zn(s) + Cu^{2+}(aq)$	→	$Zn^{2+}(aq) + Cu(s)$	(overall redox reaction)

A metal in a redox reaction either gains or loses electrons depending on its oxidation strength relative to the other reactant. A list of electrode reduction potentials ranks the elements based on their tendencies for gaining electrons; those higher on the list will take electrons spontaneously from those lower on the list.

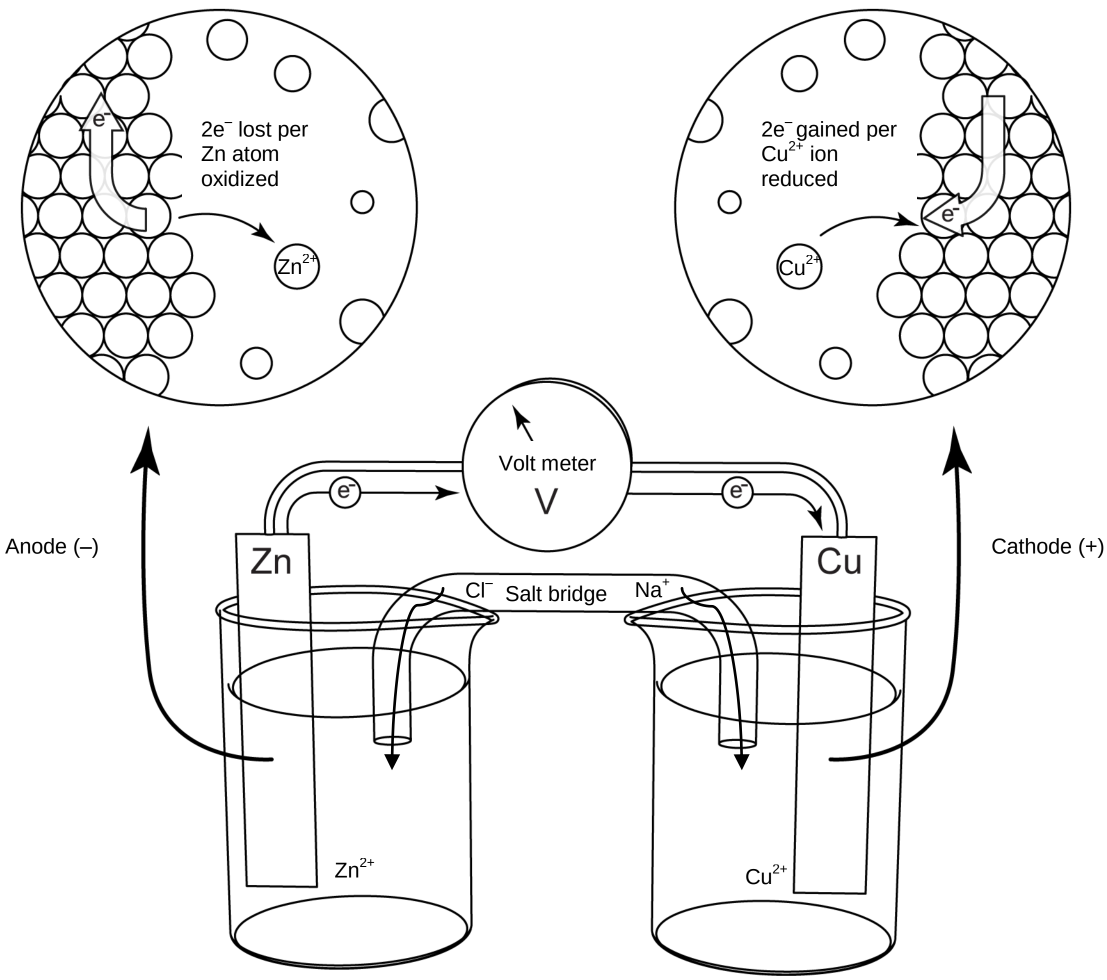
Standard reduction potentials in aqueous solution at 25 °C

Reduction Half-Reaction			E° (V)
$F_2(g) + 2e^-$	⇌	$2F^-(aq)$	+2.87
$Au^{3+}(aq) + 3e^-$	⇌	$Au(s)$	+1.50
$Ag^+(aq) + e^-$	⇌	$Ag(s)$	+0.80
$Cu^{2+}(aq) + 2e^-$	⇌	$Cu(s)$	+0.34
$2H_3O^+(aq) + 2e^-$	⇌	$H_2(g) + 2H_2O(l)$	0.00
$Sn^{2+}(aq) + 2e^-$	⇌	Sn s	–0.14
$Ni^{2+}(aq) + 2e^-$	⇌	$Ni(s)$	–0.25
$Fe^{2+}(aq) + 2e^-$	⇌	$Fe(s)$	–0.44
$Zn^{2+}(aq) + 2e^-$	⇌	$Zn(s)$	–0.76
$Al^{3+}(aq) + 3e^-$	⇌	$Al(s)$	–1.66
$Mg^{2+}(aq) + 2e^-$	⇌	$Mg(s)$	–2.37
$Na^+(aq) + e^-$	⇌	$Na(s)$	–2.71
$K^+(aq) + e^-$	⇌	$K(s)$	–2.93

For example, if comparing zinc and copper, copper has the higher reduction potential on the list and will take electrons from zinc. Comparing zinc and magnesium, however, shows that zinc has the higher reduction potential and will, therefore, take electrons from magnesium.

An electrochemical battery forces the electrons being transferred from one metal to another to travel through a circuit. The electrochemical battery then uses these flowing electrons (electrical current) to do work; the greater the flow of electrons, the higher the amperage. The voltage produced by the electrochemical battery depends on the metals that are used; metals that are farther apart on the table of electrode reduction potentials will have greater differences in electrochemical potential. This electrochemical potential is measured as voltage (V).

An electrochemical battery physically separates the two half-reactions, joining them by a wire conductor to allow the electrons to flow from one metal to another. The half-reactions are also joined by a salt bridge which allows ions to flow between the cells in order to balance the charges and complete the circuit. In electrochemical batteries, reduction takes place at the positive terminal (the electrode gaining electrons, cathode), and oxidation takes place at the negative terminal (the electrode losing electrons, anode).



An electrochemical battery (voltaic cell) made with zinc and copper electrodes.

In a dry-cell battery, the two cells are divided by a non-conductive, semi-permeable, paper membrane separator.

Pre-Lab Discussion and Activity

Energy

Engage your students in a discussion about the importance of energy in their lives and in our world. Have the students define energy, differentiate kinetic energy from potential energy, give examples of different forms of energy, and explain the law of conservation of energy.

1. What is energy? How is energy important in your life and our world?

Energy is required to do work (to push or pull an object over a distance). Energy is required for humans to live and is used to make life easier (transportation, light, washing machine) and more enjoyable (cell phones, MP3 players, televisions).

2. What are the two major classifications of energy and how are they different?

The two classifications of energy are kinetic energy and potential energy. Kinetic energy is energy in motion. It is the energy that enables us to do work. Potential energy is related to position and is stored energy.

3. List several examples of kinetic energy.

A ball flying in the air, a car driving down the street, and person running are all examples of kinetic energy. Atoms and molecules also exhibit kinetic energy when they vibrate, rotate, or move across space.

4. List several examples of potential energy.

A ball sitting on a high shelf, a rock on the edge of a cliff, an arrow pulled back in a bow, and a compressed spring are all examples of potential energy. At the molecular level, energy is stored in the chemical bonds holding atoms together.

5. What does the law of conservation of energy state?

The law of conservation of energy states that energy cannot be created or destroyed. When energy is used it changes from one form to another.

6. What are the different forms of kinetic energy?

Potential energy can be released as kinetic energy and seen as light, sound, electricity, heat, and motion.

Batteries

Hold up a few examples of batteries, such as a cell phone battery and an AA battery. Engage the students in a discussion about why we use batteries and where the power comes from. Review key terms such as power, energy, and voltage. Show the students how to use a voltage sensor to test a battery by placing the leads onto each end of the battery. End by discussing the types of energy and energy transformations that occur in batteries.

7. Why do we use batteries?

We use batteries to power things. Power is energy per unit time measured in watts (1 W = 1 J/s). Batteries provide electrical energy and this electrical energy (electricity) does work that we want it to, like powering our cell phones or turning a motor.

8. There are many different types of batteries. How do the batteries you have used differ?

There are many different types of batteries; they have different sizes, shapes, voltages, life times, and some are reusable (rechargeable).

9. There are many different types and sizes of batteries, but what do all batteries have in common?

All batteries have a "positive" end (cathode) and a "negative" end (anode) and produce electricity when the terminals are connected.

10. How can I determine if a battery is still "good"?

Use a voltage sensor to measure the voltage. If the voltage is the same as what is written on the battery, then the battery is still strong. If the voltage is lower than what is written on the battery then it is likely that the battery is no longer good.

11. What is voltage?

Voltage measures energy per unit charge. It comes from the difference in potential between one electrode and the other. This difference is important, because it is the force that moves electrons through the circuit. The greater the difference in potential, the stronger the force moving the electrons.

12. What types of energy are involved? Explain each type of energy.

A battery contains chemical potential energy. Chemical potential energy comes from the electrostatic attraction that holds atoms together (bonds). When a battery is being used, the chemical energy is changed into electricity (the flow of electrons), and the electrical energy is changed into the form of energy required by the device the battery is in (light, sound, or heat).

13. How does chemical energy in bonds become electrons flowing?

A chemical reaction occurs. Electrons are transferred from one reactant to another in a reduction-oxidation (redox) reaction. An electrochemical half-cell forces the reactant giving up their electrons to be in a different location than the reactant accepting the electrons. Thus, the electrons are forced through a wire before they reach their recipient. Thus, the energy can be harnessed.

Chemical Reaction- Copper(II) Sulfate and Iron

Demonstrate the redox reaction between copper and iron by placing a piece of iron metal (as an iron nail) into a test tube containing a 0.1 M copper(II) sulfate solution:

$$CuSO_4(aq) + Fe(s) \rightarrow Cu(s) + FeSO_4(aq)$$

Copper has a higher reduction potential than iron, so the copper(II) ions will take electrons from the iron metal. This causes the copper(II) ions to gain electrons (reduction) and become solid copper (it will coat the iron nail with a thin layer of copper). The iron metal will become iron(II) ions after giving up some of its electrons (oxidation). The color of the solution will change from blue to green as the blue copper(II) ions are consumed and replaced by green iron(II) ions. If you try putting copper into an iron(II) sulfate solution, nothing happens because the copper has a greater affinity for electrons than iron.

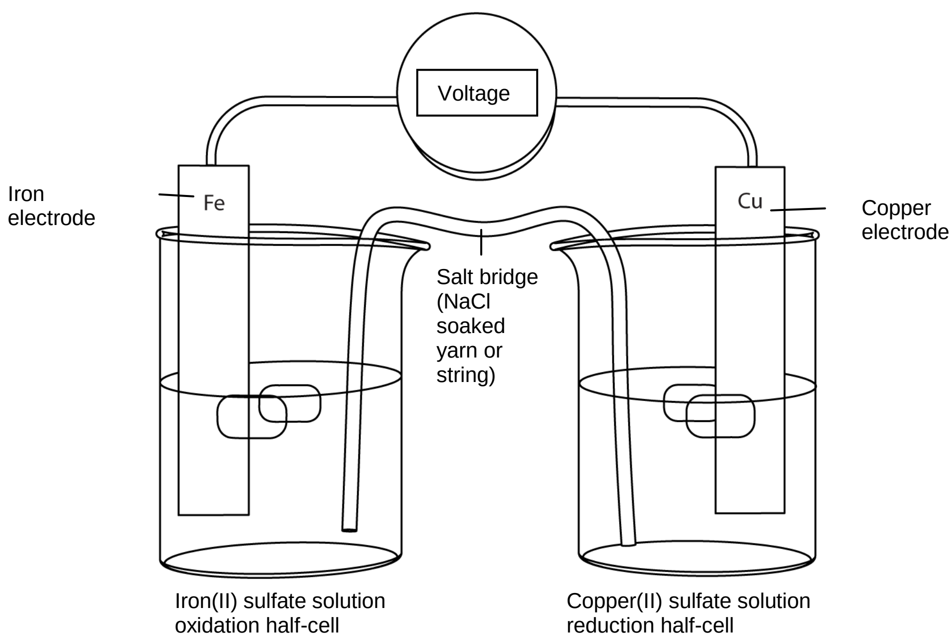
In order for the copper(II) ions to become copper metal, electrons must be transferred to them. If the electrons can be made to move through a wire on their way to the copper(II) ions, then you have produced an electric current. This is how a chemical battery works. Voltage, sometimes called *cell potential*, is the measure of the potential energy change in the battery. The larger the potential energy change that occurs in the chemical reaction, the larger the voltage.

Chemistry Inside a Battery: Electrochemical (Voltaic) Cells

In the previous demonstration, the electrons were free to transfer directly between the copper and the iron. In order to force the electrons through a wire to produce a current of electricity, the reaction must be separated into its individual parts. Demonstrate this with reduction taking place in one beaker and oxidation taking place in another. Each beaker is referred to as a half-cell, or half-reaction, since only one half of the overall redox reaction is taking place.

Construct the oxidation half-cell by placing an iron metal electrode in a beaker containing iron(II) sulfate; construct the reduction half-cell by placing a copper metal electrode in a beaker containing copper(II) sulfate. Connect the two electrodes using alligator clips attached to a voltage sensor in between. Note the absence of a voltage reading. Next, link the two beakers with a piece of yarn or heavy string dipped in a solution of sodium chloride to complete the circuit. Note the voltage reading.

Replace the iron electrode with one made of a different metal. Note the new voltage reading. As the reaction progresses, the electrolyte solution of the reduction half-cell may change color as the blue copper(II) ions are being plated as copper metal onto the copper electrode and are being replaced by green iron(II) ions traveling through the salt bridge. The color of the electrolyte solution in the oxidation half-cell will not change color since the iron(II) ions are being replaced as the iron electrode is being oxidized.



14. **Why are the electrodes in separate beakers?**

The separate beakers prevent the system from “short circuiting”. By connecting the separate beakers with a wire between the electrodes, electrons are forced to travel from one electrode to another through the wire producing an electric current as measured by the voltage sensor.

15. In which direction do the electrons flow in this system?

Electrons are produced in the oxidation half-cell at the iron electrode (– terminal, anode) and flow towards the reduction half-cell containing the copper electrode (+ terminal, cathode).

16. Why did the voltage sensor not show a reading before the salt bridge was added?

The circuit was incomplete.

17. What is the purpose of the salt bridge?

The salt bridge completes the circuit and allows ions to travel from one beaker to the other to balance the charges built up during the reaction. As electrons leave the oxidation half-cell, a positive charge builds as iron metal becomes iron(II) ions in solution requiring negative ions to flow into the beaker to counteract them. Likewise, as electrons enter the reduction half-cell, a negative charge builds as copper(II) ions leave the solution and become copper metal requiring positive ions to flow into the beaker to replace them.

18. Why did the voltage change when the iron electrode was replaced by a different metal?

Different metals have different tendencies to gain or lose electrons. Voltage is a measurement of the difference between the two electrodes' affinities for electrons (electrode potentials). The greater the difference in electrode potentials, the greater the measured voltage.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids create heat; be extra careful when handling freshly prepared solutions and glassware as they may be very hot.
- ♦ Always add acids to water, not the other way around, as the solutions may boil vigorously.
- ♦ Handle concentrated acids in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Prepare 100 mL of 0.1 M sodium chloride solution: Dissolve 0.58 g table salt (sodium chloride) in 100 mL of distilled water. This is enough for *10* lab groups.

2. Prepare 1000 mL of 0.1 M hydrochloric acid from either concentrated (12 M) or dilute (6 M) HCl. This is enough for *20* lab groups.

Starting with concentrated (12 M) HCl

- **a.** Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.
- **b.** Slowly add 8.3 mL of 12 M HCl to the beaker with continuous stirring.
- **c.** Allow the solution to cool, then carefully pour into a 1000-mL volumetric flask and dilute to the mark with distilled water.
- **d.** Cap and invert three times carefully to ensure complete mixing.

Starting with dilute (6 M) HCl

- **a.** Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.
- **b.** Add 16.7 mL of 6 M HCl to the water and dilute to the mark with distilled water.
- **c.** Cap and invert three times carefully to ensure complete mixing.

3. Electrodes:

Sources for the metals can vary from ordering directly from science suppliers online to visiting salvage yards and metal recyclers. You may already have some of the metals on hand. Remove any sharp edges.

Copper can be found as electrical wire, water pipe, and sheeting available at hardware stores.

Zinc can be found as galvanized nails available at hardware stores.

Magnesium is usually available only from chemical supply companies.

Iron is fairly common and can be found as nails, among other sources.

Nickel can be found as wire used in jewelry making at some specialty craft stores.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Beware of sharp edges on the metal electrodes.
- ♦ Do not eat any of the food used in this investigation. Discard the fruit according to your teacher's instructions.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

1	2	3	4
Assemble the electrochemical battery using 0.1 M HCl as the electrolyte and copper and zinc as the electrodes.	Test the voltage of zinc, then replace zinc with three different metals and test the voltage of each.	Replace the HCl with a lemon and then a tomato and find the voltage produced by each metal.	Use the data collected to list the metals in order by the amount of voltage they produced when used in a battery.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a voltage sensor to the data collection system. �(2.1)

3. ☐ Connect the red alligator clip to the red voltage sensor lead and connect the black alligator clip to the black voltage sensor lead.

4. ☐ Pour approximately 25 mL of 0.1 M hydrochloric acid (HCl) into each 50-mL beaker. The HCl is the electrolyte solution.

5. ☐ What is an electrolyte solution? What makes this solution electrolytic?

An electrolyte solution is a solution that conducts electricity. For a solution to conduct electricity it must contain ions. This solution contains hydrogen ions (H^+) and chloride ions (Cl^-).

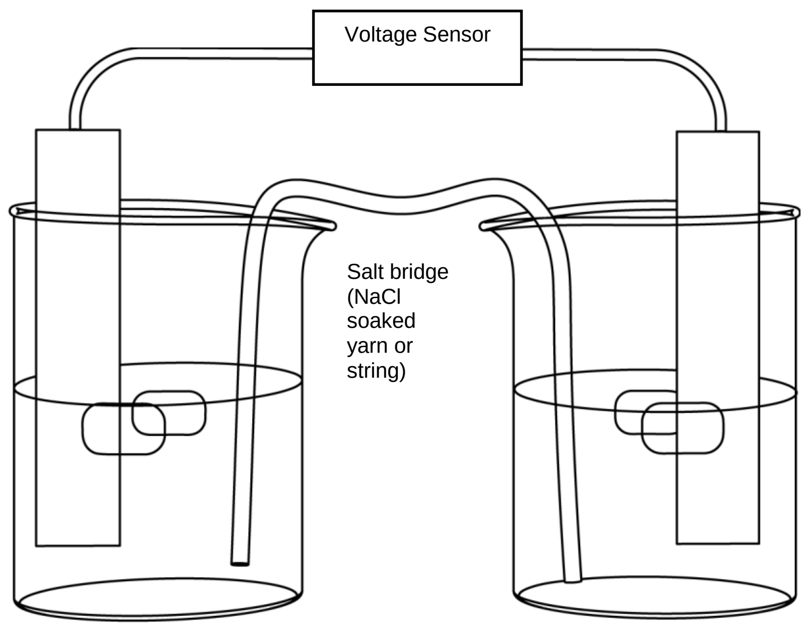
6. ☐ Place a zinc strip in one beaker and a copper strip in the other.

Note: Copper will be used in all the electrochemical batteries as a reference to compare with the other metals.

7. ☐ Is it necessary to have a separate beaker for the copper strip and the zinc strip? Explain.

For the reaction to occur, the zinc and copper do not have to be separated into two beakers. They do not have to be separated because the electrolyte in each beaker is the same. You just have to make sure that the two electrodes do not touch each other.

8. ☐ Wet a 20 cm piece of string with the 0.1 M sodium chloride solution and hang it between the two beakers with the ends submerged in the electrolyte solution. The string is called a salt bridge and allows the ions to move between the beakers to keep the solutions neutral.



9. ☐ What are the dependent variable and the independent variable as well as the units for each used in this experiment?

The dependent variable is voltage and is measured in volts.

The independent variable is the metal type.

Collect Data

10. ☐ Monitor live voltage data in a digits display. �(6.1)

11. ☐ Connect the black lead of the voltage sensor to the zinc strip and the red lead to the copper strip using the alligator clips.

12. ☐ Record the voltage of the zinc electrode in the 0.1 M HCl column in Table 1 below.

Table 1: Voltage readings of metals in different electrochemical batteries

Metal	0.1 M HCl Battery (V)	Lemon Battery (V)	Tomato Battery (V)
Zinc (Zn)	0.90	0.81	0.79
Magnesium (Mg)	1.59	1.69	1.58
Nickel (Ni)	0.04	0.04	0.00
Iron (Fe)	0.24	0.29	0.37

13. ☐ Disconnect the black lead, remove the zinc strip, replace it with a magnesium strip, and reconnect the black lead.

14. ☐ Record the voltage produced using magnesium in the 0.1 M HCl column in Table 1 above.

15. ☐ Repeat this process with a nickel strip and an iron strip and record the voltages in the 0.1 M HCl column in Table 1 above.

16. ☐ When you have collected all your data for 0.1 M HCl, disconnect the voltage sensor leads and remove both electrodes from the beakers.

17. ☐ Clean all the metal strips with water and dry them.

18. ☐ Roll the lemon firmly on the tabletop with the palm of your hand.

19. ☐ Use a knife to make two slits wide enough for the electrodes to be inserted about 2 to 3 cm apart in the lemon.

20. ☐ Insert the copper electrode deeply into the lemon through one of the slits.

21. ☐ Insert the zinc electrode in the other slit.

Note: Make sure that the electrodes do not touch each other.

22. ☐ What is acting as the electrolyte solution in the lemon?

The lemon juice.

23. ☐ Connect the black lead of the voltage sensor to the zinc electrode and the red lead to the copper electrode.

24. ☐ Record the voltage of zinc in the Lemon column in Table 1 above.

25. ☐ Disconnect the black lead, remove the zinc strip, replace it with the magnesium strip, and reconnect the black lead.

26. ☐ Record the voltage produced using magnesium in the Lemon column in Table 1 above.

27. ☐ Repeat this process with the nickel strip and the iron strip and record the voltages in Table 1 above.

28. ☐ What is voltage a measure of?

Voltage measures the energy per unit charge. It is the difference in electrode potential between the two metals. This difference in potential reflects the difference in electron affinity that exists between the two metals.

29. ☐ When you have collected all your data, disconnect the voltage sensor leads and remove the electrodes from the lemon.

30. ☐ Clean all the electrodes with water and dry them.

31. ☐ Reassemble an electrochemical battery using a tomato instead of the lemon and retest the voltage produced using the four metals.

32. ☐ Record the voltage for each metal in the Tomato column in Table 1 above.

33. ☐ Clean up according to the teacher's instructions.

Data Analysis

1. ☐ List the metals in order by the amount of voltage they produce when used in a battery with 0.1 M HCl electrolyte solution (list from highest voltage to lowest voltage).

Mg > Zn > Fe > Ni

2. ☐ List the metals in order by the amount of voltage they produce when used in a lemon battery (list from highest voltage to lowest voltage).

Mg > Zn > Fe > Ni

3. ☐ List the metals in order by the amount of voltage they produce when used in a tomato battery (list from highest voltage to lowest voltage).

Mg > Zn > Fe > Ni

Analysis Questions

1. What components are necessary to make a battery?

Two different metals, a conducting wire, an electrolyte solution, and a salt bridge.

2. Did the type of electrochemical battery used (HCl, lemon, tomato) affect the ranking of the metals? What was the role of the lemon and tomato?

The electrochemical battery used did not affect the ranking. The lemon and tomato provided the electrolyte solution that completed the circuit.

3. What was the source of the electrons in the battery?

The more active of the two metals; the one with the greatest oxidation potential (lowest reduction potential).

4. Which pair of electrodes would make the most powerful battery? How do you know?

Magnesium and copper will make the most powerful battery because it had the largest potential difference resulting in the greatest voltage.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Batteries come in all shapes, sizes, and voltages. Car batteries, cell phone batteries, computer batteries, and flashlight batteries are all different from each other. Explain how each of these batteries are similar and suggest a reason for their different voltages.

All of these batteries contain two types of metals, a wire for electrons to flow through (when connected), and an electrolyte solution separated by a membrane acting as a salt bridge. The different types of batteries use different types of metals and different electrolyte solutions.

2. Why do you think batteries "go dead"?

Because the metal electrode that produces electrons (the anode where oxidation takes place) is consumed through the chemical reaction.

3. Why do you think many electronic devices require more than one battery?

The circuitry in the electronic system either requires more voltage (or more current) than one battery can supply or to extend the life of the batteries.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. In a chemical cell, electrical current is caused by:

- **A.** Fruit
- **B.** Metals
- **C.** Protons
- **D.** **Electrons**

2. In the fruit battery, the fruit supplied:

- **A.** Metal
- **B.** Electrons
- **C.** **Electrolyte solution**
- **D.** Voltage

3. Voltage is a measure of:

- **A.** The number of electrons
- **B.** **The difference in the affinity for electrons**
- **C.** The time it takes for an electron to travel through a circuit
- **D.** The concentration of salt in fruit

4. In the fruit battery, the electrons were generated from:

- **A.** **The metal electrode**
- **B.** The fruit
- **C.** The wire
- **D.** The electrolyte solution

5. In an electrochemical battery, chemical energy is converted into:

- **A.** Potential energy
- **B.** **Electricity**
- **C.** Sound
- **D.** Light

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. All work requires the input of some form of **energy**. **Kinetic** energy is energy in motion and includes light, sound, motion, heat, and electricity. **Potential** energy is stored energy that has yet to be released. In chemistry, potential energy is stored in the **bonds** between **atoms**. These bonds form in order to arrange **electrons** in configurations with low potential energy.

2. An electrochemical battery takes advantage of the relative tendencies of **metals** to lose electrons. The transfer of electrons from one atom to another is known as a **redox** reaction. This is a combination of two separate half-reactions in which electrons are lost in **oxidation** and electrons are gained in **reduction**. Whether a metal in these reactions gains or loses electrons when paired with another depends on its relative placement on the list of electrode potentials. Those metals **higher** on the list will take electrons from the other metals. The difference in the two metals' potentials is measured as **voltage**.

3. An electrochemical battery containing an **electrolyte** solution separates the two metals and allows the transfer of electrons which flow through a wire connecting them. In order to balance the charges and complete the circuit, a **salt bridge** connects the cells allowing ions to flow. In the battery, reduction takes place at the cathode and oxidation takes place at the anode.

Extended Inquiry Suggestions

Determine if the electrolyte solution's concentration has an effect on the voltage of the cell. Use the Nernst equation to validate the results.

Replace the copper reference electrode with various metals and look for new patterns to emerge.

Test the effect of various size wires and salt bridges (thread, string, yarn, rope) on the measured voltage and current.

Determine the effect of the surface area of the electrode on the voltage of the cell.

Allow the cell to "die" while measuring voltage against time and plot the results. Make sure the dimensions of all the electrodes are as similar to each other as possible. Do some metals last longer? Is there a relationship between voltage and the "life" of the electrode?

Chemical Reactions

17. Evidence of a Chemical Reaction

Objectives

During this investigation, students:

- Observe the four main types of evidence that suggest a new chemical substance has formed
- Distinguish between physical changes and chemical reactions
- Identify processes as involving physical changes or chemical reactions
- Identify the reactants and products in a chemical reaction
- Explain the difference between exothermic and endothermic chemical reactions

Procedural Overview

Students conduct the following procedures:

- Perform three chemical reactions, collect temperature versus time data for each, and record evidence that a new substance was formed for each
- Perform three physical changes, and describe the resulting new physical appearance
- Perform three additional changes, and identify them as chemical reactions or physical changes based on the observations recorded
- Identify each chemical reaction as exothermic or endothermic

Time requirement

♦ Preparation time	20 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	50 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Fast response temperature sensor
- Balance (2-3 per class)
- Hot plate
- Graduated cylinder, 100-mL
- Graduated cylinder, 10-mL
- Beaker (2), 250-mL
- Test tube (7), 15-mm x 100-mm
- Test tube rack
- Test tube holder
- Stir rod
- Spatula
- Beaker for collecting rinse water
- Weighing paper
- Wash bottle filled with distilled (deionized) water
- Water, 255 mL
- Calcium carbonate ($CaCO_3$), ~0.2 g
- White vinegar (~5% acetic acid), 2 mL
- 1.0 M Citric acid ($C_6H_8O_7$), 2 mL[1]
- 1.0 M Sodium bicarbonate ($NaHCO_3$), 2 mL[2]
- 0.5 M Copper(II) sulfate ($CuSO_4$), 2 mL[3]
- 1.0 M Sodium hydroxide (NaOH), 2 mL[4]
- 0.05 M Silver nitrate ($AgNO_3$), 2 mL[5]
- 0.1 M Sodium chloride (NaCl), 2 mL[6]
- Lauric acid ($C_{12}H_{24}O_2$), ~0.5 g[7]
- Effervescent tablet[8]
- Colored drink powder, ~0.2 g

[1] To formulate using anhydrous citric acid ($C_6H_8O_7$) or citric acid monohydrate ($C_6H_8O_7 \cdot H_2O$), refer to the Lab Preparation section.

[2] To formulate using baking soda (sodium bicarbonate, $NaHCO_3$), refer to the Lab Preparation section.

[3] To formulate using anhydrous copper sulfate ($CuSO_4$) or copper sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$), refer to the Lab Preparation section.

[4] To formulate using sodium hydroxide pellets (NaOH), refer to the Lab Preparation section.

[5] To formulate using silver nitrate ($AgNO_3$), refer to the Lab Preparation section.

[6] To formulate using table salt (sodium chloride, NaCl), refer to the Lab Preparation section.

[7] For ideas on how to re-use the lauric acid samples, refer to the Lab Preparation section.

[8] Ensure that the effervescent tablet contains both sodium bicarbonate and citric acid. Do not use antacids with an active ingredient of calcium carbonate because the rate of the reaction occurs too slowly.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Difference between chemical properties and physical properties
- Pure substances (elements and compounds) versus mixtures
- States of matter

Related Labs in this Manual

Labs conceptually related to this one include:

- Conservation of Matter
- Percent Oxygen in Air

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting sensors to the data collection system �(2.1)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Showing and hiding data runs in a graph �(7.1.7)
- Naming a data run �(8.2)
- Measuring the difference between two points in a graph �(9.2)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

Chemistry is the study of matter and how it changes. Matter can be classified as either a pure substance or a mixture. Pure substances are either elements or compounds. Elements are the simplest form of matter because they are made up of only one type of atom. Compounds are made up of two or more different elements bonded together in fixed ratios. Pure substances have fixed physical and chemical properties because they always consist of only one type of chemical material. Mixtures, however, are different than pure substances. Mixtures are made up of two or more pure substances (either elements or compounds). The components of a mixture are not chemically bonded to each other in any way; they are simply in the same container at the same time. If the components of a mixture are distributed evenly throughout, it is called "homogeneous." Conversely, if the components are distributed such that some areas of the mixture are different than others, the mixture is called "heterogeneous."

Matter constantly undergoes change. These changes can be classified as either physical changes or chemical changes. A physical change occurs when a substance's physical appearance changes, but there is no change in the substance's chemical composition. Ice melting is a physical change because the ice has changed from a solid state to a liquid state, but its chemical composition, H_2O, remains the same. A chemical change occurs when one or more new chemical substances are produced. Rusting is an example of a chemical change because iron reacts with oxygen in the air to form a new chemical substance, iron(III) oxide. A chemical change is also called a chemical reaction.

When two substances are mixed, it can be difficult to determine whether or not a new chemical substance was formed because we cannot "see" what is happening at the molecular level. However, there are four types of evidence that, if observed, suggest that a new chemical substance has formed. If one or more of these properties are observed, then it is likely, but not definite, that a chemical reaction has occurred. The four principle types of evidence indicating a chemical reaction has occurred are:

1. A precipitate forms. A precipitate is an insoluble solid that remains suspended in the solution or sinks to the bottom. When a precipitate forms as a result of a reaction, the reaction *mixture* appears cloudy. A precipitate may also be colored.

2. A gas is evolved or absorbed. The gas appears as bubbles that effervesce from the solution. The gases may have a noticeable odor. Although it is difficult to observe, a gas can also be absorbed, such as oxygen being absorbed by iron in the formation of rust. This can be observed using an absolute pressure sensor (as used in the Percent Oxygen in Air experiment).

3. Energy is evolved or absorbed. Energy may be released in the form of light, sound, or heat. A change in heat is indicated by a temperature change during the reaction. A temperature change may be perceived by touch or may require the use of a temperature sensor. Chemical reactions nearly always involve a change in energy because chemical bonds must be broken (which requires energy) and new bonds must be formed (which releases energy). Overall, if the reaction requires more energy than it releases, energy is absorbed from the environment, and the temperature decreases. This is called an endothermic chemical reaction. If more energy is released when bonds are broken than when reformed, the reaction is exothermic.

4. There is a significant change in color. If, upon mixing or heating, the color of a substance significantly changes, or color appears from colorless solutions, a new chemical substance has formed.

It is important that students recognize these are only "pieces of evidence" suggesting that a chemical reaction has occurred. The "evidence" may not always be clear. For instance, a boiling substance might appear to show the evolution of a gas, but boiling is a phase change that does not result in a new substance forming. In another case, pouring a clear solution into a blue solution may yield a light blue solution. This may appear to be a chemical color change but only a physical change occurred as the blue solution was diluted.

In this experiment the students will perform the following chemical reactions:

1. Calcium carbonate reacts with acetic acid (vinegar) to form calcium acetate, carbon dioxide and water.

$$CaCO_3(s) + 2HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + CO_2(g) + H_2O(l)$$

2. Copper sulfate reacts with sodium hydroxide to form a copper(II) hydroxide precipitate and sodium sulfate.

$$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

3. Citric acid reacts with sodium bicarbonate to form water, carbon dioxide, and sodium citrate.

$$H_3C_6H_5O_7(aq) + 3NaHCO_3(aq) \rightarrow 3H_2O(l) + 3CO_2(g) + Na_3C_6H_5O_7(aq)$$

4. An effervescent tablet contains citric acid and sodium bicarbonate. The chemical equation will be the same as Chemical Reaction 3, above.

5. Silver nitrate reacts with sodium chloride to form a silver chloride precipitate and sodium nitrate.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

6. When copper(II) hydroxide is heated it decomposes to form copper(II) oxide and water.

$$Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(g)$$

Pre-Lab Discussion and Activity

Crumble, Tear, and Burn Paper

Engage your students in a discussion about the differences between physical and chemical changes by crumbling, tearing, and then burning a piece of paper. Guide the students into understanding that a chemical change results in a completely new chemical substance being formed, but in a physical change, the chemical substance remains the same and only its appearance changes.

Explain to the students the difference between exothermic and endothermic chemical reactions.

Start a list of evidence to indicate that a new substance has formed. In this section, introduce energy being released and the evolution of a gas as evidence to add to the list.

1. When a piece of paper is crumbled, is it a physical change or a chemical change? How do you know?

It is a physical change because no new chemical substance forms. The paper is still paper. The physical size and shape have changed, but no new substance formed.

2. When a piece of paper is torn, is it a physical change or a chemical change? How do you know?

It is a physical change because there is no new chemical substance formed. The paper is still paper. The physical size and shape have changed, but no new substance formed.

3. When a piece of paper is burned, is it a physical change or a chemical change? How do you know?

It is a chemical change because smoke and ash, two new substances, formed, as well as producing energy as light and heat.

4. What term describes a chemical reaction that releases energy?

A chemical reaction that releases energy is an exothermic chemical reaction.

5. What is the difference between a physical and a chemical change?

A physical change causes the physical appearance to change, but its chemical composition remains the same.

A chemical change forms a completely new chemical substance.

6. What did you observe to provide "evidence" that a new substance was formed?

A gas was evolved, and energy (light and heat) was released.

7. What is another name for a chemical change?

Another name for a chemical change is a “chemical reaction.”

8. How are the terms “reactant” and “product” used when describing a chemical reaction?

Reactants are the starting substances in a chemical reaction, and products are the newly formed substances.

Iodine Solution Demonstrations

Prepare an iodine solution in advance by dissolving 5.0 g of potassium iodide and 1.5 g of iodine in 500 mL of water. Additionally, prepare a starch solution by dissolving 1.0 g cornstarch in 100 mL boiling water. Allow the starch solution to cool before using.

Teacher Tips:

- Use a can of spray starch (available from the laundry aisle of your local grocery store) as a quick and easy alternative to powdered cornstarch. Spray a fine mist directly into a beaker of water while stirring.
- The starch solutions lose their effectiveness if stored for too long. Prepare fresh starch solutions no more than a couple of weeks before performing the pre-lab activity.

Use the iodine solution to engage the students in a discussion about whether or not a color change is evidence of a chemical reaction. Add a few drops of the iodine solution to 10 mL of isopropyl alcohol (available at a local pharmacy). Next, add a few drops of the iodine solution to 10 mL of the cornstarch solution.

9. Is mixing iodine and isopropyl alcohol a physical or chemical change?

This is generally regarded as a physical change. However, the evidence is not clear if the color change is due to the formation of a new substance or whether the color change is due to the dilution of iodine or the presence of iodine impurities. A physical change is further supported in that no gas was evolved, and it does not appear that energy was released.

10. Is mixing iodine and cornstarch a physical or chemical change?

There was a significant color change to dark blue indicating a chemical change. The color is significantly different than either the iodine or the cornstarch and, thus, must be due to the presence of a new chemical substance.

11. Can color change be used as evidence of a chemical reaction?

Yes, but it must be a significant color change.

12. What does the iodine solution test the presence of?

The iodine solution tests for the presence of starch. The students have probably seen this reaction in a biology class.

Precipitation Demonstration

In this section, introduce the formation of a precipitate as the final criteria to determine if a chemical reaction has occurred. Upon completion of this section, add "precipitate forms" to the list of evidence.

Demonstrate the precipitation reaction between sodium carbonate and calcium chloride. Have the students describe the solutions both before and after mixing them.

Put a fast response temperature probe into one of the solutions, and project a Temperature (°C) versus Time (s) graph for the students to see. Start collecting data, and then mix the two solutions. The precipitate will form immediately as a cloudy suspension of small particles. With some time, the solid precipitate will settle to the bottom of the test tube. Use this example and the chemical equation below to demonstrate the meaning of the term "precipitate."

$$Na_2CO_3(aq) + CaCl_2(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$$

13. Describe both the sodium carbonate solution and the calcium chloride solution.

Both solutions are clear, colorless liquids.

14. Does mixing these two solutions result in a physical change or a chemical reaction?

A new solid is formed. Therefore, this must be a chemical reaction. Additionally, the temperature of the solution changed (increased), indicating that energy was released, which provides further evidence that a chemical reaction took place.

15. What is a precipitate?

A precipitate is an insoluble solid that forms when two solutions are mixed.

16. Was this an exothermic or endothermic chemical reaction? How do you know?

It was slightly exothermic. There was a small increase in temperature during the reaction.

17. Does a change in temperature mean that a chemical reaction occurred?

Not always. However, most chemical reactions involve a change in temperature. Temperature changes also occur when solids dissolve. However, dissolving is generally regarded as a physical change.

Energy being released or absorbed is one form of evidence used to predict that a new substance has been formed. It will generally be accompanied by other changes, such as a precipitate formation, color change, or the evolution of a gas.

18. The beaker contains the newly formed precipitate and a liquid. Use the words "mixture" and "compound" to describe the contents of the beaker.

The beaker contains a mixture of at least two compounds, the newly formed precipitate (solid calcium carbonate) and the liquid (a sodium chloride solution).

19. What is the difference between a physical and a chemical change?

A physical change is one that causes the physical appearance of a substance to change, but the chemical composition remains the same. A chemical change is one in which a new chemical substance forms.

20. What are the four types of evidence that indicate a new chemical substance has formed?

Gas is evolved (usually in the form of bubbles).

Energy is released or absorbed (in the form of heat, light, or sound).

There is a significant color change.

A precipitate is formed.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. Prepare 100 mL of 1.0 M citric acid solution. This is enough for 50 lab groups.

Using anhydrous citric acid ($C_6H_8O_7$):

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 19.21 g anhydrous citric acid to the water and swirl to dissolve.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert at least three times.

Using citric acid monohydrate ($C_6H_8O_7{\bullet}H_2O$):

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 21.01 g citric acid monohydrate to the water and swirl to dissolve.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert at least three times.

2. Prepare 100 mL of 1.0 M sodium bicarbonate solution ($NaHCO_3$). This is enough for 50 lab groups.

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 8.40 g sodium bicarbonate to the water and swirl to dissolve.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert at least three times.

3. Prepare 100 mL of 0.5 M copper sulfate ($CuSO_4$) solution. This is enough for 50 lab groups.

Using anhydrous copper sulfate ($CuSO_4$):

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 7.98 g anhydrous $CuSO_4$ to the water and swirl to dissolve.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert at least three times.

Using copper sulfate pentahydrate ($CuSO_4{\bullet}5H_2O$):

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 12.48 g $CuSO_4{\bullet}5H_2O$ to the water and swirl to dissolve.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert at least three times.

4. Prepare 100 mL of 1.0 M sodium hydroxide (NaOH) solution. This is enough for 50 lab groups.

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 4.0 g NaOH to the water and swirl to dissolve. The solution may get warm.
- **c.** Dilute to the mark with distilled water.

d. Cap and invert at least three times.

5. Prepare 100 mL of 0.05 M silver nitrate ($AgNO_3$) solution. This is enough for 50 lab groups.

a. Add approximately 50 mL of distilled water to a 100-mL volumetric flask.

b. Add 0.85 g $AgNO_3$ to the water and swirl to dissolve.

c. Dilute to the mark with distilled water.

d. Cap and invert at least three times.

Teacher Tip: Silver nitrate is a light sensitive compound. If you'll store the solution keep it in a dark bottle or one that has been wrapped in aluminum foil. Avoid getting any of the solution on the skin. If this happens, wash the area with soap and water. While there is minimal threat to health, the solution, if not removed, may considerably blacken and temporarily stain the skin (for 1 to 3 days) when exposed to sunlight. Do not try to remove the stains with scrubbing, detergents, or solvents.

6. Prepare 100 mL of 0.1 M sodium chloride (NaCl) solution. This is enough for 50 lab groups.

a. Add approximately 50 mL of distilled water to a 100-mL volumetric flask.

b. Add 0.58 g NaCl to the water and swirl to dissolve.

c. Dilute to the mark with distilled water.

d. Cap and invert at least three times

Teacher Tips:

- ♦ Lauric acid can be melted by heating it in boiling water. When it is removed from the boiling water, it will form a solid at the bottom of the test tube as it cools, which is very difficult to remove. Save the lauric acid in the test tubes, and have your students reuse them from class to class, as well as from year to year.
- ♦ The waste generated in some parts of the procedure must be collected and disposed of according to your local, state, and federal regulations. Collect the copper(II) hydroxide produced in Chemical Reaction #3), the silver chloride produced in Unknown Change #2, and the copper(II) oxide produced in Unknown Change #3 in separate containers and process them following the procedures outlined by your organization. The remaining waste products may be flushed down the drain with lots of water.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ The silver nitrate ($AgNO_3$) solution may temporarily stain your skin when exposed to bright light. If the solution contacts your skin, wash it with soap and water immediately.
- ♦ Many chemicals used in this lab are hazardous to the environment and should not be disposed of down the drain. Make sure you follow your teacher's instruction on how to properly dispose of the chemicals.
- ♦ Be careful when working with hot water.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

4	2	3	1
Classify the unknown changes as chemical reactions or physical changes based on your observations.	Perform three chemical reactions followed by three physical changes; record the evidence of each type of change occurring.	After you are familiar with the evidence of chemical and physical changes, perform four unknown changes and record your observations.	Pour 150 mL of water into a 250-mL beaker, and start heating it immediately so the water is boiling when you need it later in the lab.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Part 1 – Chemical Reactions

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect the fast response temperature sensor to the data collection system. �(2.1)

3. ☐ Display Temperature (°C) versus Time (s) on a graph. �(7.1.1)

4. ☐ Fill a 250-mL beaker with approximately 150 mL of water.

5. ☐ Place the beaker on the hot plate, and allow the water to boil. The boiling water will be used later.

6. ☐ Use Table 1 below to obtain the name and quantity of Reactant #1 and Reactant #2, which you'll use in each procedure for collecting data of chemical reactions.

Reaction Procedure

7. ☐ Place reactant #1, for reaction #1, in a clean test tube labeled “reaction #1”.

8. ☐ Insert the fast response temperature sensor into the bottom of the test tube containing reactant #1.

9. ☐ Measure out reactant #2.

10. ☐ Describe at least two characteristics you observe for each reactant under their name in Table 1 below.

Table 1: Chemical reactions

Rxn #	Reactant #1 (quantity, name, and description)	Reactant #2 (quantity, name and description)	Description of the Newly Formed Product(s)	Did the Temperature Increase or Decrease?	Evidence of the Chemical Reaction
1	2 mL vinegar (acetic acid) Description: aqueous, clear, colorless	~0.2 g calcium carbonate Description: solid, white, opaque	bubbles, clear, colorless gas	Increase	Gas was produced Temperature changed
2	2 mL 1.0 M citric acid solution Description: aqueous, clear, colorless	2 mL 1.0 M sodium bicarbonate solution Description: aqueous, clear, colorless	bubbles, clear, colorless gas	Decrease	Gas was produced Temperature changed
3	2 mL 0.5 M copper(II) sulfate solution Description: aqueous, clear, blue	2 mL 1.0 M sodium hydroxide solution Description: aqueous, clear, colorless	blue, solid, cloudy	Increase	Precipitate was formed Temperature changed

Collect Data

11. ☐ Start recording data. �(6.2)

12. ☐ Add reactant #2 to the test tube containing reactant #1.

13. ☐ Adjust the scale of the graphs so that you clearly see the changes in temperature that are taking place. �(7.1.2)

14. ☐ Stop data recording when the reactants have thoroughly mixed and the temperature has stabilized. �(6.2)

15. ☐ Name the data run “reaction 1”. � (8.2)

16. ☐ Record a description of the newly formed substance, whether the temperature increased or decreased, and the evidence for the chemical reaction you observed in Table 1 above.

17. ☐ Remove the fast response temperature sensor from the test tube, and set the test tube and its contents aside for possible reuse later in the experiment.

18. ☐ Thoroughly clean the fast response temperature sensor by rinsing it several times with distilled water.

19. ☐ Repeat the steps in the Reaction Procedure and Collect Data sections for reaction #2, and again for reaction #3.

20. ☐ What is a reactant in a chemical reaction?

A reactant is a starting substance in a chemical reaction.

21. ☐ What is a product in a chemical reaction?

A product is a newly formed substance in a chemical reaction.

22. ☐ Colorless gasses cannot be “seen.” How is it possible to know if a gas is being evolved in an aqueous solution?

Bubbles indicate that a gas is being released.

Part 2 – Physical Changes

Collect Data

Physical change #1

23. ☐ Observe the water you started heating at the beginning of the lab, and record your observations for Physical Change #1 in Table 2 below. (Continue to boil the water because you will use it to heat chemicals in the next part of the lab. You may need to add more water to replace what has evaporated.)

Physical change #2

24. ☐ Add approximately 0.2 g colored drink powder to a dry, clean test tube.

25. ☐ Add 5 mL of water.

26. ☐ Record your observations for Physical Change #2 in Table 2 below.

Physical change #3

27. ☐ Break your effervescent tablet into 3 or 4 pieces.

28. ☐ Record your observations for Physical Change #3 in Table 2 below.

29. ☐ Save the pieces to use later in the lab.

30. ☐ Write at least one description in Table 2 below for how the appearance of each chemical substance changed macroscopically even though the molecules making up the substance are still the same.

Table 2: Physical changes

Physical Change #	Physical Change	Description Before the Change	Changes in Appearance (even though the chemical substance is still the same)
1	Heat water (H_2O) until it is boiling	Clear, colorless, liquid	Water is still H_2O, but it is now a gas.
2	Dissolve colored drink powder in water	Drink powder: red color, solid Water: clear, colorless, liquid	Water and drink powder are both still present, but they are mixed together. This mixing makes the water appear colored and makes the drink powder appear to be a liquid.
3	Break the effervescent tablet into 3 or 4 pieces	White, solid, opaque	The size of the effervescent tablet changed, but it is still the same chemical substance.

31. ☐ How are physical changes and chemical changes different from each other?

A chemical change occurs when a new chemical substance is formed through a chemical reaction. A physical change occurs when there is a change in the substance's physical appearance, but the chemical substance remains the same.

Part 3 – Unknown Changes

Collect Data

32. ☐ Record the description of each reactant in Table 3 below before you mix them. A description of "Reactant #2" in Unknown Changes #3 and #4 is not required because it is heat.

Table 3: Unknown changes

Unknown Change #	Reactant #1	Reactant #2	Observations	New Chemical Substance?
1	100 mL of water Description: liquid, colorless, clear	3 or 4 pieces of the effervescent tablet Description: solid, white, opaque	Bubbles were produced Colorless gas Effervescent tablet "disappeared" Temperature decreased	Yes Gas evolved
2	2 mL 0.05 M silver nitrate solution Description: aqueous, colorless, clear	2 mL 0.1 M sodium chloride solution Description: clear, colorless, aqueous	Solid white precipitate formed Cloudy solution Temperature increased	Yes Precipitate formed
3	Blue precipitate from chemical reaction #3 (copper(II) hydroxide) Description: solid, blue	Heat	Blue precipitate changed to a black color Liquid is still present	Yes Color Change
4	Test tube containing ~0.5 g of lauric acid Description: solid, white, waxy	Heat	Solid turned to a liquid Liquid is clear and colorless	No

Unknown Change #1

33. ☐ Pour 100 mL into a clean 250-mL beaker.

34. ☐ Insert the fast response temperature sensor into the beaker.

35. ☐ Start recording data. �(6.2)

36. ☐ Add the pieces of the effervescent tablet previously set aside.

37. ☐ Adjust the scale of the graph so you can see the temperature change taking place. �(7.1.2)

38. ☐ Stop recording data when the temperature has stabilized. �(6.2)

39. ☐ Name the data run “unknown 1”. �(8.2)

40. ☐ List at least three observations, including any temperature changes, in Table 3 above.

41. ☐ Record in Table 3 whether or not a new chemical substance was formed. If a new chemical substance was formed, list the evidence you can use to support your conclusion.

42. ☐ Remove the fast response temperature sensor, and rinse it thoroughly with distilled water.

Unknown change #2

43. ☐ Pour 2 mL of 0.05 M silver nitrate solution into a test tube.

44. ☐ Insert the fast response temperature sensor into the test tube.

45. ☐ Start recording data. �(6.2)

46. ☐ Add 2 mL of 0.1 M sodium chloride solution.

47. ☐ Adjust the scale of the graph so you can see the temperature change taking place. �(7.1.2)

48. ☐ Stop recording data when the temperature has stabilized. �(6.2)

49. ☐ Name the data run “unknown 2”. �(8.2)

50. ☐ List at least three observations including any temperature changes in Table 3 above.

51. ☐ Record whether or not a new chemical substance was formed in the data table above. If a new chemical substance was formed, list the evidence you can use to support your conclusion.

52. ☐ Remove the fast response temperature sensor, and rinse it thoroughly with distilled water.

Unknown change #3 and #4

53. ☐ Use a test tube holder to place the test tube containing the copper(II) hydroxide precipitate formed in chemical reaction #3 into the boiling water.

54. ☐ Use a test tube holder to place the test tube containing the lauric acid into the boiling water.

55. ☐ Allow the test tubes to sit in the boiling water until they have completely changed (about 3 to 5 minutes).

56. ☐ List at least two observations for the changes taking place in each test tube in Table 3 above.

57. ☐ Record whether or not a new chemical substance was formed in Table 3 above. If a new chemical substance was formed, list the evidence you can use to support your conclusion.

58. ☐ Remove the test tubes from the boiling water using the test tube holder.

59. ☐ Turn off the hot plate.

60. ☐ Save your data file and clean up the lab station according to the teacher's instructions, especially those concerning the disposal of used chemicals. �(11.1)

CAUTION: Be sure the hot plate and hot water have cooled before handling them.

Data Analysis

1. ☐ Determine the change in temperature for each experiment in which you collected temperature data.

a. Display the data run you plan to analyze. �(7.1.7)

b. Measure the difference between the initial temperature and final temperature. �(9.2)

c. Record the temperature change in Table 4 below.

Table 4: Identifying changes as endothermic or exothermic based on the observed change in temperature

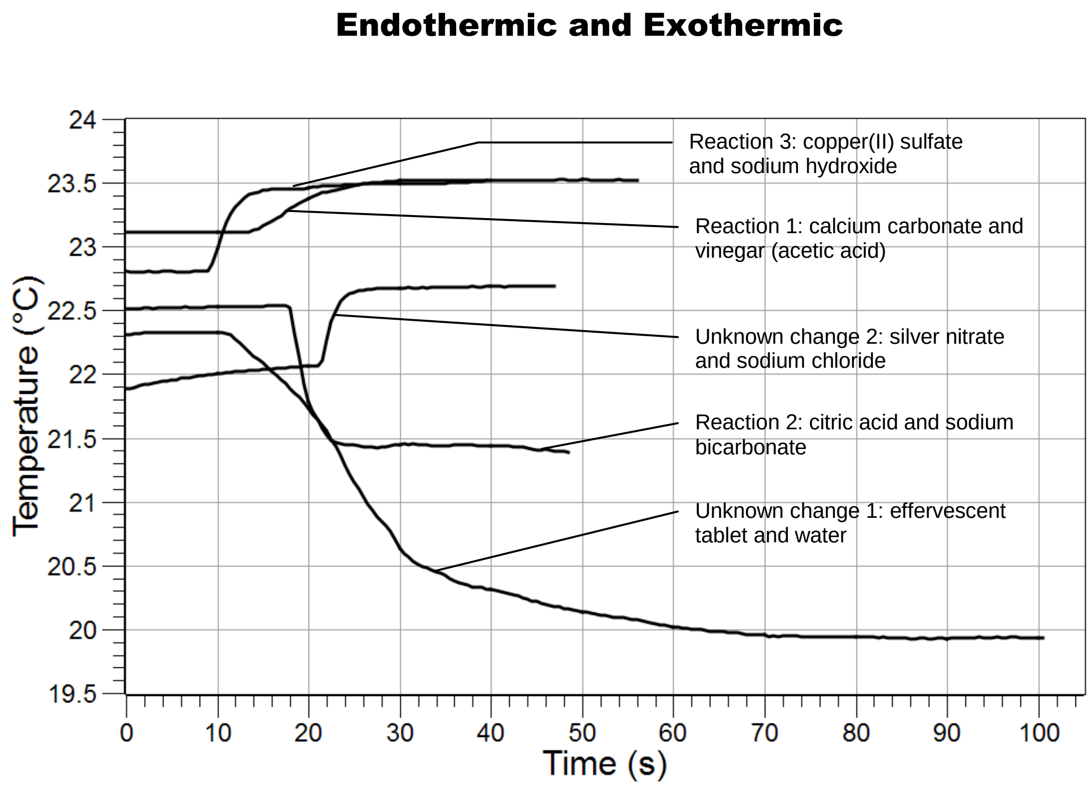
Change	Change in Temperature (°C)	Exothermic or Endothermic?
Chemical Reaction #1: calcium carbonate and vinegar (acetic acid)	0.4	exothermic
Chemical Reaction #2: citric acid and sodium bicarbonate	–1.1	endothermic
Chemical Reaction #3: copper(II) sulfate and sodium hydroxide	0.7	exothermic
Unknown Change #1: effervescent tablet and water	–2.4	endothermic
Unknown Change #2: silver nitrate and sodium chloride	0.6	exothermic

2. ☐ Determine whether the change was exothermic or endothermic using the calculated temperature changes determined above. Record your answers in Table 4 above.

3. ☐ Create a graph of data with all five runs of data displayed on your data collection system. �(7.1.3)

Note: Not all data collection systems will display all five runs of data on one set of axes. If your data collection system cannot display all the data runs, multiple graphs may be used.

4. ☐ Sketch or print a graph of Temperature (°C) versus Time (s) on one set of axes. Label each run of data as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Analysis Questions

1. Identify each "unknown change" as a physical change or a chemical reaction, and include evidence to support your answers in the table below.

#	Unknown Change	Physical Change or Chemical Reaction	Evidence
1	Effervescent tablet and water	Chemical Reaction	A new substance, a gas, was formed.
2	Silver nitrate solution and sodium chloride solution	Chemical Reaction	A new substance, a white precipitate, was formed.
3	Heating copper(II) hydroxide	Chemical Reaction	A new substance, with a black color, was formed.
4	Heating lauric acid	Physical Change	No new substance was formed, the lauric acid melted.

2. Were unknown change #3 and unknown change #4 endothermic or exothermic? Explain your reasoning.

Both of these processes were endothermic because they both absorbed heat from the boiling water to occur.

3. What is the difference between a physical change and a chemical reaction?

A physical change involves a change in the substance's physical appearance, but remains the same chemical substance.

A chemical reaction is a change in which a new chemical substance is formed.

4. What are the four main types of evidence that indicate a chemical reaction has occurred?

The evidence indicating that a chemical reaction has occurred: the evolution of a gas (seen as bubbles), the release or absorption of energy (as heat, light, or sound), the formation of a precipitate, and a significant change in the color of the substance.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Is mixing salt with water an example of a physical change or chemical reaction? Explain your reasoning.

Dissolving salt in water is a physical change. The salt is still the same substance, but it is just broken into tiny pieces (its ions).

2. List two examples in which a temperature change occurs, but no new substance is formed.

Any phase change of a substance is sufficient: melting, boiling, freezing, and condensing.

3. When a nail becomes rusty, is this an example of a physical change or a chemical reaction? Explain your reasoning.

Rusting is a chemical reaction. The nail (iron) reacts with oxygen in the atmosphere to form a new substance, rust (iron(III) oxide). Evidence for this reaction is a color change. Rust also absorbs a gas, which can be observed by using a pressure sensor, as in the Percent Oxygen in Air experiment.

4. When grass grows, is this an example of a physical change or a chemical reaction? Explain your reasoning.

Chemical reactions are involved when grass grows. Photosynthesis requires that energy from the sun is absorbed (an endothermic process). Gas is also both absorbed (carbon dioxide) and released (oxygen).

Photosynthesis: $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$

5. Is opening a can of soda an example of a physical change or a chemical reaction? Explain your reasoning.

This is an example of a physical change; no new substances are formed.

Multiple Choice Questions

1. In all chemical reactions, ________ turn into ___________.

- **A.** Products; Reactants
- **B.** Molecules; Atoms
- **C.** **Reactants; Products**
- **D.** Atoms; Elements

2. The burning of wood to form soot is an example of a _____________________ change.

- **A.** Physical
- **B.** Slow
- **C.** Fast
- **D.** **Chemical**

3. Which of the following indicates a chemical reaction has occurred?

- **A.** The change is very fast
- **B.** **A precipitate forms**
- **C.** The state of matter changes
- **D.** A dark orange solution turns light orange

4. A chemical reaction that absorbs energy is called a(n) ______________ reaction?

- **A.** **Endothermic**
- **B.** Exothermic
- **C.** Balanced
- **D.** Complete

5. Grinding a large crystal of rock candy into small pieces is an example of a

- **A.** **Physical change**
- **B.** Chemical change
- **C.** Exothermic change
- **D.** Endothermic change

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Matter can undergo both chemical and physical changes. In a **chemical change**, one substance turns into a completely different substance. A chemical change is also called a chemical **reaction**. A **physical change** occurs when a substances changes its physical appearance, but not its chemical identity. Identifying whether a process involves a chemical reaction or a physical change can be difficult because we cannot "see" at the **molecular** level to know whether or not the actual atoms have rearranged. Instead, we must rely on **macroscopic** observations. There are four main types of evidence that suggest a new chemical substance has formed. The formation of a precipitate, the evolution of a **gas**, a significant **color** change, and a change in energy all indicate that a chemical reaction has occurred. A **precipitate** is a solid that forms when two solutions are mixed.

2. A **chemical reaction** always involves a new substance being formed. Therefore, a chemical reaction has starting substances, called **reactants**, and newly formed substances called **products.** When a chemical reaction occurs, the bonds within the original substances are broken apart, the atoms rearrange, and new bonds are formed, creating completely new substances. Energy is required to break the bonds, and energy is released when the bonds are formed. In most chemical reactions, the energy required to break the bonds does not equal the energy released when the new bonds are formed. Therefore, most chemical reactions either absorb energy or release energy. Chemical reactions that absorb energy from their environment result in the temperature decreasing and are called **endothermic** reactions. Chemical reactions that release energy into their environment cause the temperature to increase and are called **exothermic** reactions.

Extended Inquiry Suggestions

Bake bread with the students, and identify each physical change and chemical reaction as you go through the process.

Have students pick a chemical reaction of biological or industrial importance, and describe what makes the process a chemical reaction.

Use a carbon dioxide sensor to monitor the amount of carbon dioxide present, and prove that a new substance is created during either the effervescent tablet experiment or the vinegar-calcium carbonate reaction.

18. Stoichiometry

Objectives

Determine the mole ratio between the reactants sodium hypochlorite and sodium thiosulfate. Through this investigation, students:

- Review the evidence of chemical reactions
- Describe the connection between coefficients in chemical reactions and the amount of product formed
- Explain the term limiting reactant

Procedural Overview

Students conduct the following procedures:

- React sodium hypochlorite and sodium thiosulfate in various proportions and measure the resulting temperature changes
- Analyze data to determine the optimal mole ratio between the two reactants
- Determine the limiting reactant for specified reaction variations

Time Requirement

- Preparation time — 20 minutes
- Pre-lab discussion and activity — 20 minutes
- Lab activity — 50 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Temperature sensor[1]
- Graduated cylinder (2), 10-mL
- Graduated cylinder (2), 50-mL or 100-mL
- Transfer pipet (2)
- Test tube (7), 20-mm x 150-mm
- Test tube rack
- Wash bottle filled with water
- Waste container
- 0.5 M Sodium hypochlorite (NaClO), 35 to 40 mL[2]
- 0.5 M Sodium thiosulfate ($Na_2S_2O_3$), in 0.2 M sodium hydroxide (NaOH), 35 to 40 mL[3]

[1] Either the fast response or stainless steel temperature sensor is suitable.

[2] To formulate the solution using household laundry bleach (4 to 6% sodium hypochlorite, NaClO), refer to the Lab Preparation section.

[3] To formulate the solution using either anhydrous sodium thiosulfate ($Na_2S_2O_3$) or sodium thiosulfate pentahydride ($Na_2S_2O_3 \cdot 5H2O$) and either solid sodium hydroxide (NaOH) or 1.0 M sodium hydroxide, refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Evidence of chemical reactions
- ♦ Moles
- ♦ Molar ratios
- ♦ Balancing chemical equations
- ♦ Law of conservation of matter
- ♦ Endothermic and exothermic reactions

Related Labs in This Guide

Labs conceptually related to this one include:

- ♦ Evidence of a Chemical Reaction
- ♦ Conservation of Matter
- ♦ Heats of Reaction and Solution

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- ♦ Starting a new experiment on the data collection system �(1.2)
- ♦ Connecting a sensor to the data collection system �(2.1)
- ♦ Recording a run of data �(6.2)
- ♦ Displaying data on a graph �(7.1.1)
- ♦ Adjusting the scale of a graph �(7.1.2)
- ♦ Finding the coordinates of a point in a graph �(9.1)
- ♦ Naming data runs �(8.2)
- ♦ Saving you experiment �(11.1)

Background

Stoichiometry refers to the relationships amongst the reactants and products in a chemical reaction. The quantities of reactants and products are calculated using the coefficients in the balanced chemical equation. The coefficients indicate the ideal mole ratio between reactants and products, and they satisfy the law of conservation of matter. When the ideal amount of each reactant is combined, the maximum amount of product is formed with the least amount of unused or wasted excess material.

Stoichiometric calculations are important because they allow chemists to control the amount of reactants and products to ensure that reactions are carried out safely, economically, and profitably. For example, in synthesizing a medicine, a pharmaceutical company needs to begin with the correct amounts of reactants to ensure that the required amount of medicine will be produced. Too little of even one reactant results in not enough medicine being produced. The reactant that runs out first is called the limiting reactant because it limits the amount of product that is formed.

If a reaction's balanced chemical equation is not known, it must be determined experimentally by combining the reactants in various proportions and measuring the extent to which products form. This is known as Job's Method of Continuous Variations. Product formation can be measured by selecting an observable property known to indicate that a chemical reaction has taken place. These properties include: the formation of a precipitate, formation of a gas, change in temperature, or a change in color. The combination that produces the greatest change in the property being observed is the ideal ratio.

In this activity, students react sodium hypochlorite, $NaClO$, with sodium thiosulfate, $Na_2S_2O_3$ (in a 0.2 M sodium hydroxide solution). In this reaction, there is no significant color change, no gas produced, and no precipitate formed. The reaction is, however, exothermic, and heat energy will be generated as one of the reaction products according to the equation below.

$$4NaClO(aq) + Na_2S_2O_3(aq) + 2NaOH(aq) \rightarrow 2Na_2SO_4(aq) + 4NaCl(aq) + H_2O(l) + \text{heat}$$

The reaction conditions that produce the most heat mark the ideal mole ratio. The amount of heat produced is measured by the change in temperature ΔT of the system. This is possible because the reaction conditions are controlled such that the total volume of water is the same in each container. The energy released is absorbed by the same amount of water, therefore the heat given off is proportional to the change in temperature ($q = mC\Delta T$, since mC is constant, q is directly proportional to ΔT).

Pre-Lab Discussion and Activity

Chemical Equations and the Law of Conservation of Matter

Write a balanced chemical equation on the board such as the synthesis of water.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Use this equation to review the components of a chemical equation and to review the law of conservation of matter.

1. Chemical equations contain a variety of numbers, letters, and symbols. Numbers are used as coefficients and subscripts. What are coefficients? What are subscripts?

Coefficients are the numbers that precede the chemical symbol or formula and are used to indicate more than one mole of a particular chemical species.

Subscripts are the numbers written after the atomic symbols and indicate the number of each atom type within the chemical formula.

2. What is the law of conservation of matter? How does the chemical equation illustrate this law?

The law of conservation of matter states that matter can neither be created nor destroyed only changed. In a chemical equation, the total number of each type of element must be the same on the reactant side as on the product side. The equation below illustrates the law of conservation of mass because there are the same number of hydrogen atoms (4) and the same number of oxygen atoms (2) both before the reaction (as reactants) and after the reaction (as products).

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Reactants	Type of atom	Products
$2H_2$ = 4H	H	$2H_2O$ = 4H
O_2 = 2O	O	$2H_2O$ = 2O

Stoichiometry – Peanut Butter and Jelly

Write the term *stoichiometry* on the board and define it as the topic in chemistry that deals with calculating the amount of reactants required or products produced in a chemical reaction. Use an actual example, such as assembling parts to make a consumer product or following a recipe when cooking, to discuss the concept of stoichiometry.

3. Stoichiometry might sound complicated, but it is very similar to making consumer products by assembling parts. What consumer products do you buy that consist of parts that have to be put together?

Students might suggest consumer products such as cell phones, bicycles, cars, and MP3 players. Food examples may also be used.

Use a peanut butter and jelly (PBJ) sandwich as an example and write it on the board in the form of a chemical equation.

2bread + 1PB + 1jelly → 1PBJ sandwich

Explain that specific amounts of reactants are required to produce certain amounts of products. In this case, 2 slices of bread are required to make 1 PBJ sandwich. So the ratio of bread to sandwich is 2:1.

4. What is the ratio between peanut butter (PB) and PBJ sandwich? What is the ratio of bread to PB?

The ratio of peanut butter (PB) to PBJ sandwich is 1:1.

The ratio of bread to PB is 2:1.

Emphasize that the balanced equation is the ideal ratio of bread to PB to jelly. It is ideal because all of the reactants are used to produce the product(s). Nothing is wasted.

5. How many PBJ sandwiches can you make if you start with four slices of bread, two spoonfuls of PB, and one spoonful of jelly? What do you run out of?

Only one sandwich can be made because you run out of jelly.

Write the term *limiting reactant* on the board. Explain that jelly is the limiting reactant. It limits the number of sandwiches that can be made. Continue changing the starting conditions (amount of each reactant) and asking how that effects the output.

Stoichiometry – Chemicals

Transition back to chemistry, by explaining that stoichiometry in chemistry is exactly the same. The only reason it seems harder is because we cannot see the individual molecules. Instead we need to observe a macroscopic property that indicates product formation (a chemical reaction has occurred). React a solution of citric acid with baking soda. To do this, take an aqueous solution of citric acid and place a temperature probe in the solution. Begin to collect data and project the data being collected for the class to see. Hold up a spoonful of baking soda (sodium bicarbonate). Add the baking soda to the citric acid solution. Discuss whether or not a chemical reaction occurred and what products were formed.

6. What macroscopic properties are evidence of a chemical reaction occurring?

The formation of a gas, a color change, the formation of a precipitate, and a change in temperature are evidence that a chemical reaction has occurred.

7. How will you know if these two substances (citric acid and baking soda) undergo a chemical reaction when they are mixed?

In a chemical reaction a new substance must be formed. Formation of a gas, color change, formation of a precipitate, and a change in temperature are all clues that suggest a new substance has been formed.

8. Did a chemical reaction take place? How do you know?

Yes, a reaction occurred. A new substance, a gas (carbon dioxide), was formed. There was also a decrease in temperature.

9. How could you increase the amount of product formed? How would you know if that suggestion worked?

To increase the amount of product formed, you could increase the amount of both reactants, but this would still cause one of them to be left over (excess reactant) while the other was completely consumed (limiting reactant). By selectively increasing the amount of just the limiting reactant, the amount of product can be increased without wasting excess reactant. By working with the proper ideal mole ratio, the amount of product can be maximized while minimizing the amount of excess reactant wasted.

When tested, an increase in the amount of product formed would be observed by an increase in the amount of gas formed as well as a larger change in temperature.

Lab Overview

Write the following equation on the board:

____NaClO + ___$Na_2S_2O_3$ → products

Demonstrate the reaction that the students will perform by mixing together sodium hypochlorite and sodium thiosulfate for the class to see (be sure to include a temperature probe and to display the data). Ask the students to explain which product will be easy be measure. Show the students the different combinations of reactant volumes they will be testing. Explain to the students that having different volumes of Solution A and B is similar to having different amounts of bread, PB, and jelly. End the discussion by asking the students to make predictions about the results they will find.

Table: Volumes of Solution A and Solution B that will be reacted

Reaction Number	1	2	3	4	5	6	7
Volume of Solution A: 0.5 M NaClO	10 mL	8 mL	6 mL	5 mL	4 mL	2 mL	0 mL
Volume of Solution B: 0.5 M $Na_2S_2O_3$	0 mL	2 mL	4 mL	5 mL	6 mL	8 mL	10 mL

10. In this experiment, sodium hypochlorite (NaClO) and sodium thiosulfate ($Na_2S_2O_3$) will be reacted. The challenge is to determine their ideal reaction ratio. How will you know what the ideal ratio is?

The ideal ratio will be the combination of reactant volumes that produces the greatest number of products. The amount of products formed will be determined by measuring one of the macroscopic properties that indicates that a chemical reaction has occurred.

11. What macroscopic property will you be able to measure? How will you know when the ideal mole ratio is met?

No gas or precipitate is formed and there is no change in color. Therefore, there must be a change in temperature.

The ideal mole ratio is the combination of reactants that produce the greatest formation of products, determined by the greatest change in temperature.

12. Are there any combinations in the table above in which you would expect no chemical reaction to occur? Explain.

There will be no chemical change in reactions 1 and 7 because there is only one reactant in each test tube.

13. Which combination do you think will be the ideal mole ratio (produce the greatest temperature change)? Why?

Answers will vary because students are not expected to know this before they perform the experiment. Many students will predict that equal volumes of each solution will produce the highest temperature. The correct answer is reaction 2, because there is a 4:1 mole ratio between sodium hypochlorite and sodium thiosulfate.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. Prepare 1000 mL of 0.5 M sodium hypochlorite (NaClO) solution by following the steps below. This is enough for 25 lab groups.

a. Determine the volume of bleach required using the table below. The percent of sodium hypochlorite may be found written on the product label.

Table: Amount of bleach required to make the sodium hypochlorite solution

Percent Sodium Hypochlorite Your Bleach Contains	Amount of Bleach to Use (mL)
4%	846 mL
5%	677 mL
5.25%	645 mL
6%	564 mL
other %	see Teacher Tips below

b. Measure the amount of bleach required and add enough distilled water to the bleach to create 1000 mL of solution.

Teacher Tips:

- Do not use color-safe bleach.
- A new container of bleach works best because the percent of sodium hypochlorite gradually decreases with storage.
- Determining the volume of bleach required to make a 0.5 M hypochlorite solution by using the following steps:

a. The volume of bleach required is determined by converting the percent sodium hypochlorite into molarity using the density of bleach (1.10g/mL) and the molar mass of sodium hypochlorite (74.44 g/mol).

$$\text{Molarity (mol/L)} = \left(\frac{\%\ \text{NaClO}}{100}\right)\left(\frac{1.10\ \text{g}}{\text{mL}}\right)\left(\frac{1\ \text{mol}}{74.44\ \text{g}}\right)\left(\frac{1000\ \text{mL}}{1\ \text{L}}\right)$$

b. The formula $M_1V_1 = M_2V_2$ is then used to calculate the volume needed. To determine the volume of bleach to be diluted for 1.0 L of a 0.5 M solution, substitute the molarity calculated above into the following equation:

$$\text{volume (mL) of bleach} = \frac{(1000\ \text{mL})(0.5\ \text{M})}{\text{molarity of NaClO}}$$

2. Prepare 1000 mL of 0.5 M sodium thiosulfate ($Na_2S_2O_3$) in a 0.2 M sodium hydroxide (NaOH) solution by following the steps below:

a. Weigh 79.0 g anhydrous $Na_2S_2O_3$ or 124.1 g $Na_2S_2O_3 \cdot 5H_2O$.

b. Measure out 8.0 g NaOH or 200 mL of 1 M NaOH solution.

c. Mix the reagents in steps a and b together and add enough distilled water to create 1000 mL of solution.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Sodium hypochlorite (NaClO) is the active ingredient in household bleach. Beware of spills onto clothing or skin.
- ♦ The solutions in this lab are alkaline, and harmful to skin and eyes. If the solutions come in contact with the skin or eyes, flush with a copious amount of water.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

4	3	1	5	2
Begin data collection just before Solution B ($Na_2S_2O_3$) is added to a test tube. Stop when the temperature stabilizes.	Pour the indicated amount of Solution A (NaClO) into each labeled test tube.	Obtain equipment, glassware, and solutions.	Record the highest temperature produced. Use this to determine the mole ratio between the reactants.	Label all solutions, graduated cylinders, and test tubes.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Use a large graduated cylinder (50-mL or 100-mL) to obtain 35 to 40 mL of 0.5 M sodium hypochlorite (NaClO). Label the graduated cylinder "Solution A".

2. ☐ Describe the sodium hypochlorite solution.

The sodium hypochlorite solution is a clear, slightly yellow liquid.

3. ☐ Use the other large graduated cylinder (50-mL or 100-mL) to obtain 35 to 40 mL of 0.5 M sodium thiosulfate ($Na_2S_2O_3$). Label the graduated cylinder "Solution B".

4. ☐ Describe the sodium thiosulfate solution.

The sodium thiosulfate solution is a clear, colorless liquid.

5. ☐ Label two 10-mL graduated cylinders. The first cylinder should be labeled "Solution A" and the second cylinder should be labeled "solution B".

6. ☐ Label two transfer pipets. The first pipet should be labeled "Solution A" and the second pipet should be labeled "Solution B".

7. ☐ Why is it necessary to label all the glassware and solutions so carefully?

Proper lab safety requires that all solutions be labeled.

Labeling is also necessary to avoid contamination and ensure reproducible results.

8. ☐ Using Table 1 below, label each test tube with the reaction number and place it in the test tube rack.

Table 1: Volumes of Solution A and Solution B reacted and the highest temperature produced

Reaction Number	**1**	**2**	**3**	**4**	**5**	**6**	**7**
Volume of Solution A: 0.5 M NaClO (mL)	10	8	6	5	4	2	0
Volume of Solution B: 0.5M $Na_2S_2O_3$	0	2	4	5	6	8	10
Highest temperature observed (°C)	23.8	32.9	30.3	29.3	28.0	26.0	23.9

9. ☐ Identify the independent and dependent variables of the experiment and their units.

The independent variable is the amount of each reactant being mixed expressed in milliliters (mL).

The dependent variable is the maximum temperature reached expressed in degrees Celsius (°C).

10. ☐ Predict which combination of reactant volumes is the ideal mole ratio (produces the largest change in temperature). Explain your prediction.

Answers will vary because students are not expected to know this before they perform the experiment. Many students will predict that equal volumes of each solution will produce the highest temperature. The correct answer is reaction 2, because there is a 4:1 mole ratio between sodium hypochlorite and sodium thiosulfate.

11. ☐ Using the 10-mL graduated cylinder labeled “Solution A” and the transfer pipet labeled “Solution A”, measure and pour 10 mL of the 0.5 M NaClO solution into the test tube labeled for “reaction 1”.

12. ☐ Continue to fill each test tube with the proper amount of Solution A until all the test tubes contain Solution A (except test tube 7, which has 0 mL of Solution A).

13. ☐ Start a new experiment on your data collection system. �(1.2)

14. ☐ Connect the temperature sensor to the data collection system. �(2.1)

15. ☐ Display Temperature (°C) versus Time (s) on a graph. �(7.1.1)

Collect Data

16. ☐ Place the temperature sensor into test tube 1 and start recording data. �(6.2)

17. ☐ Adjust the scale of the axes as necessary to see any temperature changes. �(7.1.2)

18. ☐ When the temperature stabilizes, stop recording data. �(6.2)

19. ☐ Determine the highest temperature achieved and record it in Table 1 above, for Reaction 1. �(9.1)

20. ☐ Name the data run “Reaction 1”. �(8.2)

21. ☐ Remove the temperature sensor and rinse it with water over the waste container. Place the clean temperature sensor into test tube 2 and start recording data. �(6.2)

22. ☐ Using the 10-mL graduated cylinder labeled “Solution B” and the transfer pipet labeled “Solution B”, measure and pour 2 mL of Solution B into test tube 2 which is already recording data.

23. ☐ Adjust the scale of the axes as necessary to see any temperature changes. �(7.1.2)

24. ☐ When the temperature has stabilized, stop recording data. �(6.2)

25. ☐ Determine the highest temperature achieved and record it in the Table 1 for reaction 2.�(9.1)

26. ☐ Name the data run “Reaction 2”. �(8.2)

27. ☐ Why is it necessary to start recording data before Solution B is added to Solution A?

The reaction occurs very quickly. If you start recording data after the reactants have been mixed, you might not observe the highest temperature reached.

28. ☐ According to your data, is this an endothermic or exothermic reaction? How do you know?

The reaction is exothermic. The temperature increased, and therefore, heat was being released as the chemicals reacted.

29. ☐ Continue recording data for the remaining reactions. For each reaction:

a. Place a clean temperature sensor into the test tube and start recording data. �(6.2)

b. Add the correct amount of Solution B.

c. Adjust the scale of the axes as necessary to see any temperature changes. �(7.1.2)

d. When the temperature stabilizes, stop recording data. �(6.2)

e. Find the highest temperature reached and record it in Table 1 above. �(9.1)

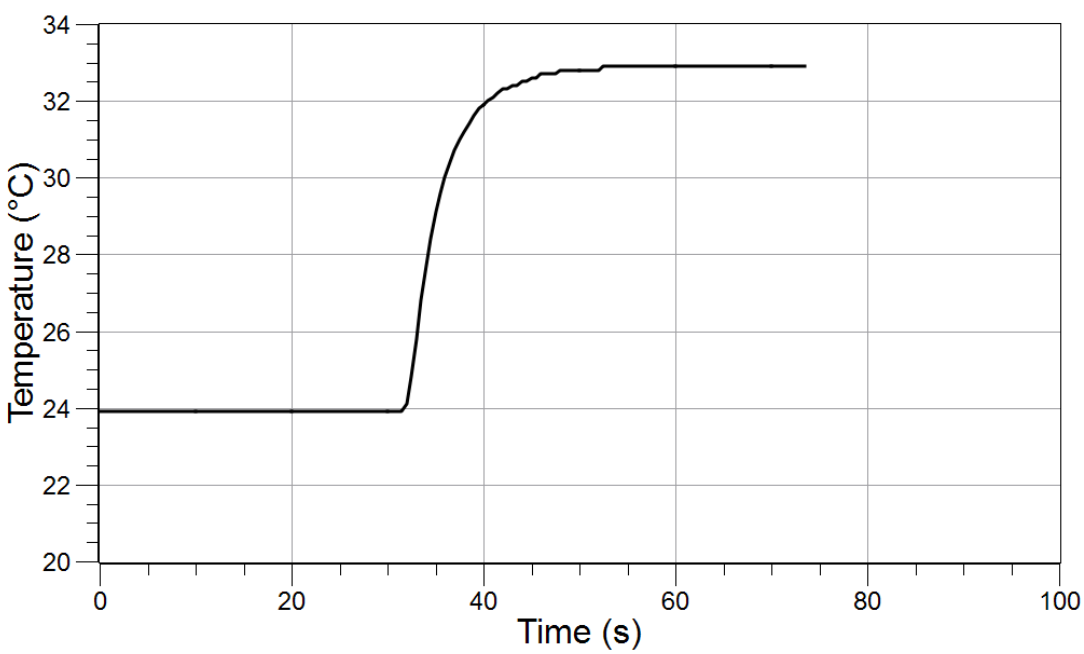
f. Name the data run using the correct reaction number. �(8.2)

30. ☐ Save the data file and clean up the lab station according to the teacher's instructions. �(11.1)

Sample Data

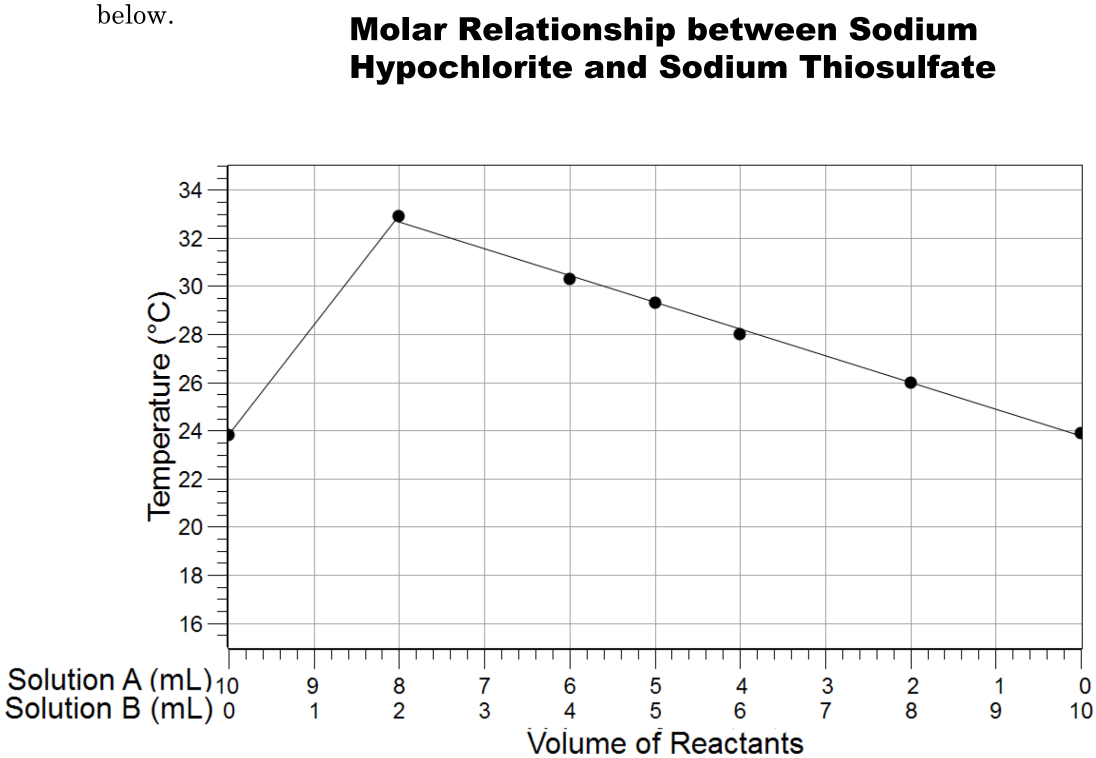
Each run the students collect will look similar to the graph below.

Data Run for Reaction 2



Data Analysis

1. ☐ Plot the graph of Temperatures (°C) versus Volumes of Reactants (mL) on the axes below.



2. ☐ On the graph above, draw two different best fit lines. Draw the first best fit line through the data points where the maximum temperature of the reaction is increasing and the second best fit line through the data points where the maximum temperature of the reaction is decreasing.

3. ☐ Which volumes of reactant produced the highest temperature? List the volume of the each reactant below.

Solution A: 0.5 M $NaClO$ ____________ 8 mL

Solution B: 0.5 M $Na_2S_2O_3$ ____________ 2 mL

4. ☐ Find the number of moles of each reactant for the reaction that produced the highest temperature. Show your work and record your answer in Table 2 below.

Table 2: Moles of each reactant that produced the highest temperature

Reactant	Show Your Work Here	Moles of Each Reactant (mol)
Solution A: 0.5 M $NaClO$	8 mL : $(0.008\ L)\left(\frac{0.5\ mol}{L}\right) = 0.004$	0.004
Solution B: 0.5 M $Na_2S_2O_3$	2 mL : $(0.002\ L)\left(\frac{0.5\ mol}{L}\right) = 0.001$	0.001

5. ☐ What values of x and y express the mole ratio between sodium hypochlorite and sodium thiosulfate in the following chemical equation: $xNaClO + yNa_2S_2O_3 \rightarrow$ products?

x=4; y=1

Analysis Questions

1. How was the ideal mole ratio determined from the data collected?

The ideal mole ratio was determined by identifying the combination of reactant volumes that produced the most products. The reaction is exothermic, therefore, the combination that produced the highest temperature was the one that gave the ideal mole ratio. The volumes of each reactant used were converted to moles and then compared to each other in order to determine the mole ratio.

2. What happens when two reactants are combined in amounts not consistent with their ideal mole ratio?

One reactant will be completely consumed (limiting reactant) and the other reactant will be left over (excess reactant).

3. Explain the difference between a limiting reactant and an excess reactant?

A limiting reactant is the reactant that is used up first in a chemical reaction. An excess reactant is a reactant that still remains present after the reaction has occurred.

4. What was the limiting reactant in Reaction 5 (4 mL Solution A : 6 mL Solution B)? How do you know?

Sodium hypochlorite (Solution A) was the limiting reactant. This can be determined using the ideal mole ratio between sodium hypochlorite and sodium thiosulfate (4:1). This means that four moles of sodium hypochlorite (Solution A) will react with a single mole of sodium thiosulfate (Solution B). Because the molarity of both solutions is the same, it is possible to compare the number of milliliters. Test tube *5* contains 6 mL (0.003 mol) of sodium thiosulfate (Solution B). Each milliliter of sodium thiosulfate (Solution B) will react with 4 mL of sodium hypochlorite (Solution A), so 24 mL (0.012 mol) of sodium hypochlorite (Solution A) is needed. The test tube, however, only contains 4 mL (0.002 mol) of sodium hypochlorite (Solution A). The sodium hypochlorite (Solution A) runs out first and, therefore, limits the reaction.

Synthesis Questions

Use available resources to help you answer the following questions.

1. The production of heat is one indicator of a chemical reaction. List at least two additional indicators of a chemical reaction and explain how each indicator could be measured experimentally.

Formation of a precipitate – The mass of the precipitate formed can be weighed after the precipitate is separated and dried.

Change in color – A colorimeter can be used to measure the concentration of the colored reactant or product.

Evolution of a gas – The volume or pressure of the gas can be measured.

2. Explain how stoichiometry of chemical reactions is similar to following a recipe while cooking.

Stoichiometry calculations allow chemists to determine the amounts of reactants they need to combine for a reaction to yield a particular amount of product(s) in the same way that a recipe tells you the amount of each ingredient needed to make a certain number of servings.

3. Explain how an understanding of stoichiometry can help chemical manufacturing companies maximize their profits.

By mixing reactants in their ideal ratio, the maximum amount of product(s) will be produced using the least possible amount of reactants. There will not be any wasted material and, therefore, the company will save money.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. How is the ideal ratio of reactants for a specific chemical reaction determined in the experiment?

- **A. By mixing various concentrations of the reactants and determining the combination that produced the greatest amount of products**
- **B.** By mixing various concentrations of the reactants and determining the combination that produced the smallest amount of products
- **C.** By mixing one chemical with a variety of different chemicals and determining which set of reactants produces a chemical change
- **D.** The ideal ratio of reactants cannot be determined in the experiment. You need to look up the balanced chemical reaction in the textbook before you begin the experiment

2. Stoichiometry is useful because it allows chemists to know:

- **A.** The temperature change that occurs when chemicals are mixed
- **B.** The extent to which a physical change has occurred
- **C.** The elements that different substances are made up of
- **D. The amount of reactants required to form a specific amount of products**

3. Given the following reaction, 3A + 2B → 2C, how many moles of B are required to react exactly with 10.0 moles of A?

- **A.** 3.33 mol B
- **B.** **6.67 mol B**
- **C.** 10.0 mol B
- **D.** 15.0 mol B

4. In exothermic reactions, heat is considered ________________________________.

- **A.** A reactant
- **B.** **A product**
- **C.** Neither a reactant nor a product
- **D.** Both a reactant and a product

5. What is a limiting reactant?

- **A.** **The reactant that is completely used up**
- **B.** The reactant of which there is an excess
- **C.** The product formed in the greatest amount
- **D.** The product formed in the least amount

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Chemical **stoichiometry** refers to the relationships that exist amongst reactants and products in chemical reactions. Stoichiometry allows chemists to calculate the amount of each **reactant** required to produce a product or determine the amount of product expected from a set of reactants. In symbol form, the stoichiometry of a reaction is indicated by the coefficients used to balance the chemical equation and to indicate the ideal **mole ratio** of reactants to products. A properly balanced equation satisfies the **law of conservation of matter** because exactly the same number and type of atoms on the reactant side are also represented on the product side.

2. If a reaction's stoichiometry (balanced chemical equation) is not known, chemists determine it experimentally by measuring an **observable property** that indicates a chemical reaction has taken place. The formation of a precipitate, formation of a gas, change in temperature, or a change in color indicate that a chemical reaction has taken place. The ideal ratio of reactants can be determined by reacting various concentrations of each reactant together. The combination that produces the greatest change in the property being observed is the **ideal ratio.**

3. When reactants are combined in a ratio that is not ideal, one of the reactants is completely consumed (stopping the reaction) and any other reactants still remain. The reactant that is completely used up is called the **limiting reactant**. Reactants that are left over are considered to be present in **excess**.

4. In this activity, the observable property being measured is a change in **temperature**. This is possible because the reaction is exothermic. An **exothermic reaction** is a reaction in which energy (in the form of heat) is produced. Energy can therefore be considered a **product.**

Extended Inquiry Suggestions

Try additional exothermic reactions: react sodium hypochlorite with sodium sulfite (Na_2SO_3) or potassium iodide (KI) to discover the correct stoichiometric mole ratio for those reactions using the same procedure described above.

Have students perform a similar analysis for a reaction where other properties can be measured: formation of a precipitate, evolution of a gas, or a color change. Guide students to understand that the same principle applies. The measured change in the amount of products formed corresponds to the overall "success" of the reaction.

19. Single Replacement Reactions

Objective

Determine the mass of copper consumed and silver deposited in a single replacement reaction. Through this investigation, students:

- Understand single replacement reactions
- Learn about the activity series of metals and how it is used to determine whether or not a chemical reaction will occur

Procedural Overview

Students conduct the following procedures:

- React solid copper with silver nitrate solution and measure the absorbance of the resulting copper(II) nitrate solution
- Determine the concentration of Cu^{2+} ions by analyzing a graph of absorbance versus concentration
- Calculate the amount of copper consumed and silver deposited in the reaction using stoichiometry

Time Requirement

♦ Preparation time	30 minutes
♦ Pre-lab discussion and activity	30 minutes (50 minutes if demonstrating reactions)
♦ Lab activity	50 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Colorimeter
- Sensor extension cable
- Glass cuvette with cap
- Balance, centigram
- Test tube, 20-mm x 150-mm
- Test tube rack
- Graduated cylinder, 100-mL
- Sand paper or steel wool
- Non-abrasive cleaning tissue
- 0.5 M Silver nitrate solution, 30.0 mL[1]
- Copper wire, 20.0 cm
- Paper towels

[1] To prepare using solid silver nitrate ($AgNO_3$), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Mole
- Molarity
- Evidence of chemical reactions
- Chemical equations
- Stoichiometry

Related Labs in This Guide

Labs conceptually related to this one include:

- Evidence of a Chemical Reaction
- Stoichiometry
- Electrochemical Battery: Energy from Electrons
- Molar Mass of Copper
- Concentration of a Solution: Beer’s Law

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Monitor live data in a digits display �(6.1)

Background

To better study and understand the incredible variety of possible chemical reactions, chemists classify reactions in a number of different ways. One common classification scheme recognizes four major types of chemical reactions:

- Synthesis reactions
- Decomposition reactions
- Single replacement reactions
- Double replacement reactions

Starting from this classification scheme, reactions can be further classified according to the type of chemistry that occurs: acid-base neutralization and oxidation-reduction reactions are two examples. In this activity, students investigate one example of a single replacement reaction. This type of reaction occurs when one element replaces another element in a compound. The general formula for a single replacement reaction is given below.

$$A + BC \rightarrow B + AC$$

You can determine whether one metal replaces a different metal that is already in a compound by the relative reactivity of the two metals. The activity series for metals ranks the elements in order of decreasing reactivity (the activity series of metals is found at the end of the Background section). Therefore, a metal high in the activity series replaces all metals beneath it if given the chance to react with one of them as part of a compound.

In all single replacement reactions, one element changes from its elemental metallic form into an ion while the other element changes from its ionic form into its elemental form. Each of these processes requires the transfer of electrons between the two metals, so all single replacement reactions are examples of oxidation-reduction, or redox, reactions.

In this experiment, students investigate the reaction of copper metal with a solution of silver nitrate. The silver ions receive electrons from the copper, changing them into solid silver metal atoms that deposit onto the surface of the copper wire as a dull, grey coating. The atoms comprising the solid copper wire lose their electrons and dissolve into the solution as copper(II) ions, resulting in a blue aqueous solution.

The greater the concentration of copper(II) ions, the more intense the blue color of the solution becomes. Because of this color change, students can use a colorimeter to measure the absorbance of the solution after the reaction has occurred. By using a calibration curve for copper(II) ions in solution, students are able to determine the concentration of the copper(II) ions. Students can then use stoichiometry to calculate both the amount of metallic copper that was consumed and the amount of solid silver deposited during the reaction.

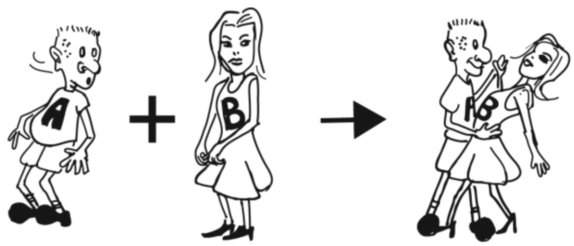
Table: Activity series of metals

	Metal Name	Symbol
Decreasing reactivity	Lithium	Li
	Potassium	K
	Calcium	Ca
	Sodium	Na
	Magnesium	Mg
	Aluminum	Al
	Zinc	Zn
	Iron	Fe
	Lead	Pb
	(Hydrogen)	(H)
	Copper	Cu
	Mercury	Hg
	Silver	Ag
	Platinum	Pt
	Gold	Au

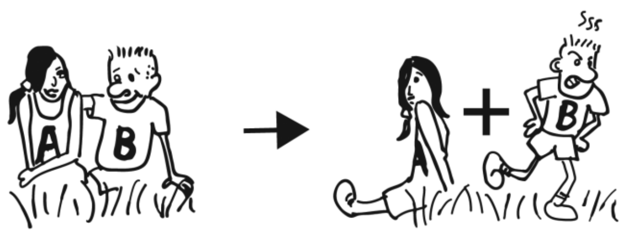
Pre-Lab Discussion and Activity

Classification of Chemical Reactions

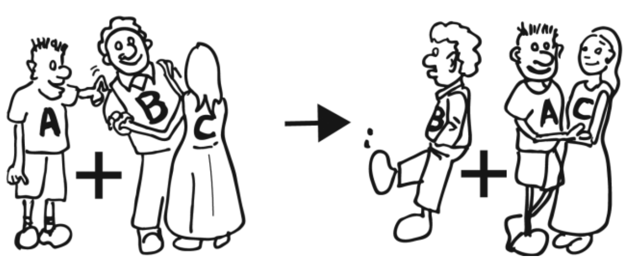
Review the basic types of chemical reactions: synthesis, decomposition, single replacement, and double replacement. Use the following cartoon examples below to help students understand reaction types in terms that they understand:



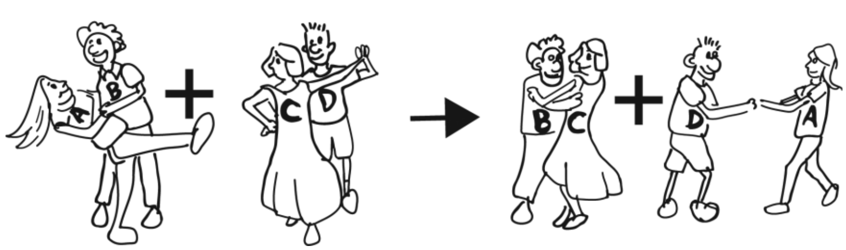
Synthesis



Decomposition



Single replacement



Double replacement

1. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Because two separate reactants are coming together to form only one product, this is an example of a synthesis reaction.

Teacher Tip: Demonstrate this synthesis reaction by using a gas burner to ignite a 3- to 5-cm strip of magnesium ribbon in a dark room. Hold the strip using clean tongs. The white magnesium oxide product coats the tongs and can easily be removed by wiping with a dry paper towel. Avoid looking directly at the brilliant white light produced by the reaction. Mention that energy being released is evidence of a chemical reaction.

2. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

Because one reactant is breaking down into multiple products, this is an example of a decomposition reaction.

Teacher Tip: Demonstrate this decomposition reaction by heating 5 g of sodium bicarbonate in a 125-mL Erlenmeyer flask. A popping sound might occur and the condensation of water vapor can be seen on the inside of the flask. Ignite a wood splint. Using tongs, lower the splint into the Erlenmeyer flask. The flame is extinguished by the carbon dioxide. Mention that the production of a gas (both carbon dioxide and water vapor) are evidence that a chemical reaction has occurred.

3. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

Because the aluminum metal replaces the iron ion producing elemental iron metal and aluminum ion, this is an example of a single replacement reaction.

Teacher Tip: Demonstrate this single replacement reaction using thermite obtained from a supplier (thermite is a mixture of iron(III) oxide and very fine aluminum powder). Because of the shower of sparks and large volume of smoke produced, this dramatic reaction must be performed in an open area outside! Molten iron is produced. Be sure to follow all safety instructions provided with the thermite. Mention that the production of energy is evidence that a chemical reaction has occurred.

4. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$Mg(s) + 2HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

Because the magnesium metal replaces the hydrogen ion producing elemental hydrogen gas and magnesium ion, this is also an example of a single replacement reaction.

Teacher Tip: Demonstrate this single replacement reaction by adding 1 to 2 g magnesium metal to 25 mL of 0.1 M HCl. The bubbles produced are hydrogen gas. Use an inverted small test tube to collect the hydrogen gas. Mention that the production of gas is evidence that a chemical reaction has occurred.

5. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

Because two reactants are coming together to form one product, this is another example of a synthesis reaction.

Teacher Tip: Demonstrate this synthesis reaction by igniting the hydrogen produced above. While the test tube remains inverted, bring a lit wooden splint near the open end of the test tube. The hydrogen "barks" and some condensation might be seen on the inside of the test tube. Mention that the production of energy is evidence that a chemical reaction occurred.

6. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$$

Because the calcium and sodium cations (+ ions) are exchanging anions (- ions), this is an example of a double replacement reaction. Specifically, this is a precipitation reaction creating a calcium carbonate precipitate.

Teacher's Tip: Demonstrate this double replacement reaction by combining 5 mL of calcium chloride solution with 5 mL of sodium carbonate solution (each made by pre-dissolving 1 to 2 g of ionic compound in separate vials). The cloudy white precipitate is the calcium carbonate product. Mention that the formation of a precipitate is evidence that a chemical reaction occurred.

7. Classify the following reaction as one of the four reaction types: synthesis, decomposition, single replacement, or double replacement.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + HOH(l)$$

Because the hydrogen and sodium cations (positive ions) are exchanging anions (negative ions), this is another example of a double replacement reaction. Specifically, this is a neutralization reaction between an acid (HCl) and a base (NaOH), producing a salt (NaCl) and water (HOH).

Teacher Tip: Demonstrate this by mixing 6 mL of 0.1 M NaOH into an Erlenmeyer flask containing 10 mL of water, 1 to 2 drops of 1% phenolphthalein indicator solution, and 5 mL of 0.1 M HCl. The solution turns from colorless to magenta, indicating that the acid has been consumed by the base. Mention that the color change is evidence that a chemical reaction has occurred.

Copper Reaction with Silver Nitrate

Write the following reaction on the board and tell the students they will be performing this reaction in the experiment:

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

Use the activity series table to show that copper is "more active" than silver and is able to displace it from the nitrate ion. Evidence of this reaction is the formation of solid silver and the blue color change of the solution produced by the copper(II) ions. Explain that a piece of copper wire is submerged in a solution of silver nitrate. With time, the solid copper oxidizes and the resulting copper(II) ions dissolve into the solution. At the same time, the silver ions are reduced to the solid form of silver and are deposited onto the copper wire.

8. How might the amount of copper consumed or silver deposited be determined?

Possible ideas might include taking the mass of the copper wire before and after the experiment or filtering the solution and finding the mass of the solid silver after allowing it to dry. Students with experience in colorimetry might suggest using Beer's law to determine the concentration of copper(II) ions by measuring the absorbance of the blue solution.

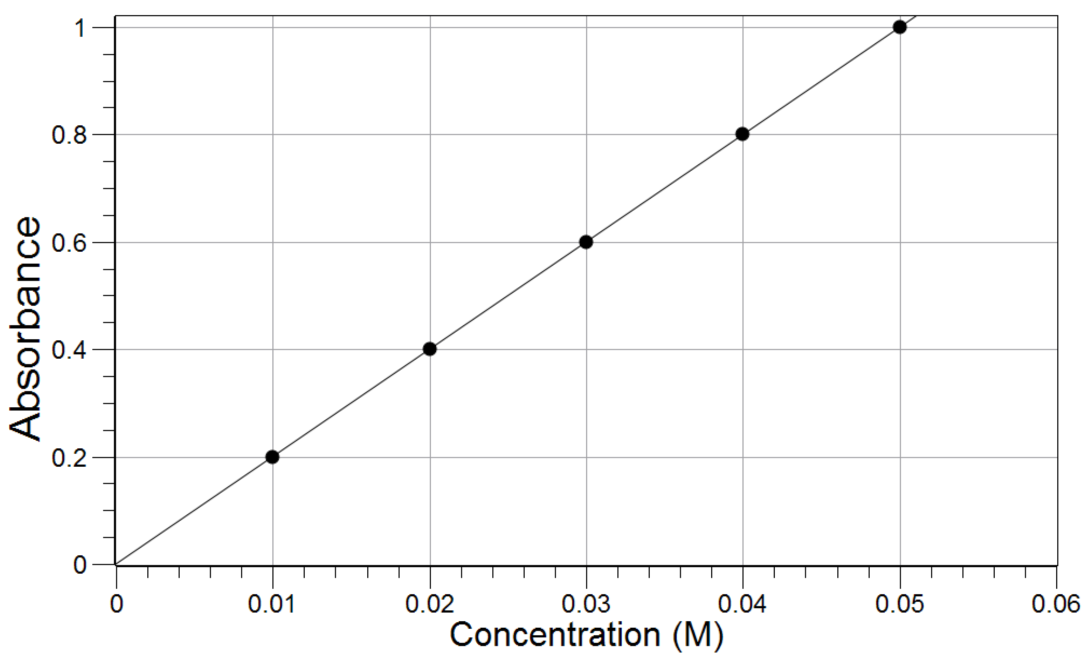
Tell students they will use the method of colorimetry to determine the concentration of blue colored copper ions and then use stoichiometry to find the amount of copper consumed and silver deposited. Review the concept of colorimetry by showing two solutions of food coloring. One solution should be very dark in color and almost opaque. The other should be very light in color.

9. Which solution is more concentrated?

The more concentrated, opaque solution absorbs more light than the less concentrated solution. In fact, the absorbance should be directly proportional to the concentration.

Display the following sample graph and demonstrate the method to predict the concentration of a solution that has an absorbance of 0.90.

Calibration Curve



Guide students in finding the slope of the line, and writing the equation of the line in the form *y* = *mx* + *b*. In this case, the equation of the line is:

$$\text{slope} = \frac{\Delta y}{\Delta x} = \frac{1 - 0.2}{0.05 - 0.01} = \frac{0.8}{0.04} = 20$$

y-intercept = 0

absorbance = 20(concentration) + 0

Therefore, the concentration of a solution that has an absorbance of 0.9 is 0.045 M.

10. Determine the concentration of a solution that has an absorbance of 0.50.

$$\text{absorbance} = 20(\text{concentration}) + 0$$

$$0.50 = 20(\text{concentration})$$

$$(\text{concentration}) = \frac{0.50}{20} = 0.025\ \text{M}$$

Lab Preparation

These are the materials and equipment to set up prior to the lab.

The following instructions make 500 mL of 0.5 M silver nitrate solution. This is enough for *16* lab groups.

1. Prepare 500 mL of 0.5 M silver nitrate solution by following the steps below.

   a. Add approximately 200 mL of distilled water to a 500-mL volumetric flask.

   b. Add 42.47 g $AgNO_3$ to the water and swirl to dissolve.

   c. Dilute to the mark with distilled water.

   d. Cap and invert three times to ensure complete mixing.

Teacher Tips:

- Silver nitrate is a fairly expensive compound. The experiment will work with 0.2 M silver nitrate (dissolve 16.99 g $AgNO_3$) solution as well, but the amount of copper used up is smaller (0.19 g) and, therefore, more difficult to measure. This also makes the absorbance lower, but the experiment will work.

- The waste generated in the procedure must be collected and disposed of according to your local, state, and federal regulations. You can collect the silver residue and the copper(II)/silver nitrate solution in a single container and process it following the procedures outlined by your organization.

- In the data analysis section, a general calibration curve is provided for the students to use to analyze their data. If time permits, a more rigorous treatment of the data can be applied by asking the students to create their own curves specific for their colorimeter from a serial dilution of a standard copper(II) nitrate solution as described in the Determining the Concentration of a Solution: Beer's Law experiment.

Safety

Add these important safety precautions to your normal laboratory procedures:

- The silver nitrate ($AgNO_3$) solution may temporarily stain your skin when exposed to bright light. If the solution contacts your skin, wash it with soap and water immediately.

- Both the silver nitrate and copper(II) ions are hazardous to the environment and should not be disposed of down the drain. Make sure you follow the teacher's instructions about how to properly dispose of the chemicals.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	5	4	3
Suspend the coiled copper wire in the test tube and then fill the test tube with silver nitrate solution.	Clean a piece of copper wire with steel wool or sandpaper and then find the mass of the copper.	Use a colorimeter to determine the absorbance of the solution that remained in the test tube after the copper was removed.	Remove the copper wire from the test tube, wipe off the silver deposits, and measure the mass of the copper wire.	Allow the reaction between the copper wire and silver nitrate to proceed for 15 to 20 minutes.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

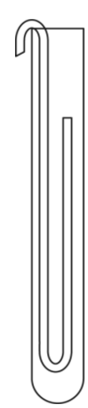
1. ☐ Clean a 20-cm piece of heavy copper wire using steel wool or sand paper.

2. ☐ Why must the copper wire be cleaned before the experiment begins?

You clean the copper wire to ensure that it is free of oxidation allowing the maximum amount of surface area of copper to be exposed to the silver nitrate solution.

3. ☐ Use a centigram balance to determine the mass of the copper to the nearest 0.01 g, and record the mass below.

Mass of copper before the reaction (g): 2.64 g



4. ☐ Form the lower part of the wire into an elongated u-shape, and bend the wire to form a hook at the opposite end.

5. ☐ Suspend the u-shaped copper in the test tube by hanging the hooked end of the copper over the side of the test tube. Do not let the copper wire touch the bottom of the test tube. Place the test tube in a test tube rack.

6. ☐ Measure 30.0 mL of 0.5 M silver nitrate solution and pour it into the test tube containing the copper wire.

7. ☐ Use three adjectives to describe the silver nitrate solution.

Clear, colorless, liquid

8. ☐ Allow the reaction to proceed undisturbed for 15 to 20 minutes.

9. ☐ Why must you wait for at least 15 to 20 minutes?

Chemical reactions do not occur instantaneously and sometimes require time for the molecules to react with each other. In this case, 15 to 20 minutes will ensure that enough copper metal has been oxidized to have a measurable result.

10. ☐ Remove the copper wire from the silver nitrate solution. Save the solution.

11. ☐ Use a paper towel to wipe the silver deposits off the copper wire. Dispose of the silver in the proper receptacle according to your teacher's instructions.

12. ☐ Use the centigram balance to find the mass of the copper wire after the chemical reaction. Record the mass to 0.01 g below.

Mass of copper after the reaction (g): 2.33 g

13. ☐ Why is this procedure likely to give an inaccurate mass?

Wiping the copper with a paper towel does not guarantee that you have removed all the water or silver deposits. Heating the copper wire would remove the water, however it would promote the formation of copper(II) oxide, CuO. This would lead to an additional inaccurate measurement for the amount of copper consumed during the reaction. Scraping the wire or cleaning with steel wool or sandpaper would remove additional copper from the wire causing further inaccuracies in the amount of copper consumed by the reaction.

14. ☐ Start a new experiment on the data collection system. �(1.2)

15. ☐ Connect the colorimeter to the data collection system using a sensor extension cable. �(2.1)

16. ☐ Monitor the Absorbance of Orange (610 nm) Light in a digits display. �(6.1)

17. ☐ Calibrate the colorimeter by following the steps below:

a. Fill a cuvette with distilled water and cap it.

b. Holding the cuvette by the lid, wipe the outside of the cuvette with a non-abrasive cleaning tissue.

c. Place the cuvette in the colorimeter and close the lid.

d. Push the green calibrate button on the colorimeter.

e. When the light turns off, the calibration is complete and the cuvette can be removed from the colorimeter.

18. ☐ Rinse the cuvette with some of the solution saved in the test tube by following the steps below:

a. Fill the cuvette with some of the solution that is saved in the test tube.

b. Dispose of the solution you just poured into the cuvette according to the teacher's instructions.

19. ☐ Fill the cuvette with more solution from the test tube and cap the cuvette.

20. ☐ Why was the cuvette rinsed (filled and then emptied) with the solution from the test tube?

Rinsing the cuvette one time removes excess water that could dilute the solution being measure. This could lead to inaccurate results.

21. ☐ Describe any changes to the solution compared to the original solution that was in the test tube. What may have caused these changes?

The solution is now light blue in color, while the solution was originally clear and colorless. The color change is due to the presence of copper(II) ions in the solution.

22. ☐ Use a non-abrasive cleaning tissue to wipe the outside of the cuvette containing the solution from the test tube, and then place the cuvette inside the colorimeter. Close the lid of the colorimeter.

23. ☐ Why is it necessary to wipe the outside of the cuvette before placing it in the colorimeter?

You need to wipe the cuvette because any substances on the outside of the cuvette, including finger prints, will absorb (or scatter) light and cause the absorbance values to be lower than the actual values.

24. ☐ Why is it necessary to close the lid of the colorimeter before recording the data values?

The colorimeter measures the absorbance and transmittance of light. If the lid is open, the light from the room interferes with the reading, causing the absorbance to be lower than the actual value.

25. ☐ Record the value for the absorbance below.

Absorbance of Orange (610 nm) Light: 0.365

26. ☐ Remove the cuvette and properly dispose of the solution according to the teacher's instructions.

27. ☐ Clean up the lab station according to the teacher's instructions, including any special instructions for disposing of the silver deposit and copper(II)/silver nitrate solution.

Data Analysis

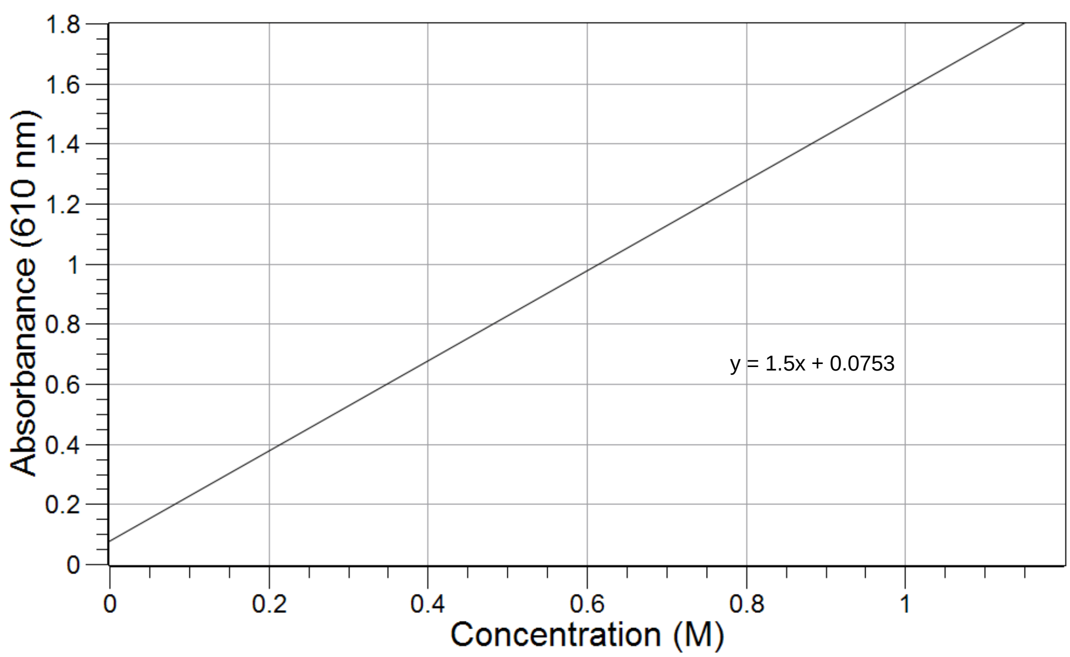
1. ☐ Calculate the mass of copper consumed in the reaction using the mass of the copper wire measured before the reaction after the reaction.

mass copper consumed = mass copper before reaction – mass copper after reaction

0.31 g = 2.64 g – 2.33 g

2. ☐ Using the calibration curve for copper(II) ions in solution provided below and your absorbance reading for the solution in the test tube, determine the concentration of copper(II) ions in the solution.

Calibration Curve for Copper(II) Solution



y = 1.50x + 0.0753

absorbance = 1.50(concentration) + 0.0753

0.365 = 1.50(concentration) + 0.0753

concentration of copper(II) ions = 0.193 M

3. ☐ Based on the concentration of copper ions in solution determined above, calculate the number of moles of copper(II) ions (as copper(II) nitrate) in the final solution. Remember, the volume of the solution was 30.0 mL.

$$\left(\frac{0.193\text{ mol}}{\text{L}}\right)(0.0300\text{ L}) = 5.97 \times 10^{-3}\text{ mol Cu}^{2+} = 5.79 \times 10^{-3}\text{ mol Cu(NO}_3)_2$$

4. ☐ Based on the previous calculation and the stoichiometry of the reaction given below, calculate the mass of metallic copper that reacted with the silver nitrate solution.

$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

The mole ratio of metallic copper to copper(II) nitrate is 1:1, so the amount of metallic copper that must have reacted must also be 5.79×10^{-3} mol.

$$(5.79 \times 10^{-3}\text{ mol})\left(\frac{63.546\text{ g}}{\text{mol}}\right) = 0.368\text{ g copper reacted}$$

Analysis Questions

1. You have found the mass of copper consumed using two different methods. What are the two methods and how do the results compare?

The first method was to use a centigram balance to weight the copper before and after the reaction occurred. The second method was to measure the amount of copper(II) ions that were present in the solution after the reaction. The mass of copper was then calculated from the concentration measured.

Using the first method 0.31 g of copper were transferred to copper(II) ions. In the second method 0.368 g of copper were transferred to copper(II) ions. The two methods varied by 0.1 grams. This is a significant difference considering that the total amount of copper was less than half a gram.

2. List some possible sources of error with each method.

Possible sources of error for the first method include not cleaning enough silver off the metal wire, not drying the metal wire enough, the formation of copper(II) oxide on the wire, and so on.

Possible sources of error for the second method include uneven mixing of the copper(II) ion solution, error in reading the absorbance, error in the calibration curve, and so on.

3. Predict the mass of silver produced in the reaction.

The mole ratio of copper(II) nitrate to metallic silver is 1:2, so the amount of metallic silver that must have been produced is 1.16×10^{-2} mol.

$$5.79 \times 10^{-3}\text{ mol Cu(NO}_3)_2 \times \frac{2\text{ mol Ag}}{1\text{ mol Cu(NO}_3)_2} = 1.16 \times 10^{-2}\text{ mol Ag}$$

$$(1.16 \times 10^{-2}\text{ mol})\left(\frac{107.87\text{ g}}{\text{mol}}\right) = 1.25\text{ g Ag produced}$$

Synthesis Questions

Use available resources to help you answer the following questions.

1. During the Middle Ages, alchemists searched unsuccessfully for a way to "transmute" or change one element into another, most notably lead into gold. Was copper transmuted into silver in this experiment? Why or why not?

No. The copper was replaced by silver that already existed in the silver nitrate solution. When metallic ions are dissolved in solution, they are not visible.

2. How can you account for the color change of the solution?

After the reaction, the solution contained copper(II) ions (Cu^{2+}). This caused the blue color.

3. Using the activity series table in the Background section, predict the outcome of this experiment if you used a gold wire instead of a copper wire. Explain your prediction.

There would be no reaction. Silver is ranked higher than gold on the activity series and would, therefore, stay in solution. In this experiment, we used copper wire. Copper is ranked higher than silver on the activity series and was, therefore, able to replace silver from the solution.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

Note: Students will need a copy of the activity series of metals to answer some of the following questions.

1. What can be predicted using an activity series?

- **A.** The amount of energy released by a chemical reaction
- **B.** The electronegativity values of elements
- **C.** The melting points of elements
- **D. Whether or not a reaction will occur**

2. If one metal is lower than a second in the activity series, then it:

- **A.** Replaces the ions of the second metal in solution
- **B.** Loses electrons more readily than the second metal
- **C. Loses electrons less readily than the second metal**
- **D.** Forms positive ions more readily than the second metal

3. What happens when lead is added to nitric acid?

- **A.** No reaction occurs
- **B. Hydrogen is released**
- **C.** Lead oxide forms
- **D.** Oxygen is released

4. What happens when iron is added to a solution of lead(II) sulfate?

- **A.** No reaction occurs
- **B.** Hydrogen is released
- **C.** Lead oxide forms
- **D.** **Lead will deposit on the surface of the iron**

5. What happens when zinc is added to a solution of magnesium chloride?

- **A.** **No reaction occurs**
- **B.** Hydrogen is released
- **C.** Zinc chloride forms
- **D.** Magnesium will deposit on the surface of the zinc

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. A **single replacement** reaction occurs when one element replaces another element in a compound. You can predict whether or not one metal will replace another metal by use of an **activity series**. Metals that are **higher** in the activity series replace metals that are **beneath** them. In these types of reactions, one element changes into its ionic form and the other element changes into its elemental form. This process involves a transfer of electrons from one substance to another and is referred to as a **redox reaction**. If an element loses electrons, it is being **oxidized**. If an element gains electrons, it is being **reduced**.

Extended Inquiry Suggestions

Single replacement reactions are oxidation-reduction (redox) reactions. Repeat this lab, or perform a demonstration with another redox reaction such as:

$$Mg(s) + CuSO_4(aq) \rightarrow Cu(s) + MgSO_4(aq)$$

$$Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$$

In these reactions the more active metal will replace the copper ions in solution at a faster rate and the solution loses its color. The solution with the magnesium will lose its color more rapidly than the solution containing the iron.

20. Molar Mass of Copper

Objectives

Determine the molar mass of copper through electroplating in an electrolytic cell. Through this investigation, students:

- Differentiate between oxidation and reduction
- Use the area under a curve to provide meaningful data
- Convert current (amperes) to moles of electrons using the Faraday constant
- Use stoichiometric calculations

Procedural Overview

Students conduct the following procedures:

- Electroplate an object with copper
- Determine the mass of deposited copper by finding the difference between the mass of the object before and after it is plated
- Determine the number of electrons transferred during electroplating from the area under the curve of a Current versus Time graph
- Use stoichiometry and the number of electrons transferred to determine the number of moles of copper reduced
- Calculate the molar mass of copper by dividing the mass of copper plated by the moles of copper determined experimentally

Time Requirement

♦ Preparation time	30 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	50 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Voltage-current sensor
- Balance, centigram
- Beaker, 250-mL
- Utility clamp (2), insulated[1]
- DC power supply
- Red patch cord (2), 4-mm banana plug
- Black patch cord, 4-mm banana plug
- Alligator clip (2)
- Copper electrode[2]

- Ring stand
- Magnetic stirrer
- Magnetic stir bar
- Stainless steel spoon (or other item to electroplate)
- 0.5 M Copper(II) sulfate ($CuSO_4$), 150 mL[3]

[1]If the utility clamps are not insulated, use electrical tape and wrap it around the spoon and copper electrode. This ensures that the electrons only go through the electroplating circuit.

[2]Copper can be found as electrical wire, water pipe, and sheeting available at hardware stores.

[3]To formulate using solid copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Ions and electrolyte solutions
- Molar mass
- Stoichiometry

Related Labs in This Guide

Labs conceptually related to this one include:

- Electrochemical Battery: Energy from Electrons
- Single Replacement Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Finding the area under a curve �(9.7)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

Elements have certain qualities that classify them as either metals or nonmetals. Metals have luster, are malleable and ductile, and conduct both heat and electricity. They tend to lose electrons to become positively charged ions (cations). The ability to lose electrons, however, is relative to competition with other metals.

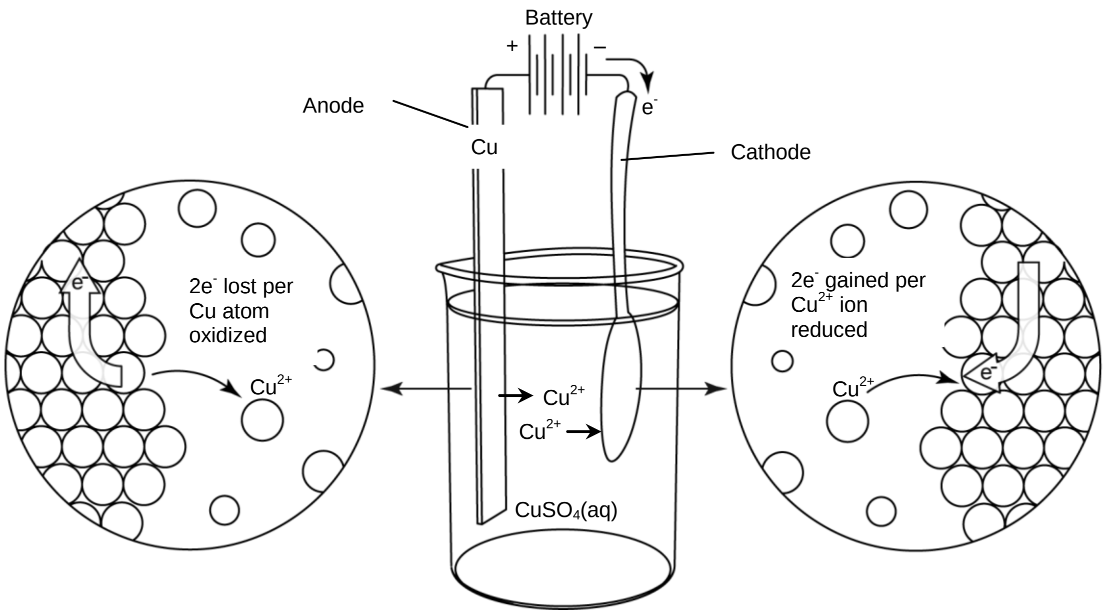
In single replacement reactions, metals are arranged in a hierarchy according to how active they are. More active metals are able to displace those in compounds with a lower activity by giving electrons to them. In general, any reaction that involves the transfer of electrons between reactants is a reduction-oxidation, or redox, reaction.

Specifically, the chemical species that loses electrons in the process are said to be oxidized. Those that gain electrons are said to be reduced. The mnemonic, "oil rig" = *o*xidation *i*s a *l*oss of electrons, *r*eduction *i*s a *g*ain of electrons is one way to remember the direction of electron transfer.

The potential difference between the two chemical reactants involved in a redox reaction can be measured as voltage (V). A table of standard reduction potentials can be used to predict the voltage between two materials used in a electrochemical battery (voltaic cell). In an electrochemical battery, a chemical reaction occurs spontaneously and produces energy in the form of electricity. In such a battery, the two reactants are physically separated from one another in two half-cells. This forces the transfer of electrons to occur through a wire and requires an inert "salt bridge" to balance the charges in each cell. The electrochemical cell then uses these flowing electrons (electrical current) to do work. The greater the flow of electrons is, the higher the amperage (A). (A table of standard reduction potentials is provided at the end of the Background section.)

An electrolytic cell is the opposite of an electrochemical battery. An electrolytic cell is not spontaneous, but instead uses electric current to make a chemical reaction occur. The process of using an electrolytic cell is called electrolysis. Electrolysis is used to electroplate and purify substances.

In an electrolytic cell, the reactants are not physically separated from one another. Instead, flowing electrons are transferred to a solution containing the ions of the metal that will be plated onto the object (a spoon in this investigation). This causes these ions to be reduced (gain of electrons) to their solid form and deposited on the object to be plated. The object being plated is the cathode (the electrode that is the site of reduction). The concentration of the ions in solution should not change over time. This is because the ions in solution are replaced by the oxidation of the anode at the same rate they are removed from solution at the cathode.



Electrolytic cell

A current of one ampere running for one second delivers one Coulomb (C) of electrical charge. The Faraday constant (96485 A•s/mol) is the amount of electrical charge associated with one mole of electrons. Thus, the number of electrons can be determined from knowing how long a particular current runs. For example, a 2.0 A current running for 100 seconds delivers 1.2 x 10^{21} electrons.

$$2.0\ \text{A} \times 100\ \text{s} \times \frac{1\ \text{mol}}{96{,}485\ \text{A}\cdot\text{s}} = 2.1\times10^{-3}\ \text{mol}\ e^{-}$$

$$2.1\times\ 10^{-3}\ \text{mol}\ e^{-}\times \frac{6.022\times10^{23}\ e^{-}}{1\ \text{mol}\ e^{-}} = 1.2\times10^{21}\ e^{-}$$

The stoichiometry of the half-reaction can then be used to determine the amount in moles of material electroplated. A mole is a measurement that describes a set number of particles. Just as a dozen describes twelve units of anything, a mole describes 6.022×10^{23} units. The units can be atoms, molecules, or anything else. The mass (in grams) of one mole (6.022×10^{23} units) of an element or a compound is the molar mass of that substance.

By measuring the mass of the electroplated material, the moles can then be used to calculate a value for molar mass. Using the example above, if the experimental difference in mass before and after electroplating is found to be 0.070 g after applying 2.0 A of current for 100 seconds, then the experimentally determined molar mass of copper is 64 g/mol.

$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu^{0}(s)$ (reduction half-reaction)

$$2.1\times10^{-3}\ \text{mol}\ e^{-}\times\frac{1\ \text{mol Cu}^{0}(\text{s})}{2\ \text{mol}\ e^{-}} = 1.1\times10^{-3}\ \text{mol Cu}^{0}(\text{s})$$

$$\frac{0.070\ \text{g}}{1.1\times10^{-3}\ \text{mol Cu}^{0}(\text{s})} = 64\ \text{g/mol}$$

Table: Standard reduction potentials in aqueous solution at 25 °C

Reduction Half-Reaction			E° (V)
$F_2(g) + 2e^-$	⇌	$2F^-(aq)$	+2.87
$Au^{3+}(aq) + 3e^-$	⇌	$Au(s)$	+1.50
$Ag^+(aq) + e^-$	⇌	$Ag(s)$	+0.80
$Cu^{2+}(aq) + 2e^-$	⇌	$Cu(s)$	+0.34
$2H_3O^+(aq) + 2e^-$	⇌	$H_2(g) + 2H_2O(l)$	0.00
$Sn^{2+}(aq) + 2e^-$	⇌	Sn s)	–0.14
$Ni^{2+}(aq) + 2e^-$	⇌	$Ni(s)$	–0.25
$Fe^{2+}(aq) + 2e^-$	⇌	$Fe(s)$	–0.44
$Zn^{2+}(aq) + 2e^-$	⇌	$Zn(s)$	–0.76
$Al^{3+}(aq) + 3e^-$	⇌	$Al(s)$	–1.66
$Mg^{2+}(aq) + 2e^-$	⇌	$Mg(s)$	–2.37
$Na^+(aq) + e^-$	⇌	$Na(s)$	–2.71
$K^+(aq) + e^-$	⇌	$K(s)$	–2.93

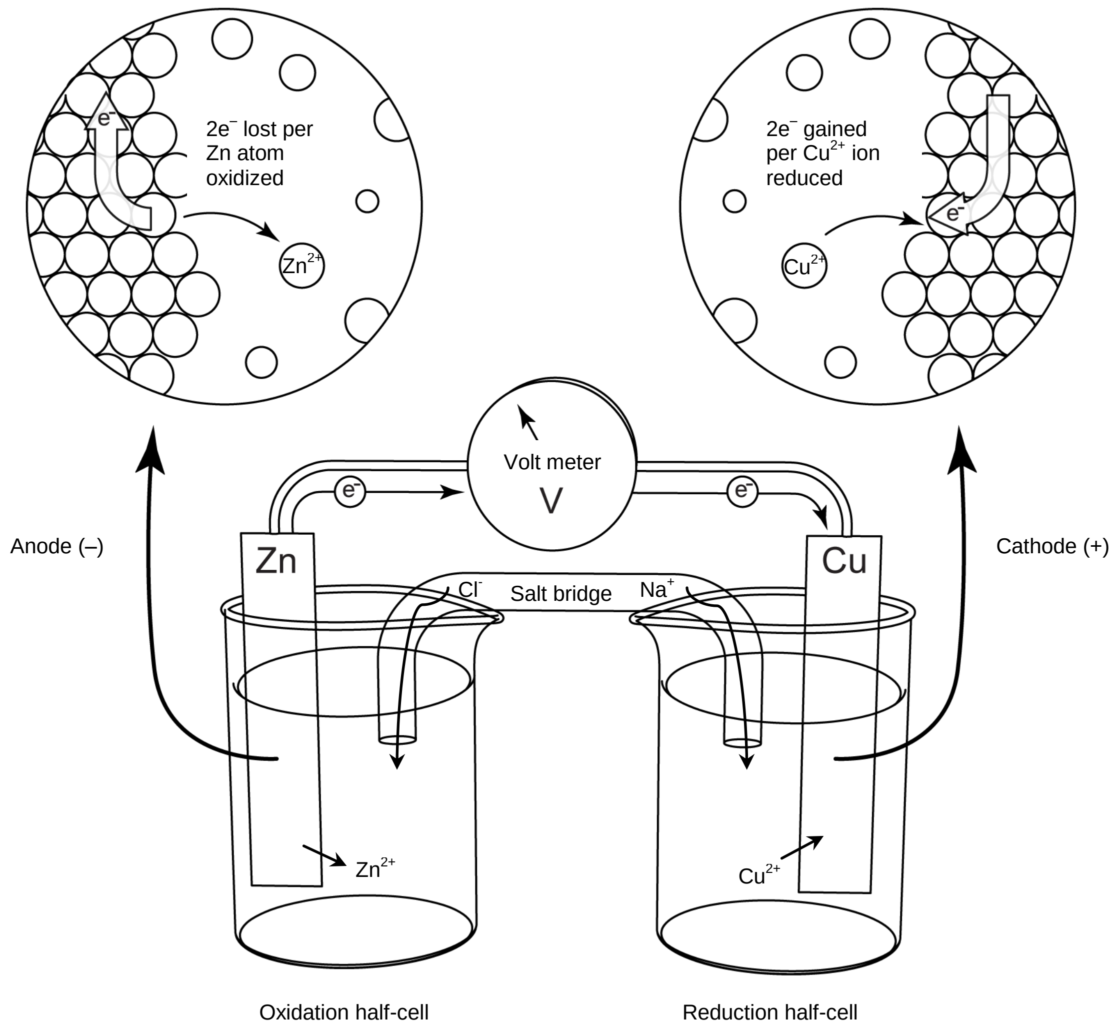
Pre-Lab Discussion and Activity

Redox Reactions

Reactions involving the transfer of electrons between atoms are referred to as reduction-oxidation, or redox, reactions. Because of the law of conservation of matter, one reactant must lose electrons (oxidation) in order for the other reactant to gain them (reduction).

The separate processes are referred to as half-reactions because they each represent half of the overall reaction. The reactants can be physically separated into half-cells and connected by a wire through which electrons travel. Electricity is a flow of electrons. If the reactants spontaneously react, the flow of electrons through the wire can be measured as electricity.

Construct the simple voltaic cell pictured below. Explain each component:



Oxidation half-cell: **Electrons are lost by the zinc metal electrode (anode). They flow toward the reduction half-cell through the wire. The zinc atoms that lose electrons become ionized and dissolve into the solution as Zn^{2+}.**

Reduction half-cell: **Electrons are gained by the copper(II) ions in solution. They become neutral copper atoms which deposit on the copper metal electrode (cathode).**

Teacher's Tip: A mnemonic for anode and cathode is "an ox" (anode is oxidation) and "red cat" (reduction at the cathode).

Wire: **Electrons flow through a conduit from the oxidation half-cell to the reduction half-cell. By placing a meter or a light bulb between the cells, the flowing electrons can do work and be seen as electricity.**

Salt-bridge: **The flow of electrons changes the balance of charge, both positive and negative. Ions must be able to flow as well in order to neutralize the build up of either positive charge (in the oxidation half-cell) or negative charge (in the reduction half-cell).**

1. How do you know electrons are flowing through the wire between the cells?

The light bulb is illuminated or the meter indicates a flow of electrons. Electricity is a flow of electrons.

2. What is the purpose of the salt bridge?

As electrons are transferred, positive ions build up in the oxidation half-cell solution and are removed from the reduction half-cell solution. The salt bridge allows positive and negative ions to move between the cells to neutralize the excess charge.

3. What happens if the cells are not separated?

The system would short circuit and electrons would move directly between the electrodes through the solution and not pass through the wire.

4. What would you see if the solution surrounding the anode is different, such as silver sulfate?

The silver ions would gain the electrons and become silver metal atoms. These silver atoms would coat the copper electrode on its outside surface.

Voltage and reduction potentials

Write the overall reaction between zinc and copper on the board: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu^{0}(s)$.

The reaction between copper ions and zinc is spontaneous. The standard reduction potentials (listed in the Background section) assists in the prediction of the voltage between two reactants. Display the table of standard reduction potentials.

Reversing the direction of a reaction changes the sign on the value for the potential. Because of this, one table can be used for oxidation reactions as well. The reaction is spontaneous if the sum of the half-reactions and their potentials gives a positive value for the overall redox reaction.

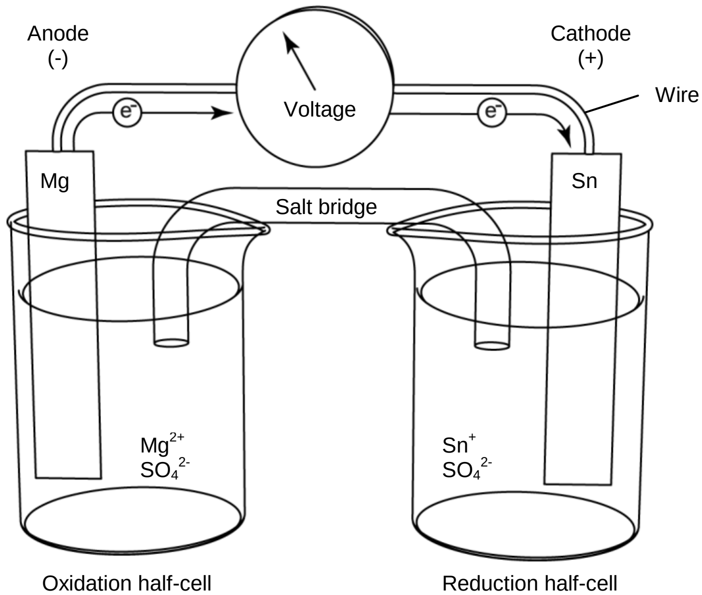
Separate the overall reaction into the two half-reactions and give their potentials. Sum the half-reactions to produce the overall reaction with its overall potential, emphasizing the positive value (spontaneous).

	Oxidation half-reaction:	**$Zn(s)$**	→	**$Zn^{2+}(aq) + 2e^{-}$**	**$E^{o} = +0.76$ V**
+	**Reduction half-reaction:**	**$Cu^{2+}(aq) + 2\,e^{-}$**	→	**$Cu^{0}(s)$**	**$E^{o} = +0.34$ V**
=	**Overall redox reaction:**	**$Zn(s) + Cu^{2+}(aq)$**	→	**$Zn^{2+}(aq) + Cu^{0}(s)$**	**$E^{o} = +1.10$ V**

5. Give the overall redox reaction and calculate the voltage between magnesium and tin, showing that the reaction is spontaneous.

	Oxidation half-reaction:	$Mg(s)$	→	$Mg^{2+}(aq) + 2\,e^{-}$	$E^{o} = +2.37$ V
+	Reduction half-reaction:	$Sn^{2+}(aq) + 2\,e^{-}$	→	$Sn^{0}(s)$	$E^{o} = -0.14$ V
=	Overall redox reaction:	$Mg(s) + Sn^{2+}(aq)$	→	$Mg^{2+}(aq) + Sn^{0}(s)$	$E^{o} = +2.23$ V

6. Sketch a picture of the apparatus allowing magnesium and tin to react spontaneously. Label the wire, salt bridge, reduction half-cell, and oxidation half-cell. Also label which metal is the anode and which metal is the cathode. Use arrows to show the direction of electron flow.



7. Which electrode is eventually consumed?

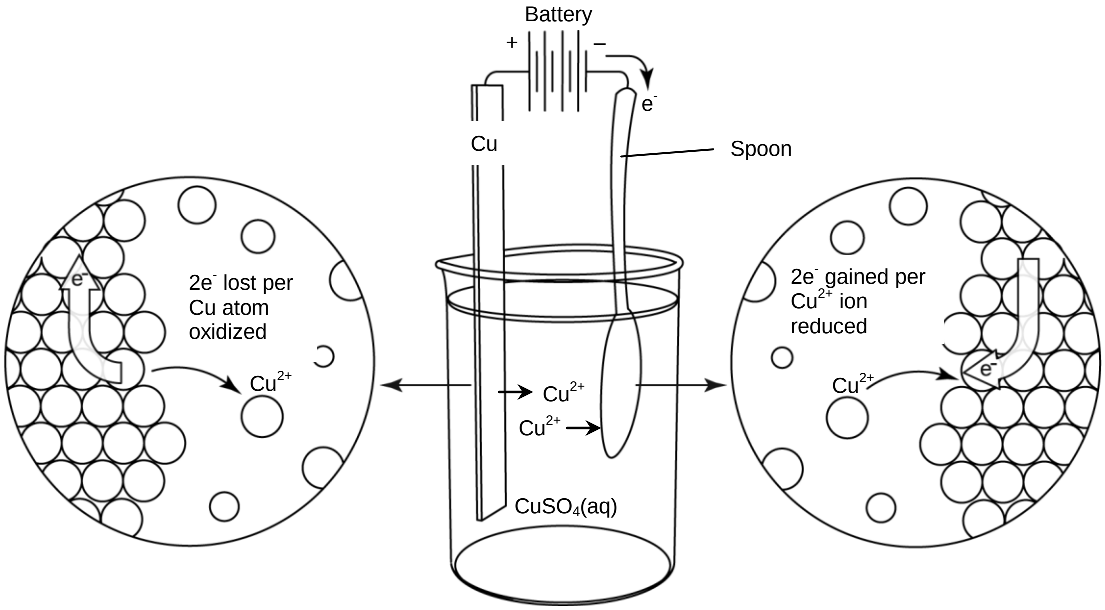
The magnesium electrode is eventually consumed because the magnesium metal is dissolving as the atoms become magnesium ions in solution.

Electrolytic cells

Batteries can be constructed using materials that produce a positive voltage when combined. These are known as voltaic (or galvanic) cells. However, reactions that are not spontaneous (negative voltages when combined) require an input of energy.

In electrolytic cells, the reactions are not spontaneous and require a source of electricity. The half-reactions are not separated into half-cells. Some applications of electrolysis include the following: purify metals such as aluminum, produce sodium metal, chlorine gas, as well as sodium hydroxide, and to electroplate objects.

In electroplating a spoon, the amount of product is a direct result of the number of electrons transferred. The copper that is deposited on the spoon is replaced by copper dissolving into solution from the copper strip. The overall amount of copper(II) ion in solution remains constant.



The reduction reaction at the spoon (cathode): $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu^{0}(s)$**.**

The oxidation reaction at the copper strip (anode): $Cu^{0}(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$ **.**

8. When does the reaction stop?

The reaction stops when the battery is disconnected, the battery dies, or the copper strip is completely consumed.

9. What happens to the concentration of copper(II) ions in solution as electrons are transferred?

Because the copper(II) ions are being replaced from the copper strip at the same rate as they are deposited onto the surface of the spoon, the concentration does not change.

10. How many moles of electrons are transferred for every mole of copper deposited?

Two moles of electrons are required for every one mole of copper that is electroplated onto the spoon.

11. How many moles of copper are deposited if 6 moles of electrons are transferred?

$$6 \text{ mol } e^{-} \times \frac{1 \text{ mol } Cu^{0}(s)}{2 \text{ mol } e^{-}} = 3 \text{ mol } Cu^{0}(s)$$

Current and the Faraday constant

Just as the current of a river describes how fast water is flowing, the current of electricity describes how fast electrons are flowing through a wire. There are many units describing various properties of electricity, but this experiment uses only one: the ampere (A). By multiplying the current in amperes by the length of time the current runs, the number of coulombs can be determined. A coulomb (C) is the amount of electric charge that flows in one second (1 C = 1 A • 1 s).

By multiplying the number of coulombs by the Faraday constant (96485 C/mol), the number of electrons (as moles) can be determined. A better approach is to find the area under the curve from a graph of amperes versus time, which is part of this experiment. From this, the number of moles of copper deposited can be determined using the stoichiometry of the reduction half-reaction.

By knowing the number of moles and the mass of metal deposited, the unit of molar mass for the metal can be calculated. The mass of the metal can be determined by measuring the mass of the object before and after it is electroplated.

12. How many moles of electrons would be transferred by a 5.0 A current running for 5.0 minutes?

First convert minutes to units of seconds. Also, by using ampere-seconds instead of coulombs in the Faraday constant (1 C = 1 A•s) the equation is easier to solve using dimensional analysis.

$$5.0\text{ min} \times \frac{60\text{ s}}{1\text{ min}} = 300\text{ s}$$

$$5.0\text{ A} \times 300\text{ s} \times \frac{1\text{ mol } e^-}{96{,}485\text{ A} \cdot \text{s}} = 0.016\text{ mol } e^-$$

Teacher Tip: Instead of multiplying the set amperage and the time the current ran, the experimental procedure calls for finding the area under the ampere-second curve. Using the area under the curve is a much more accurate method for determining the product of amperes and seconds.

13. The metal parts on many vehicles are electroplated with a thin layer of chrome to give them a mirror-like finish: $Cr^{6+}(aq) + 6e^- \rightarrow Cr^0(s)$. How many moles of chrome can be deposited using a 5.0 A current running for 5.0 minutes?

The number of electrons calculated in the previous question is substituted into the calculation below.

$$0.016\text{ mol } e^- \times \frac{1\text{ mol } Cr^0(s)}{6\text{ mol } e^-} = 2.7 \times 10^{-3}\text{ mol } Cr^0(s)$$

14. Assume a small metal knob with a mass of 15.27 g before electroplating with chromium and 15.41 g after 5.0 A of current runs for 5.0 minutes. What is the molar mass for chromium?

The mass of chromium plate is found by the difference in the mass of the knob before and after the process.

15.41 g – 15.27 g = 0.14 g

The number of moles of chromium calculated in the previous question is substituted into the calculation below.

$$\frac{0.14\text{ g } Cr^0}{2.7 \times 10^{-3}\text{ mol } Cr^0} = 52\text{ g/mol}$$

Lab Preparation

These are the materials and equipment to set up prior to the lab:

Make 1000 mL of 0.5 M copper(II) sulfate solution from solid copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$). This is enough for 6 lab groups.

1. Add approximately 400 mL of distilled water to a 1000-mL volumetric flask.

2. Add 124.8 g of $CuSO_4 \cdot 5H_2O$ to the flask and swirl until completely dissolved.

3. Dilute to the mark with distilled water.

4. Cap and invert the flask three times carefully to ensure complete mixing.

Teacher Tip: The waste generated in this experiment must be collected and disposed of according to your local, state, and federal regulations. It is suggested that used copper(II) sulfate be collected in a container and processed following the procedures outlined by your organization.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Use caution when operating the DC power supply. Injury to people or damage to equipment can occur due to electric shock. Refer to the instruction manual of your particular power supply for specific safety precautions.

- Never connect electrodes while the power supply is connected to an electrical outlet. Be sure the power supply is unplugged while attaching patch cords and alligator clips to the various components.

- Be sure that the power supply, data collection system, sensors, and general area around the experiment remain dry. Do not handle equipment with wet hands.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	4	5	3
Put the spoon and a piece of copper metal into a beaker of copper(II) sulfate solution.	Measure the mass of a clean, dry stainless steel spoon (or other metal to be plated).	Turn on the power and graph Current versus Time. Determine the area under the curve.	Determine the mass of the copper plated on the spoon and then calculate the molar mass of copper.	Connect the spoon and the copper electrode (that are in the beaker) to the power supply and the voltage-current sensor.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Add approximately 150 mL of 0.5 M copper(II) sulfate solution ($CuSO_4$) to a 250-mL beaker containing a magnetic stir bar.

2. ☐ Place the beaker on a magnetic stirrer.

3. ☐ After cleaning and drying a stainless steel spoon (or other metal to electroplate), measure and record its mass.

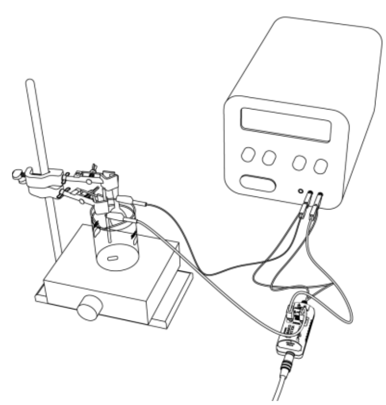
Mass of spoon (g): 41.12 g

4. ☐ Why does the spoon need to be clean and dry?

The spoon needs to be clean so that the layer of copper can adhere to the spoon. Oils can prevent the coating from sticking.

Also, the spoon needs to be dry because the mass of the spoon is subtracted from the mass of the spoon plus the copper plating. Any mass due to water introduces error in the calculation of how much copper is plated.

5. ☐ Use the ring stand and utility clamps to position the spoon and the copper electrode in the $CuSO_4$ solution so they are not touching the bottom of the beaker or each other.



6. ☐ Turn on the stirrer and set it to a low or medium speed. Make sure the stir bar does not hit the copper electrode or the spoon.

7. ☐ Why is it necessary to stir the electrolyte solution?

Stirring is necessary to ensure that the ions in the electrolyte solution are evenly distributed, so that the chemical reaction continues to occur.

8. ☐ Start a new experiment on the data collection system. �(1.2)

9. ☐ Connect the voltage-current sensor to the data collection system. �(2.1)

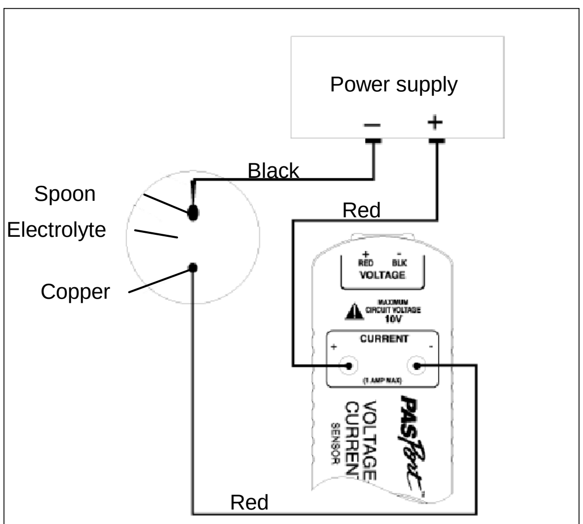
10. ☐ With the power supply unplugged, assemble the electroplating circuit by following the steps below:

Note: It is customary to use a red patch cord for the positive terminal and a black patch cord for the negative terminal of the power supply. The color of the cord, however, does not affect the lab results. The voltage sensor leads are not used during this experiment.

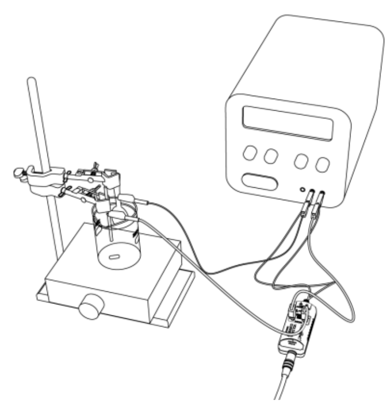
a. Plug a red patch cord into the negative terminal on the voltage-current sensor. Attach the other end of this red patch cord to the copper electrode using an alligator clip adapter.

b. Attach a black patch cord to the negative terminal of the DC power supply. Attach the other end of this black patch cord to the spoon using the second alligator clip.

c. Connect the second red patch cord to the positive terminal on the DC power supply. Attach the other end of this red patch cord to the positive terminal of the voltage-current sensor.



Circuit diagram



11. ☐ Double check that your circuit is set up correctly. A black cord connects the negative terminal of the power supply to the spoon. A red cord connects the copper electrode to the negative terminal of the sensor and a second red cord connects the positive terminal of the sensor to the positive terminal of the power supply.

Collect Data

12. ☐ Create a graph of Current (A) versus Time (s). �(7.1.1)

13. ☐ Start recording data. �(6.2)

14. ☐ Plug the power supply into a power outlet and set the power supply to output 0.4 amperes.

CAUTION: Do not exceed 0.4 A. This much current can be dangerous!

15. ☐ Adjust the scale of the graph so that you can clearly see the data being collected. �(7.1.2)

16. ☐ Why do you have to begin collecting data before you turn on the power supply?

You must begin collecting data before you turn on the power supply so that all the current that travels through the circuit is collected in the graph. If the power supply is turned on first, this data is not collected, which introduces error.

17. ☐ After 15 minutes, turn off and unplug the power supply.

18. ☐ Stop recording data. �(6.2)

19. ☐ Remove the spoon from the electrolyte solution and allow it to air dry.

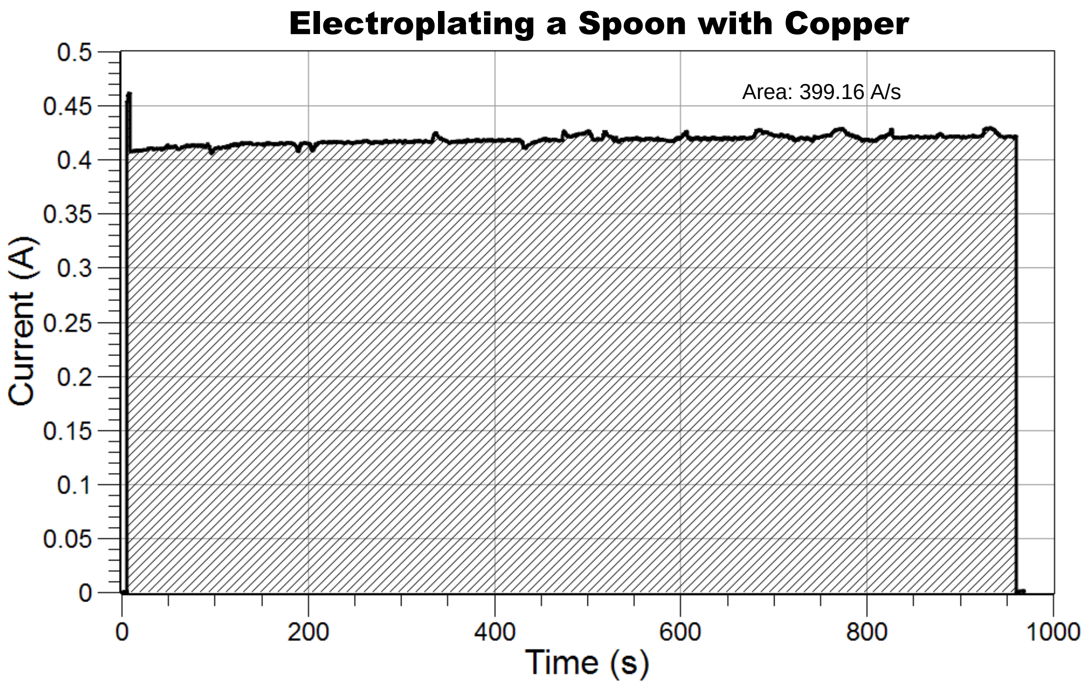
20. ☐ Measure and record the mass of the spoon that is now plated with copper.

Mass of copper plated spoon (g): ______________________ 41.26 g

21. ☐ Save the data file and clean up the lab station according to your teacher's instructions, including any special instructions for the disposal of the copper(II) sulfate solution. �(11.1)

Data Analysis

1. ☐ Sketch or print a copy of the graph of Current (A) versus Time (s). Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



2. ☐ Determine the area under the curve on the graph of Current (A) versus Time (s). �(9.7)

Area under the curve (A•s): 399.16 A•s

3. ☐ Using the area under the curve, calculate the moles of electrons transferred.

$$399.16 \text{ A} \times \text{s} \times \frac{1 \text{ mol } e^{-}}{96,485 \text{ A} \cdot \text{s}} = 4.1370 \times 10^{-3} \text{ mol } e^{-}$$

4. ☐ Give the balanced chemical equation for the reduction half-reaction used in this experiment.

$Cu^{2+}(aq) + 2\ e^{-} \rightarrow Cu^{0}(s)$

5. ☐ Using the stoichiometry of the reduction half-reaction, calculate the number of moles of copper plated onto the spoon.

$$4.1370 \times 10^{-3} \text{ mol } e^{-} \times \frac{1 \text{ mol } Cu^{0}(s)}{2 \text{ mol } e^{-}} = 2.0685 \times 10^{-3} \text{ mol } Cu^{0}(s)$$

6. ☐ Calculate the mass of copper that was plated onto the spoon.

mass of Cu = mass of copper plated spoon – mass of the spoon

mass of Cu = 41.26 g – 41.12 g = 0.14 g

7. ☐ Calculate the molar mass of copper using the number of moles of copper and the mass of copper determined in the above calculations.

$$\text{molar mass Cu} = \frac{\text{g Cu}}{\text{mol Cu}}$$

$$\frac{0.14\text{ g Cu}}{2.0685 \times 10^{-3}\text{ mol Cu}} = 67.682\text{ g/mol} \xrightarrow{\text{2 significant digits}} 68\text{ g/mol}$$

Analysis Questions

1. What is the percent error for the molar mass of copper calculated?

$$\textbf{percent error} = \left|\frac{\textbf{accepted value} - \textbf{experimental value}}{\textbf{accepted value}}\right| \times \mathbf{100}$$

$$\left|\frac{63.55 - 67.682}{63.55}\right| \times 100 = 6.502\% \xrightarrow{\text{2 significant digits}} 6.5\%$$

2. What are some possible combinations of amperage and time that would produce the same amount of copper plating as observed in the experiment?

Doubling the amperage decreases the amount of time by half. The product of each combination should produce 6 A•min (360 A•s). Combinations include:

0.1 A for 60 min = 0.2 A for 30 min = 0.4 A for 15 min = 0.6 A for 10 min = 1.0 A for 6 min

3. The electrolyte solution used in this experiment is 0.5 M copper(II) sulfate. What happens to the concentration of this electrolyte solution during the course of the reaction?

The concentration should remain the same. The copper(II) ions that are plated onto the spoon are replaced by new copper(II) ions from the copper electrode dissolving into the solution.

Synthesis Questions

Use available resources to help you answer the following questions.

1. If you stop collecting data before you turn off the power supply, where would the error show up? Would the moles of copper or the mass of copper you used to calculate the atomic mass be incorrect? Explain.

Some transferred electrons would not be counted. This makes the measured moles of copper that forms too small.

2. If you stop collecting data before you turn off the power supply, would your calculated atomic mass be too big or too small? Explain.

The moles of copper is in the denominator, so if the moles of copper is too small, the fraction, which is the atomic mass, would be too big.

3. Which electrode is the anode and which electrode is the cathode during this electroplating experiment? How do you know?

The spoon is the cathode, because this is where reduction occurs (copper ions gaining electrons to form solid copper).

The copper is the anode, because this is where oxidation occurs (copper from the electrode losing electrons and forming copper ions).

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

Note: Students need a table of standard electrode potentials to answer some of the following questions.

1. In the reaction between aluminum and chlorine, which substance is undergoing oxidation?

$$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$$

- **A. Aluminum (Al)**
- **B.** Aluminum chloride ($AlCl_3$)
- **C.** Chlorine (Cl_2)
- **D.** Not enough information

2. What process is occurring at the cathode?

- **A.** Oxidation
- **B.** Precipitation
- **C.** Redox
- **D. Reduction**

3. Using a table of standard reduction potentials, what is the calculated voltage associated with the spontaneous reaction between nickel and tin sulfate?

$$Ni(s) + SnSO_4(aq) \rightarrow Sn(s) + NiSO_4(aq)$$

- **A.** –0.39 V
- **B.** –0.11 V
- **C. +0.11 V**
- **D.** +0.39 V

4. What is the calculated number of moles of electrons transferred if the area under the curve is found to be 5566 A·s.

- **A. 5.769×10^{-2} mol e^-**
- **B.** 1.733×10^{1} mol e^-
- **C.** 5.566×10^{3} mol e^-
- **D.** 5.370×10^{8} mol e^-

5. Imagine another electroplating lab with an unknown metal. You know the metal ion has a 3+ charge. If the mass of the metal deposited is 0.500 g and the area under the curve is 5566 A·s, what is the identity of the metal?

- **A. Aluminum**
- **B.** Arsenic
- **C.** Chromium
- **D.** Iron

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. In a **reduction-oxidation** reaction, electrons are transferred between reactants. In the **reduction** half-reaction, ions **gain** electrons to become solid metal. In the **oxidation** half-reaction, metals **lose** electrons to become dissolved ions in solution. Because electroplating is a **non-spontaneous** process, electricity must be supplied to the reaction. The object to be plated is connected to the **cathode** where ions are reduced. The ions are replaced at the **anode** where metal atoms are oxidized.

2. **Electricity** is a flow of electrons. The rate of flowing electrons (current) is measured in **amperes**. Knowing the current and **time** the electricity is applied, you can determine the charge supplied to the reaction. Incorporating **the Faraday constant**, allows you to determine the number of moles of electrons transferred in the reaction. Using the **stoichiometry** of the reduction half-reaction then assists you in calculating the number of moles of metal electroplated.

Extended Inquiry Suggestions

Experiment with electroplating other metals and objects.

Compare the amount of metal (as moles) plated with the same current and time but using metals of different ionic charge, for instance using silver (1+), copper (2+), and aluminum (3+).

Determine the relationship between current and time on the amount of copper plated.

Even though aluminum is the third most abundant element in the crust of the Earth, it was, at one time, more expensive than both gold and platinum because it was difficult to purify. Investigate how electrolysis is now used to purify aluminum.

21. Double Replacement Reactions

Objectives

Determine the amount of chloride ion in water samples. Through this investigation, students:

- Are introduced to the process of dissolving
- Learn rules governing the solubility of ionic compounds in aqueous solutions
- Understand double replacement reactions
- Become familiar with net ionic equations
- Are introduced to the method of analytical titration

Procedural Overview

Students conduct the following procedures:

- Titrate known and unknown water samples, employing a double replacement reaction between silver nitrate and sodium chloride
- Calculate the concentration of chloride ion in the water samples

Time Requirement

♦ Preparation time	30 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	60 minutes

Materials and Equipment

For each student or group:

- Ring stand
- Buret clamp
- Buret, 50-mL
- Funnel
- Magnetic stirrer
- Magnetic stir bar
- Transfer pipet
- Waste container
- Erlenmeyer flask (4),125-mL
- Graduated cylinder, 50-mL
- 0.2% Disodium salt fluorescein indicator, 2 mL[1]
- 1% Dexrin solution, 100 mL[2]
- 0.020 M Silver nitrate ($AgNO_3$), 200 mL[3]
- 0.010 M Sodium chloride (NaCl), 100 mL[4]
- Swimming pool water, 100 mL[5]

[1] To formulate using fluorescein, disodium salt, refer to the Lab Preparation section.

[2] To formulate using dextrin, refer to the Lab Preparation section.

[3] To formulate using silver nitrate ($AgNO_3$), refer to the Lab Preparation section.

[4] To formulate using table salt (sodium chloride, NaCl), refer to the Lab Preparation section.

[5] If swimming pool water is not available, a substitute unknown may be created using table salt (sodium chloride, NaCl). Refer to the Lab Preparation section to prepare the unknown.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Evidence of chemical reactions
- Writing chemical equations
- Ionic compounds
- Molarity
- Mole calculations
- Stoichiometry

Related Labs in This Guide

Labs conceptually related to this one include:

- An Acid-Base Titration
- Electrolyte versus Non-electrolyte Solutions
- Evidence of a Chemical Reaction
- Single Replacement Reactions

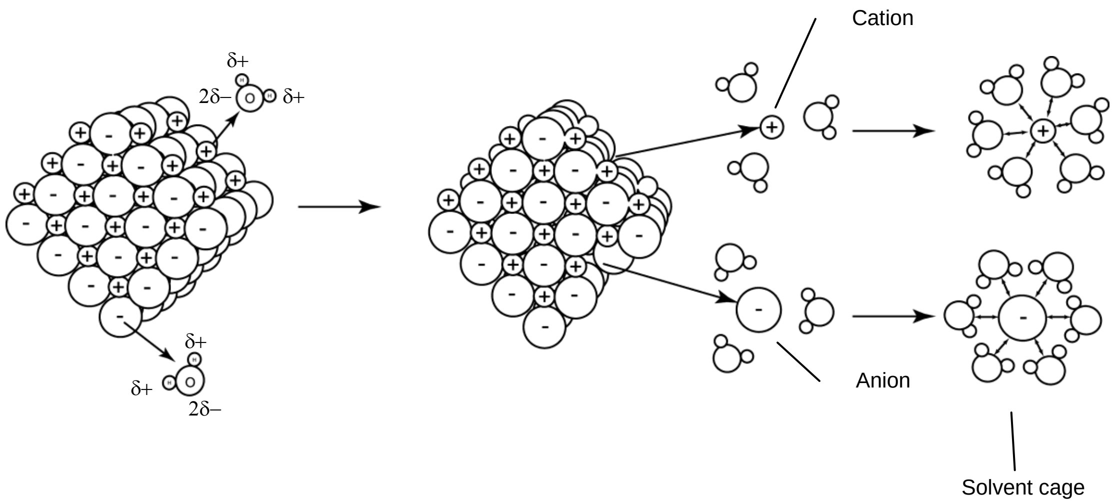
Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Note: There are no Tech Tips to list in this section as this activity does not use a data collection system.

Background

Dissolution is the process of dissolving a substance into a solvent to form a solution. Substances dissolve when solvent molecules completely surround individual components (molecules or ions) of the substance. Several solvent molecules surround the free particle creating a "solvent cage" around it (solvation), allowing that substance to be fully integrated into the solution.



Some substances do not dissolve because the attractions between the individual components of the substance are stronger than the attractions to the components of the solvent. They are insoluble in that solvent. For ionic salts with water as a solvent, a set of solubility rules assists in predicting which combinations of cations and anions produce insoluble ionic compounds. The solubility rules are listed in a Table at the end of the Background section.

In a double replacement reaction, the anions of two soluble reactants exchange cations. One type of double replacement reaction is a precipitation reaction between two ionic compounds. (The other is a neutralization reaction between an acid and a base.) Knowing the solubility rules is important for determining if a precipitation reaction can occur.

Reactions can occur only if they can form an insoluble product. A precipitate is the insoluble product that forms from two soluble reactants. For example, in the reaction between soluble silver nitrate and soluble sodium chloride, the exchange of cations produces silver chloride and sodium nitrate.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Following the solubility rules, it can be predicted that sodium nitrate is soluble, but silver chloride is insoluble. The silver chloride is the precipitate.

If a double replacement reaction cannot form an insoluble product, a reaction does not occur. For example, no reaction between sodium hydroxide and potassium chloride occurs, because the exchange of cations would produce sodium chloride and potassium hydroxide. As the solubility rules indicate, both of these are soluble, and a double replacement reaction must form at least one insoluble product.

$$NaOH(aq) + KCl(aq) \rightarrow NR \quad (NR = No\ Reaction)$$

No reaction occurs because the soluble ions do not undergo any change. When dissolved, sodium hydroxide produces sodium and hydroxide ions without precipitating as an insoluble solid. They remain sodium and hydroxide ions on the product side of the chemical equation. The same is true for the potassium chloride. These ions are referred to as “spectator ions.” Writing a net ionic equation removes the unchanging components (spectator ions) from the chemical equation, highlighting only that portion of the reaction that actually produces a new compound.

For the precipitation reaction with silver nitrate, the aqueous compounds can be rewritten as their soluble ions.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

Spectator ions do not change during the reaction. Removing them from both sides of the chemical equation (in this case, Na^+ and NO_3^-),gives a net ionic equation. It contains only the components that change.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

A titration is the quantitative method in which an exact amount of one reactant that is known is allowed to react with an unknown amount of a second reactant. The addition of known reactant is stopped when an indicator changes color, signifying that the unknown reactant is completely consumed.

In this experiment, an unknown amount of chloride ion is titrated with a known amount of silver ion. For example, if 25.0 mL of a water sample containing chloride ions required 10.0 mL of a 0.0500 M silver nitrate solution to change the color of the indicator from yellow to red, then the original concentration of chloride ion in the water sample can be calculated.

$$0.0100\ \text{L AgNO}_3 \times \frac{0.0500\ \text{mol AgNO}_3}{1\ \text{L AgNO}_3} \times \frac{1\ \text{mol Ag}^+}{1\ \text{mol AgNO}_3} = 0.000500\ \text{mol Ag}^+\ \text{added}$$

$$0.000500\ \text{mol Ag}^+ \times \frac{1\ \text{mol AgCl}}{1\ \text{mol Ag}^+} \times \frac{1\ \text{mol Cl}^-}{1\ \text{mol AgCl}} \times \frac{35.45\ \text{g Cl}^-}{1\ \text{mol Cl}^-} \times \frac{1000\ \text{mg}}{1\ \text{g}}$$

$$= 17.7\ \text{mg Cl}^-\ \text{reacted}$$

$$\frac{17.7\ \text{mg Cl}^-}{0.0250\ \text{L}} = 708\ \text{mg/L Cl}^-$$

Reporting the concentration in milligrams per liter (mg/L) is convenient because it is equivalent to a common unit, parts per million (ppm). Environmental sampling, for instance, often uses parts per million. One ppm equals one mg/L.

The unknown chloride solution tested in this experiment is a sample of treated swimming pool water. Students may know that chlorine is used to prevent bacterial growth. However, this form of chlorine (the hypochlorite ion, OCl^-) is not the chloride ion (Cl^-) being precipitated.

Swimming pool water has chloride ion added as calcium chloride ($CaCl_2$) used to adjust how hard the water is. Water that is too hard can build residue known as "scale." Water that is too soft can leach minerals from any available source, which can destroy grout and other surfaces over time.

Solubility rules

#	Rules	Examples	
1	Any compound with Na^+, K^+, Li^+, or ammonium (NH_4^+) cations or with acetate ($C_2H_3O_2^-$) or nitrate (NO_3^-) anions is *soluble*.	Soluble	NH_4Cl, KOH, $Mg(NO_3)_2$, $Ca(C_2H_3O_2)_2$
		Insoluble	None
2	Any compound with carbonate (CO_3^{2-}) or chromate (CrO_4^{2-}) or phosphate (PO_4^{3-}) anions is *insoluble*, *except* if with those cations listed in rule number 1.	Insoluble	$CaCrO_4$, $MgCO_3$
		Soluble	Na_2CrO_4 $(NH_4)_3PO_4$
3	All halides (F^-, Cl^-, Br^-, I^-, and At^-) are *soluble*, *except* those of Ag^+, Hg_2^{2+}, and Pb^{2+}.	Soluble	$CaCl_2$, AlI_3
		Insoluble	AgBr, $PbCl_2$
4	All sulfates (SO_4^{2-}) are *soluble*, *except* those of Sr^{2+}, Ba^{2+}, and Pb^{2+}.	Soluble	$CaSO_4$, $FeSO_4$
		Insoluble	$BaSO_4$, $PbSO_4$
5	All sulfides (S^{2-}) and hydroxides (OH^-) are *insoluble*, *except* those of Ca^{2+}, Sr^{2+}, and Ba^{2+} and those cations listed in rule number 1.	Insoluble	Fe_2S_3, $Mg(OH)_2$
		Soluble	$Ca(OH)_2$, Na_2S

Pre-Lab Discussion and Activity

Solubility Rules

Introduce the students to solubility by asking if sodium chloride dissolves in water. From their own experiences, they should recognize that sodium chloride (table salt) dissolves in water. Less familiar, is magnesium sulfate (Epsom salts), which is soluble, or lead(II) hydroxide, which is insoluble. Explain that not all ionic compounds (salts) dissolve in water. List the solubility rules presented in the Background section, above, on the board. Leave them displayed for students to reference.

1. For each of the compounds listed below, is it soluble or insoluble, what rule number applies, and what explanation applies based on the solubility rules?

Compound	Soluble or Insoluble	Rule	Explanation
Ammonium chloride	Soluble	1,3	Contains NH_4^+
Potassium hydroxide	Soluble	1,5	Contains K^+ even though it contains OH^-
Magnesium nitrate	Soluble	1	Contains NO_3^-
Calcium acetate	Soluble	1,5	Contains $C_2H_3O_2^-$
Calcium chromate	Insoluble	2	Contains CrO_4^{2-}
Magnesium carbonate	Insoluble	2	Contains CO_3^{2-}
Sodium chromate	Soluble	1,2	Contains Na^+ even though it also contains CrO_4^{2-}
Ammonium carbonate	Soluble	1,2	Contains NH_4^+ even though it also contains CO_3^{2-}
Silver bromide	Insoluble	3	Contains Br^- except that it also contains Ag^+
Lead(II) chloride	Insoluble	3	Contains Cl^- except that it also contains Pb^{2+}
Calcium chloride	Soluble	3	Contains Cl^-
Ammonium iodide	Soluble	3,1	Contains I^-
Calcium sulfate	Soluble	4	Contains SO_4^{2-}
Barium sulfate	Insoluble	4	Contains SO_4^{2-} except that it also contains Ba^{2+}
Iron(II) sulfate	Soluble	4	Contains SO_4^{2-}
Lead(II) sulfate	Insoluble	4	Contains SO_4^{2-} except that it also contains Pb^{2+}
Magnesium hydroxide	Insoluble	5	Contains OH^-
Calcium hydroxide	Soluble	5	Contains Ca^{2+} even though it contains OH^-
Iron(III) sulfide	Insoluble	5	Contains S^{2-}

Demonstrating Solubility

Using the solubility rules, ask the students if each of the following salts is soluble or insoluble: KNO_3, $NaCl$, $NaNO_3$, KCl, $CaCl_2$, Na_2CO_3, $CaCO_3$. Record their answers on the board.

In a set of small screw cap vials containing water, add a small portion of the solid salts to the vials. Shake to dissolve (all should dissolve except calcium carbonate).

2. Why does calcium carbonate not dissolve?

It is the only salt on the list that is insoluble. This is because compounds with carbonate (CO_3^{2-}) are usually insoluble (rule 2). Sodium carbonate is soluble because it contains a sodium ion (rule 1).

3. What is present in the other vials that contain the salts that do dissolve?

The vials containing the dissolved salts contain the dissolved (dissociated) ions from the ionic compounds. For example, $KNO_3(s) \rightarrow K^+(aq) + NO_3^-(aq)$.

Predicting Precipitation Reactions

Remind students of the evidence for a chemical reaction: energy is released or absorbed, significant color change, production of gas, or the formation of a precipitate. Write the molecular equation for potassium nitrate and sodium chloride on the board, explaining how the products are a result of the two salts exchanging anions.

$$KNO_3 + NaCl \rightarrow NaNO_3 + KCl$$

Refer back to the previous results pointing out that both products are also soluble. Mix a portion of the potassium nitrate solution with a portion of the sodium chloride solution in a clean, empty vial. Since no evidence is produced, it can be concluded that a reaction did not occur.

Write the molecular equation for calcium chloride and sodium carbonate on the board. Again, indicate that the products are a result of the two salts exchanging anions, but this time one of the products is insoluble, forming a precipitate (an insoluble solid produced from soluble reactants). This is a double replacement that results in a reaction.

$$CaCl_2 + Na_2CO_3 \rightarrow 2NaCl + CaCO_3$$

Mix a portion of the calcium chloride solution with a portion of the sodium carbonate solution in a clean, empty vial. Note the color of the solid precipitate and compare it with the demonstration involving the calcium carbonate solution that did not dissolve.

4. What is present in the vial containing the reaction of potassium nitrate with sodium chloride?

The soluble ions K^+, Na^+, Cl^-, and NO_3^- are present.

5. What is present in the vial containing the reaction of calcium chloride and sodium carbonate?

The insoluble precipitate $CaCO_3$ as well as the soluble ions Na^+ and Cl^- are present.

Net Ionic Equations

Rewrite the calcium carbonate reaction, explicitly showing all of the ions on both sides of the equation (ionic equation). Note how the sodium and chloride ions remain unchanged. By removing these spectator ions, only the portion of the reaction that undergoes a change is highlighted. The resulting chemical equation is known as a net ionic equation.

Molecular equation: $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$

Ionic equation: $Ca^{2+}(aq) + \cancel{2Cl^-}(aq) + \cancel{2Na^+}(aq) + CO_3^{2-}(aq) \rightarrow \cancel{2Na^+}(aq) + \cancel{2Cl^-}(aq) + CaCO_3(s)$

Net ionic equation: $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

Titrations

Hold up a small Erlenmeyer flask containing 25.0 mL of sodium chloride (0.05 M, although the students do not know this) in 1% dextrin solution. Ask the students how much chloride ion is present. The only way to know is to react all of the chloride until it is completely consumed using a precipitation reaction. Adding an indicator signals when we have added just enough reactant to completely consume the chloride ions. If we know how much reactant we added, we can calculate the amount of chloride originally present. Add 1 or 2 drops of fluorescein indicator to the sodium chloride solution and swirl to mix thoroughly.

Begin adding and stirring in 1.0 M silver nitrate, drop by drop, to the sodium chloride solution until the indicator changes from yellow to red (approximately 25 drops, or 1.25 mL). Using a few calculations, we can convert the amount of silver nitrate added to the amount of chloride ion originally present.

Net ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

$$(25\ \text{drops})\left(\frac{1\ \text{mL}}{\sim\ 20\ \text{drops}}\right)\left(\frac{1.0\ \text{mol Ag}^+}{1\ \text{L}}\right)\left(\frac{1\ \text{L}}{1000\ \text{mL}}\right) = 0.00125\ \text{mol Ag}^+$$

$$(0.00125\ \text{mol Ag}^+)\left(\frac{1\ \text{mol Cl}^-}{1\ \text{mol Ag}^+}\right)\left(\frac{35.45\ \text{g Cl}^-}{\text{mol Cl}^-}\right) = 0.044\ \text{g Cl}^-\ \text{present}$$

$$\frac{0.044\ \text{g Cl}^-}{25.0\ \text{mL}} = 0.0018\ \text{g/mL Cl}^-$$

Lab Preparation

These are the materials and equipment to set up prior to the lab.

1. The following instructions make 100 mL of 0.2% fluorescein indicator. This is enough for 50 lab groups.

- **a.** Add approximately 20 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 0.2 g fluorescein, disodium salt to the flask and swirl to dissolve.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert the flask three times carefully to ensure complete mixing.

Teacher Tip: Fluorescein is a very potent, but relatively safe dye (Yellow No. 8). With contact, it stains most items. Gloves and lab aprons are suggested, as is bench liner to prevent stains to work surfaces.

2. The following instructions make 1000 mL of 1% dextrin (starch) solution. This is enough for 10 lab groups.

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 1.00 g of dextrin to the flask and allow it to dissolve completely with continuous swirling.

c. Dilute to the mark with distilled water.

d. Cap and invert the flask three times carefully to ensure complete mixing.

3. The following instructions make 1000 mL of 0.020 M silver nitrate ($AgNO_3$) from either solid $AgNO_3$ or 0.050 M $AgNO_3$. This is enough for 5 lab groups.

Starting with solid AgNO3

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 3.40 g of $AgNO_3$ to the flask and allow it to dissolve completely with continuous swirling.

c. Dilute to the mark with distilled water.

d. Cap and invert the flask three times carefully to ensure complete mixing.

Starting with 0.050 M AgNO3

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 400.0 mL 0.050 M $AgNO_3$ to the flask.

c. Dilute to the mark with distilled water.

d. Cap and invert the flask three times carefully to ensure complete mixing.

Teacher Tip: Silver nitrate is a light sensitive compound. If the solution is to be stored for any length of time, it should be kept in a dark bottle or one that has been wrapped in aluminum foil. Avoid getting any of the solution on the skin. If this happens, wash the area with soap and water. While there is minimal threat to health, any remaining solution may blacken considerably and stain the skin for one to three days when exposed to sunlight. Do not attempt to remove the stains through scrubbing, detergents, or solvents.

4. The following instructions make 1000 mL of 0.010 M sodium chloride (NaCl) from solid NaCl. This is enough for 10 lab groups.

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 0.58 g of NaCl to the flask and allow it to dissolve completely with continuous swirling.

c. Dilute to the mark with distilled water.

d. Cap and invert the flask three times carefully to ensure complete mixing.

5. It is highly recommended that a titration of the swimming pool water is performed in advance in order to assure the chloride concentration is great enough to be used as an unknown in the experiment, as well as to determine the accepted value for comparison purposes.

If you do not have access to a swimming pool, you can use the following instructions make 1000 mL of 0.015 M sodium chloride (NaCl) unknown from solid NaCl as a substitute for swimming pool water. This is enough for 10 lab groups.

- **a.** Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.
- **b.** Add 0.88 g of NaCl to the flask and allow it to dissolve completely with continuous swirling.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert the flask three times carefully to ensure complete mixing.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Silver nitrate ($AgNO_3$) and the resulting silver chloride (AgCl) precipitate in this lab are hazardous to the environment and should not be disposed down the drain. Make sure you follow your teacher's instructions on how to properly dispose of the chemicals.
- ♦ The fluorescein indicator is a dye and may stain your skin or clothes. The silver nitrate ($AgNO_3$) solution may also stain your skin and clothes, especially after being exposed to bright sunlight. If either solution contacts your skin, wash it with soap and water immediately.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

1	2	4	3	5
Clamp a buret to the ring stand and fill the buret with 50 mL of silver nitrate solution. Record the initial buret volume.	Add 50.0 mL of sample solution, dextrin, and the indicator to the flask. Place the flask under the buret filled with silver nitrate.	Stop adding titrant when the solution turns pink.	Slowly add titrant to the flask.	Record the final buret volume in the data table. Use the results to calculate the concentration of chloride ion in the sample solution.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Set Up

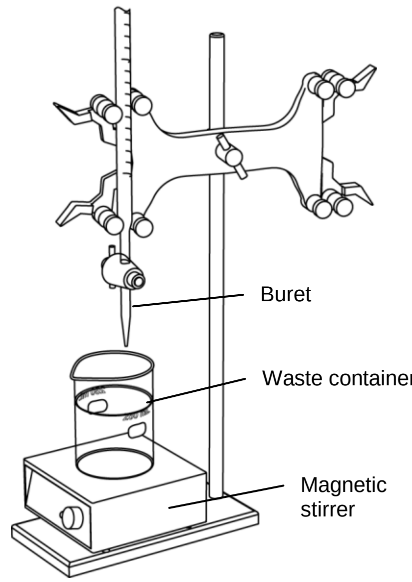
1. ☐ Set up the titration apparatus, using the illustration as a guide.

a. Assemble the ring stand.

b. Position the magnetic stirrer on (or next to) the base of the ring stand.

c. Place a waste container on the magnetic stirrer.

d. Use the buret clamp to attach the buret to the ring stand.



2. ☐ Make sure the stopcock on the buret is off.

3. ☐ With a funnel, fill the buret with approximately 50 mL of $AgNO_3$ solution (titrant).

4. ☐ Drain a small amount of the titrant into the waste beaker to remove any air in the tip of the buret.

5. ☐ Why is it important to remove air from the tip of the buret?

Any air trapped in the buret tip is counted as volume of $AgNO_3$. If this happens, the final amount of titrant used is inaccurate.

6. ☐ Place a magnetic stir bar in a clean 125-mL Erlenmeyer flask.

7. ☐ Using the 50-mL graduated cylinder, measure 50.0 mL of sodium chloride known solution (0.010 M) and add it to the Erlenmeyer flask.

8. ☐ Why is a solution with a known concentration being tested first?

It is being used to practice the titration technique as well as to ensure the apparatus and chemical reagents are set up and prepared properly.

9. ☐ Using the same 50-mL graduated cylinder, measure 25 mL of 1% dextrin solution and add it to the Erlenmeyer flask containing the sodium chloride solution.

10. ☐ Using a transfer pipet, add 1 or 2 drops of the fluorescein indicator to the solution in the flask.

Collect Data

11. ☐ Read the initial volume of the $AgNO_3$ solution that is in the buret. Record this value in Table 2 below, in the NaCl Known, Trial 1 column (record within 0.01 mL).

Table 2: Initial and final buret readings

	NaCl Known (0.010 M)		Swimming Pool Water	
	Trial 1	Trial 2	Trial 1	Trial 2
Initial Buret Volume (mL)	0.03	0.01	0.00	0.01
Final Buret Volume (mL)	24.89	24.71	36.51	36.76

12. ☐ Position the flask containing the sodium chloride solution on the magnetic stirrer.

13. ☐ Turn on the magnetic stirrer and begin stirring at a slow to medium speed.

14. ☐ Why is it necessary to stir the solution during a titration?

Stirring ensures that the ions in the solution are thoroughly mixed so that the indicator changes color for the entire solution.

15. ☐ Open the stopcock on the buret carefully, allowing the titrant to drip about 2 or 3 per second into the sodium chloride solution.

16. ☐ Reduce the flow of the titrant when the solution becomes cloudy.
Note: The solution becoming cloudy occurs when the solution begins to precipitate.

17. ☐ Continue to add titrant slowly until the solution turns pink. When the solution turns pink, close the stopcock.

18. ☐ What substance is being formed in the beaker? What type of reaction is occurring?

A precipitate of silver chloride, AgCl, is being formed. The reaction is a double replacement reaction.

19. ☐ What equation represents the chemical reaction that is occurring in the beaker?

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl\ (s) + NaNO_3(aq)$

20. ☐ Record the final volume of $AgNO_3$ that is in the buret. Record this value in Table 2 above, in the NaCl Known, Trial 1 column (record within 0.01 mL).

21. ☐ Turn off the magnetic stirrer.

22. ☐ Remove the flask and dispose of its contents according to your teacher's instructions.

23. ☐ Clean the flask and the magnetic stir bar with distilled water.

24. ☐ For the "NaCl Known Trial 2," repeat the titration so that you can record the initial and final volumes in the buret of the titrant $AgNO_3$. Refer to the process above for details.

25. ☐ For the "Swimming Pool Water Trial 1," repeat the titration so that you can record the initial and final volumes in the buret of the titrant $AgNO_3$. Refer to the process above for details.

26. ☐ For the "Swimming Pool Water Trial 2," repeat the titration so that you can record the initial and final volumes in the buret of the titrant $AgNO_3$. Refer to the process above for details.

27. ☐ Clean your lab station according to your teacher's instructions, especially those regarding unused and waste chemicals.

Data Analysis

1. ☐ Determine the volume of titrant added to the flask by subtracting the initial volume from the final volume. Record this value in the Table 3 below.

Table 3: Concentration of chloride ions

	NaCl Known (0.010 M)		Swimming Pool Water	
	Trial 1	**Trial 2**	**Trial 1**	**Trial 2**
Total silver nitrate added, final minus intial (mL)	24.86	24.70	36.51	36.75
Chloride ion in flask (mg)	17.6	17.5	25.9	26.1
Chloride ion in 1000 mL of sample (mg/L)	352	350	518	522
Average chloride ion concentration (mg/L)	351		520	

2. ☐ Determine the amount of chloride ion in each flask using the following hint. Record the values in Table 3 above. Show your work for the NaCl Known, Trial 1, below.
Hint:

$$(\text{AgNO}_3\text{ added})\left(\left[\text{Ag}^+\right]\right)\left(\text{mole ratio of Cl}^-\text{:Ag}^+\right)\left(\text{molar mass Cl}^-\right)\left(\text{conversion to mg}\right) = \text{mg Cl}^-$$

$$\left(0.02486\text{ mL AgNO}_3\right)\left(\frac{0.0200\text{ mol Ag}^+}{1\text{ L}}\right)\left(\frac{1\text{ mol Cl}^-}{1\text{ mol Ag}^+}\right)\left(\frac{35.45\text{ g Cl}^-}{1\text{ mol Cl}^-}\right)\left(\frac{1000\text{ mg}}{1\text{ g}}\right) = 17.6\text{ mg Cl}^-$$

3. ☐ Determine the concentration of chloride in each flask by dividing the amount of Cl^- by the volume of the water sample (50.0 mL). You will also have to convert from milliliters to liters. Record the values in Table 3 above. Show your work for the NaCl Known, Trial 1, below.

$$\left(\frac{17.6\text{ mg Cl}^-}{50.0\text{ mL}}\right)\left(\frac{1000\text{ mL}}{1\text{ L}}\right) = 352\text{ mg/L Cl}^-$$

4. ☐ Calculate the average concentration of Cl^- for the different water samples by adding the values of the two trials and dividing by the number of trials (2). Record this value in Table 3 above. Show your work for the NaCl Known, Trials 1 and 2, below.

$$\frac{\left(\text{Cl}^-\text{ concentration Trial 1} + \text{Cl}^-\text{concentration Trial 2}\right)}{2} = \text{average concentration of Cl}^-$$

$$\frac{\left(352\text{ mg/L Cl}^- + 350\text{ mg/L Cl}^-\right)}{2} = 351\text{ mg/L Cl}^-$$

5. ☐ The chloride concentration for the NaCl Known sample is 355 mg/L. Using your calculated average for the NaCl Known, report your percent error for this experiment.

$$\text{percent error} = \left|\frac{(\text{accepted value} - \text{experimental value})}{\text{accepted value}}\right| \times 100$$

$$\text{percent error} = \left|\frac{\left(351\text{ mg/L Cl}^- - 355\text{ mg/L Cl}^-\right)}{355\text{ mg/L Cl}^-}\right| \times 100 = 1.13\%$$

Analysis Questions

1. For the titration of the swimming pool water chloride, what are the equations?

a. Molecular equation:

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

b. Ionic equation:

$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

c. Net ionic equation:

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

2. What is it likelihood that the amount of silver ion and chloride ion is exactly equal at the point at which the solution turns pink and the titration is stopped (equivalence point)? Explain your reasoning.

It is highly unlikely that, at the equivalence point, the amount of silver ion added exactly equals the chloride ion that was originally present. This is because the smallest amount of silver ion that can be added to the flask with the analyte (the chloride ion) is one drop. All of the Ag^+ in the last drop added to get to the point of the color change may not have been necessary to react with the last amount of Cl^-.

3. If additional silver nitrate is added after the indicator changes color, what effect would this have on the calculated results for the chloride ion concentration in the sample?

If too much silver nitrate is added to the sample, the resulting concentration of chloride ion in the sample is erroneously high.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Based on the results for both the sodium chloride known and the swimming pool water chloride, what are their concentrations converted into parts per million (ppm)?

1 mg/L is the same as 1 ppm, so:

$$\text{Known}: (351\ \mathrm{mg/L})\left(\frac{1\ \text{ppm}}{1\ \mathrm{mg/L}}\right) = 351\ \text{ppm}\ Cl^-$$

$$\text{Pool water}: (520\ \mathrm{mg/L})\left(\frac{1\ \text{ppm}}{1\ \mathrm{mg/L}}\right) = 520\ \text{ppm}\ Cl^-$$

2. Based on the results for both the sodium chloride known and the swimming pool water chloride, what are their concentrations converted into units of molality?

$$\text{Known}: \left(\frac{351\ \text{mg}\ Cl^-}{\text{L}}\right)\left(\frac{1\ \text{g}\ Cl^-}{1000\ \text{mg}\ Cl^-}\right)\left(\frac{1\ \text{mol}\ Cl^-}{35.45\ \text{g}\ Cl^-}\right) = 0.00990\ \mathrm{mol/L}\ Cl^- = 0.00990\ \text{M}\ Cl^-$$

$$\text{Pool Water}: \left(\frac{520\ \text{mg}\ Cl^-}{\text{L}}\right)\left(\frac{1\ \text{g}\ Cl^-}{1000\ \text{mg}\ Cl^-}\right)\left(\frac{1\ \text{mol}\ Cl^-}{35.45\ \text{g}\ Cl^-}\right) = 0.0147\ \mathrm{mol/L}\ Cl^- = 0.0147\ \text{M}\ Cl^-$$

3. What is the source of chloride ions in the swimming pool water? Is it the same as the chlorine used to kill bacteria?

Calcium chloride is added to swimming pool water to control its hardness. It is not the same as the chlorine used to kill bacteria (hypochlorite ion, ClO^-).

4. Why are multiple trials of the same sample analyzed and then averaged together? What effect do multiple trials have on the precision of the results?

Every experiment includes errors associated with the instruments and the method used to collect the data. By repeating the experiment, error can be minimized by averaging the results. Increasing the number of trials can decrease the uncertainty and optimize the precision of the results by eliminating much of the experimental error.

5. What is the maximum chloride ion concentration that can be determined by one buret (50.00 mL) of the silver nitrate solution? How would you analyze a sample that had a greater chloride concentration?

$$(50.00 \text{ mL } AgNO_3)\left(\frac{0.0200 \text{ mol } Ag^+}{1000 \text{ mL } AgNO_3}\right)\left(\frac{1 \text{ mol } Cl^-}{1 \text{ mol } Ag^+}\right)\left(\frac{35.45 \text{ g } Cl^-}{1 \text{ mol } Cl^-}\right)\left(\frac{1000 \text{ mg}}{1 \text{ g}}\right) = 35.45 \text{ mg } Cl^-$$

$$\left(\frac{35.45 \text{ mg } Cl^-}{50.0 \text{ mL}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 709 \text{ mg/L } Cl^-$$

In order to analyze samples with concentrations greater than 709 mg/L, you could either refill the buret multiple times and keep track of the total volume of silver nitrate added or use a smaller sample size.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

Note: Students will need a copy of the solubility rules to answer some of the following questions.

1. Which of the following reactions represents a double replacement?

- **A.** $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
- **B.** $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
- **C.** $\mathbf{Pb(C_2H_3O_2)_2 + Na_2CrO_4 \rightarrow 2NaC_2H_3O_2 + PbCrO_4}$
- **D.** $CuSO_4 + Zn \rightarrow Cu + ZnSO_4$

2. Which of the following compounds is soluble in water?

- **A.** $CaCO_3$
- **B.** $\mathbf{CaI_2}$
- **C.** AgI
- **D.** PbI_2

3. Which of the following compounds is *not* soluble in water?

- **A.** $LiOH$
- **B.** $\mathbf{FePO_4}$
- **C.** $(NH_4)_2S$
- **D.** $NiSO_4$

4. Which two reactants form a precipitate when mixed in solution?

- **A.** $\mathbf{2NH_4OH + BaI_2 \rightarrow 2NH_4I + Ba(OH)_2}$
- **B.** $Li_2S + Ca(C_2H_3O_2)_2 \rightarrow 2LiC_2H_3O_2 + CaS$
- **C.** $3KCl + FeI_3 \rightarrow 3KI + FeCl_3$
- **D.** $Mg(OH)_2 + Na_2SO_4 \rightarrow 2NaOH + MgSO_4$

5. Which of the following reactants forms a precipitate when mixed with NaCl in solution?

- **A.** $(NH_4)_2SO_4$
- **B.** **$Pb(NO_3)_2$**
- **C.** $Fe_2(OH)_3$
- **D.** None of the above

Key Term Challenge

Fill in the blanks from the list of randomly ordered words in the Key Term Challenge Word Bank.

1. Substances that dissolve have **stronger** attractions with solvent particles than with themselves. Substances that do not dissolve are **insoluble**. The solubility rules assist chemists in predicting if a substance dissolves in water. For example, any compound having Na^+, K^+, Li^+, NH_4^+, $C_2H_3O_2^-$, or NO_3^- ions is **soluble**. There are no exceptions for these components. For other components, some exceptions apply. For instance, all **halides** are soluble except those of Ag^+, Hg_2^{2+}, and Pb^{2+}.

2. In a **double replacement** reaction, the reactants switch anions to form new products. The reaction takes place only if one of the resulting products is insoluble. The insoluble product formed is called a **precipitate**. Knowing the solubility rules assist in predicting the products in a double replacement reaction. By using a **net ionic** equation, only those ions and products undergoing a change in the reaction are shown in the written chemical equation.

Extended Inquiry Suggestions

Develop and test methods for determining the chloride concentrations of various water samples, for instance, from a river, stream, lake, pond, or ocean.

Create a series of NaCl salt dilutions ranging from 0 to 700 mg/L. Have students arrange them in order of increasing concentration based on taste. Titrate a sample of each to verify their predictions.

Visit a water treatment facility or invite a water treatment plant operator to talk about their occupation as well as the daily tests performed to ensure a safe, clean supply of water to the community.

22. Rates of Reaction

Objectives

Determine the effect of temperature, concentration, and surface area on the rate of a chemical reaction. Through this investigation, students:

- ♦ Learn the effects of temperature, concentration, and particle size on the rate at which a chemical reaction occurs.
- ♦ Measure changes in absolute pressure as a reaction proceeds

Procedural Overview

Students conduct the following procedures:

- ♦ Record the absolute pressure as magnesium ribbon reacts with hydrochloric acid (HCl) at different temperatures
- ♦ Record the absolute pressure as magnesium powder reacts with HCl
- ♦ Record the absolute pressure as magnesium ribbon reacts with different concentrations of HCl
- ♦ Determine the rates of the reaction between magnesium and HCl for the different reaction parameters
- ♦ Draw conclusions about how temperature, surface area, and concentration effect the rate at which magnesium reacts with HCl

Time Requirement

♦ Preparation time	30 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	60 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Absolute pressure sensor
- Sensor extension cable
- Test tube (3), 20-mm x 150-mm
- Test tube rack
- One-hole stopper to fit the test tubes
- Quick-release connector[1]
- Tubing connector[1]
- Tubing, 1- to 2-cm[1]
- Glycerin
- 4.0 M Hydrochloric acid (HCl), 5 mL [2]
- 2.0 M Hydrochloric acid (HCl), 5 mL[3]
- 1.0 M Hydrochloric acid (HCl), 20 mL[3]
- 0.1 M Hydrochloric acid (HCl), 5 mL[3]
- Warm and cold water baths[4] (one per class)
- Magnesium ribbon (18), 1-cm pieces
- Magnesium powder, 0.05 g

[1]Included with most PASCO absolute pressure sensors.

[2]To prepare using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section, below.

[3]To prepare using a series of dilutions from 4.0 M hydrochloric acid (HCl), refer to the Lab Preparation section, below.

[4]For instructions about how to prepare warm and cold water baths, refer to the Lab Preparation section, below.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Kinetic molecular theory
- Evidence of a chemical reaction
- Pressure

Related Labs in This Guide

Labs conceptually related to this one include:

- Evidence of a Chemical Reaction
- Conservation of Matter
- Boyle's Law

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting sensors to the data collection system �(2.1)
- Displaying data on a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Showing and hiding runs in a graph �(7.1.7)
- Recording a run of data �(6.2)
- Naming runs of data �(8.2)
- Applying a linear fit to the data �(9.5)
- Finding the slope and intercept of a best fit line �(9.6)
- Saving data files �(11.1)
- Printing �(11.2)

Background

Kinetics is the branch of chemistry that studies the rates of chemical reactions. Some reactions, such as lighting a match, have high rates of reaction and occur very quickly. Other reactions, such as the formation of rust, have low rates of reaction and occur more slowly.

In order for a reaction to occur the reacting particles must collide together with a sufficient amount of energy. This is known as collision theory and is used to explain reaction rates. The rate of a chemical reaction is proportional to the number of times the reacting particles collide with sufficient amounts of energy. The rate of a chemical reaction can be increased by increasing the number of collisions between reacting particles; this can be accomplished by increasing the temperature of the reaction, the concentration of the reactants, or the surface area of the reactants.

By increasing the temperature of the reaction, the kinetic energy of the system is increased. This results in the particles moving at greater speeds. With increased speed comes an increase in the number of collisions in a given amount of time. Additionally, an increase in temperature also gives a greater proportion of the reactant particles the energy required to react, overcoming the activation energy barrier.

When you increase the concentration of one or more of the reactants, the probability of the reactants coming in contact with one another increases. A greater probability of collisions in a given amount of time increases the rate of reaction.

Reactants must physically come into contact with one another. Therefore, increasing the surface area of one or more of the reactants increases the likelihood that a substance is available to collide with another reactant.

In this experiment, students will be investigating various factors that affect the reaction of magnesium metal with hydrochloric acid. The chemical equation for this reaction is given below.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

When added to hydrochloric acid, magnesium metal reacts immediately and generates hydrogen gas. The hydrogen gas that is evolved causes the pressure inside the reaction vessel (a test tube) to increase. The rate at which the pressure increases is related to the rate of production of hydrogen gas, which, in turn, is proportional to the rate of the overall reaction. Students can assess the rate by applying a linear fit to the initial linear portion of the pressure graph in each experiment.

Pre–Lab Discussion and Activity

Engage your students with the following activities and questions.

Collision Theory and Rates of Reaction

Engage the students in a discussion about what occurs at the molecular level during a chemical reaction. Review the kinetic molecular theory and introduce the collision theory. Explain that reactant particles need to collide with enough energy for products to be formed. The rate of collisions between reactants is what determines the rate of a reaction. The idea of activation energy and transition states might or might not be discussed in conjunction with this experiment. It might add some depth to the activity, but it is not required in order to understand the factors governing the rates of reactions.

1. What is the kinetic molecular theory and how does it relate to chemical reactions?

The kinetic molecular theory explains that matter is made up of particles that are in constant motion. In a chemical reaction, reactant particles must collide with each other for the products to form.

2. What does the collision theory explain?

The collision theory explains that the rate of a chemical reaction depends on the number of collisions (with sufficient energy) between the reacting particles that occur in a given amount of time.

3. What is meant by "reaction rate"? What are other rates with which you are familiar?

Reaction rate is the speed at which a reaction occurs, or the time it takes for the reaction to occur. Another rate that students are likely to be familiar with is automobile speed (time it takes a car to travel a certain distance).

4. What are some examples of chemical reactions that occur quickly, or have high rates of reaction?

Lighting a match, burning, formation of precipitates, baking soda reacting with vinegar, reacting sodium with water, explosions, and other similar reactions are possible examples.

5. What are some examples of chemical reactions that occur slowly, or have low reaction rates?

Decomposing plants and animals, formation of fossil fuels from dead plants and animals, the formation of rust, and other similar reactions are possible examples.

6. At the molecular level, what needs to happen for the rate of a reaction to increase?

The number of collisions (with sufficient energy) between the reactants needs to increase.

7. How can the number of collisions be increased? Why do you think this will work?

Increasing the temperature of the reactants causes the number of collisions, and, therefore, the rate of reaction, to increase. Increasing the temperature of the reactants works because temperature is a measure of the average molecular motion of the particles. If the particles are moving faster, they collide more often and they have more energy when they collide.

Other answers include increasing the surface area of the reactants and increasing the concentration of the reactants.

Teacher Tip: Use this question as an introduction to the next section. The students might or might not be able to answer it at this time.

Influencing Reaction Rates

To demonstrate the effects of concentration and temperature on the rate of reaction, prepare the following two solutions in advance: 0.5 g of sodium metabisulfite with 2 g starch in 200 mL of distilled water (prepared by first dissolving the starch in warm water and then allowing it to cool to room temperature) and 1.28 g of potassium iodate in 300 mL of distilled water. Transfer two 10-mL aliquots of the starch solution to two separate large test tubes (one in a warm water bath, the other in an ice bath); repeat for the potassium iodate solution. These four test tubes will be used in the temperature demonstration. The remaining starch and potassium iodate solutions should be transferred to separate bottles to be used in the concentration demonstration.

Demonstrate the effects of temperature on the rate of reaction by performing a Landolt (iodine clock) reaction. Remove the test tubes from the warm water bath and put them into a test tube rack. Do the same for the test tubes in the ice bath. Simultaneously pour the warm starch solution into the warm potassium iodate solution while pouring the cold starch solution into the cold potassium iodate solution. Ask the students to count aloud and record the time to react on the board. The colder reactants should require a longer time to react.

8. At the molecular level, what is different between the warm and cold solutions?

The kinetic energies of the two solutions are different. The warm solution has higher kinetic energy; the particles are moving at a greater speed. The cold solution has lower kinetic energy; the particles are moving at a slower speed.

9. Which combination of reactants has the higher rate of reaction? Explain using collision theory.

The warmer reactants have the higher rate of reaction because the color appears first in the warm solution. This is because the reactant particles are moving faster and experience more collisions, and therefore have more opportunities to react and form products.

Demonstrate the effects of concentration on the rate of reaction. Show the bottle containing the starch solution. Show the other bottle containing the potassium iodate solution. In front of the students, pour 2/3 of the potassium iodate solution into one 500-mL beaker and the remaining 1/3 of the solution into another 500 mL beaker. Add enough distilled water to the second beaker to make the volumes equivalent (essentially halving the concentration of the potassium iodate in the second beaker). In two additional 500-mL beakers, divide the starch solution equally. Simultaneously pour the contents from each beaker containing the starch solutions into the different beakers containing the potassium iodate solutions.

Ask the students to count aloud and record the time to react on the board. The diluted reactant (half the concentration) should require twice as long to react.

10. At the molecular level, what is different between the two solutions?

The solution that has water added to it is less concentrated than the other solution. In the more concentrated solution, the particles are closer together.

11. Which combination of reactants has the higher rate of reaction? Explain using collision theory.

The more concentrated reactants have the higher rate of reaction because the color appears first in the concentrated solution. This is because the reactant particles are closer together and experience more collisions, and, therefore, have more opportunities to react and form products.

Teacher Tip: A helpful analogy is to describe the difficulty you might experience in trying to cross a crowded dance floor compared to an empty room. The likelihood of bumping into someone increases as the concentration of people increases.

Demonstrate the effect of surface area on the rate of reaction. Be sure all combustible materials within two meters of the demonstration area are removed. Make a small pile of lycopodium powder on a watch glass and try to light it with a match. The pile will not ignite. Create a fine cloud of suspended lycopodium powder particles by squeezing a wash bottle containing lycopodium powder through a burner or candle flame. The resulting fireball demonstrates the effects of the increased surface area for combustion to take place. This is a small scale example of the dangers associated with grain elevator explosions.

12. What is different between the pile and the stream of powder?

In the pile, the particles are packed together; in the stream, the particles have space between them.

13. Which combination of reactants has the higher rate of reaction? Explain using collision theory.

The particles that are separated from each other have the higher rate of reaction because the powder burns easily. This is because the particles are exposed to oxygen and can react (burn) on all sides, essentially increasing the availability of reaction sites. The surfaces in the separated powder are more exposed to the oxygen (reactant) than those in the pile, and, therefore, have more opportunities to react and form products.

Measuring Reaction Rates

While holding an inverted small test tube over the opening, drop a piece of magnesium into a test tube of 1.0 M hydrochloric acid. As you do this, explain to the students that this is the reaction they will be performing in the lab (the equation is given below). While the inverted test tube is filling with hydrogen product, engage the students in a review about evidence for a chemical reaction and single replacement reactions as they apply to this situation.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Confirm the presence of hydrogen gas as the product of this reaction by igniting the contents of the inverted test tube with a glowing splint. Discuss with the students how it will be possible to measure the reaction rate of this reaction by monitoring the production of hydrogen gas. Help the students realize that a rate involves measuring both the extent to which the reaction has occurred and the time it takes for the reaction to occur. End the discussion with a review about how a rate can be determined from a graph.

14. What two components must be measured in the lab to determine the rate of a reaction?

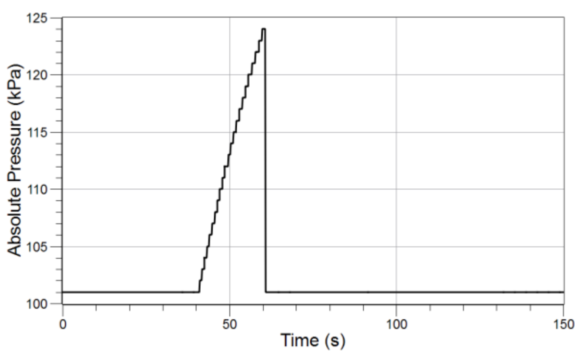
The rate of a chemical reaction is determined by measuring the time it takes for a measured increase in the formation of one or more products.

15. How can the formation of products be measured?

In this chemical reaction, a gas is produced. The formation of the gas can be followed by measuring the change in pressure of a closed vessel as the reaction occurs or by measuring the volume of gas produced using the water displacement method. In this lab, an absolute pressure sensor is used to measure the pressure increase as gaseous hydrogen is produced in a closed test tube.

16. How can the rate of reaction be determined from a graph of Absolute Pressure (kPa) versus Time (s)?

Rate of Reaction



The rate of reaction is the change in the amount of product formed compared to the amount of time it took to form the product. By graphing the formation of product measured as the pressure of the system (on the y-axis) against the amount of time elapsed (on the x-axis), the resulting slope of the best fit-line (change in y divided by change in x) gives a value proportional to the rate of reaction (amount of product divided by the time it took to form the product).

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids creates heat. Be extra careful when handling freshly prepared solutions and glassware because they might be very hot.
- ♦ Always add acids to water, not the other way around, because the solutions may boil vigorously.
- ♦ Handle concentrated acids in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Prepare 500 mL of 4.0 M hydrochloric acid from either concentrated (12 M) or dilute (6 M) HCl. Half of this solution will be used to make the 2.0 M HCl solution and the remaining 250 mL of solution is enough for 50 lab groups.

Starting with concentrated (12 M) HCl:

- **a.** Add approximately 300 mL of distilled water to a 1000-mL beaker with a stir bar.
- **b.** Slowly add 167 mL of 12 M HCl to the beaker with continuous stirring.
- **c.** Allow the solution to cool, then carefully pour it into a 500-mL volumetric flask and dilute to the mark with distilled water.
- **d.** Cap and invert three times carefully to ensure complete mixing.

Starting with dilute (6 M) HCl:

- **a.** Add approximately 100 mL of distilled water to a 1000-mL beaker with a stir bar.
- **b.** Slowly add 333 mL of 6 M HCl to the beaker with continuous stirring water and dilute to the mark with distilled water.
- **c.** Allow the solution to cool, then carefully pour it into a 500-mL volumetric flask and dilute to the mark with distilled water.
- **d.** Cap and invert three times carefully to ensure complete mixing.

2. Prepare 500 mL of 2.0 M hydrochloric acid from the recently prepared 4.0 M HCl. Half of this solution will be used to make the 1.0 M HCl solution and the remaining 250 mL of solution is enough for 50 lab groups.

- **a.** Add approximately 200 mL of distilled water to a 500-mL volumetric flask.
- **b.** Add 250.0 mL of 4.0 M HCl to the volumetric flask.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert three times to ensure complete mixing.

3. Prepare 500 mL of 1.0 M hydrochloric acid from the recently prepared 2.0 M HCl.
Ten milliliters of this solution will be used to make the 0.1 M HCl solution and the remaining 490 mL is enough for 24 lab groups.

- **a.** Add approximately 200 mL of distilled water to a 500-mL volumetric flask.
- **b.** Add 250.0 mL of 2.0 M HCl to the volumetric flask.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert three times to ensure complete mixing.

4. Prepare 100 mL of 0.1 M hydrochloric acid from the recently prepared 1.0 M HCl. This is enough for 20 lab groups.

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 10.0 mL of 1.0 M HCl to the volumetric flask.
- **c.** Dilute to the mark with distilled water.

d. Cap and invert three times to ensure complete mixing.

5. Prepare community water baths in the laboratory by immersing test tube racks in warm (not to exceed 50 °C) and cold (ice cooled) water. Students might leave their labeled test tubes with HCl in the appropriate bath for several minutes until the desired temperature is obtained.

Teacher Tip: Laboratory time is greatly reduced if the magnesium ribbon is precut into strips of approximately 1- to 2-cm. The lengths need not be exactly the same, but they should be similar to enable comparison of rates between experiments. Before cutting the magnesium strip, remove any oxidation on the ribbon surface using steel wool or fine sand paper.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Do not hold the stopper in the tube beyond a pressure of 125 kPa (1.2 atm).
- ♦ Do not point the test tube towards yourself or anyone else.
- ♦ Hydrochloric acid is a corrosive irritant. Avoid contact with the skin and eyes.
- ♦ Be sure that all acids and bases are neutralized before being disposed of down the drain.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	4	3	5
Begin recording data (before the reactants are mixed).	Add hydrochloric acid solution to a test tube.	Carefully release the stopper when the pressure reaches 125 kPa (1.2 atm). Stop recording data.	Add magnesium metal to test tube containing HCl and quickly seal the test tube using the stopper.	Analyze the results to determine the various factors affecting reaction rates.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

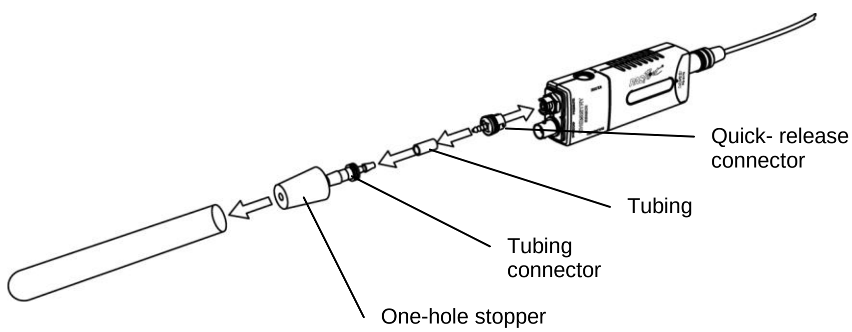
1. ☐ Complete the following steps to attach the one-hole stopper to the absolute pressure sensor:

a. Insert the thicker end of the tubing connector into the hole in the rubber stopper. If this is difficult, add a drop of glycerin.

b. Connect the 1- to 2-cm piece of tubing to the other, thinner end of the tubing connector.

c. Insert the barbed end of a quick-release connector into the open end of the 1- to 2-cm piece of tubing. If this is difficult, add a drop of glycerin.

d. Insert the quick-release connector to the port of the absolute pressure sensor and then turn the connector clockwise until the fitting "clicks" onto the sensor (about one-eighth turn).



2. ☐ Start a new experiment on the data collection system. �(1.2)

3. ☐ Use a sensor extension cable to connect the absolute pressure sensor to the data collection system. �(2.1)

4. ☐ Display Absolute Pressure (kPa) versus Time (s) on a graph. �(7.1.1)

5. ☐ When you mix magnesium ribbon and hydrochloric acid together inside a closed test tube what do you expect to happen to the pressure? Why?

The pressure should increase because the reaction produces hydrogen gas. An increase in the number of gas molecules in the test tube increases the pressure.

Part 1 – Establish a baseline reaction rate

Collect Data

6. ☐ Add 5 mL of 1.0 M HCl into a test tube labeled “baseline”.

7. ☐ Start recording data. �(6.2)

8. ☐ Add three 1-cm pieces (0.05 g) of magnesium ribbon to the test tube and quickly place the stopper attached to the absolute pressure sensor in the test tube.

CAUTION: Apply gentle pressure to keep the stopper in, but do *not* hold the stopper in the tube beyond a pressure of 125 kPa.

CAUTION: Do not point the test tube towards yourself or anyone else.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

9. ☐ Carefully remove the stopper when the pressure reaches 125 kPa and then stop recording data. �(6.2)

10. ☐ Why is it important not to let the pressure in the test tube exceed 125 kPa?

If the pressure in the test tube becomes too great, the glass is no longer able to withstand the pressure and might break, causing serious injury.

11. ☐ Rename the data run “baseline”. �(8.2)

12. ☐ Why is it important to clearly label each run?

When the experiment is complete, there will be at least seven different runs of data that are easily confused. To avoid confusion, it is good scientific practice to clearly label data sets.

13. ☐ Is the magnesium dissolving in or reacting with the HCl? Explain your reasoning.

The magnesium is reacting with the HCl. A reaction has occurred because a new product, a gas (seen as bubbles), was formed. If the magnesium had simply dissolved, a gas would not have been produced.

14. ☐ Dispose of the contents of the test tube according to the teacher’s instructions and then rinse the test tube with water so it can be re-used in the next part of the lab.

Analyze Data

15. ☐ Find the slope (reaction rate) of the initial linear portion of the data run by following the steps below.

a. Apply a linear fit to the first 10 to 20 seconds of the data run after the magnesium was added. �(9.5)

b. Determine the slope of the linear fit line. �(9.6)

c. Record the slope below·

Baseline slope = 0.52 kPa/s

16. ☐ What does the slope of this line represent?

The slope is proportional to the rate of the reaction. The faster the hydrogen is produced, the faster the pressure increases in the test tube. The faster the pressure increases, the steeper the slope of the linear best fit line will be.

Part 2 – The effect of temperature on the rate of reaction

Set Up

17. ☐ Add 5 mL of 1.0 M HCl into each of two different test tubes (one labeled "hot" and the other labeled "cold").

18. ☐ Place the test tube labeled "hot" into a warm water bath (40 to 50 °C) and the test tube labeled "cold" into an ice-water bath (0 to 5 °C). Leave the test tubes in the water baths for 5 to 10 minutes to allow the HCl solutions to attain the appropriate temperatures.

19. ☐ Predict the effect of temperature on the rate of reaction. Explain your prediction using your knowledge of what is happening at the molecular level.

The rate of reaction increases when the reaction occurs at higher temperatures. The reaction rate increases because the molecules are moving faster and are colliding more frequently and with more energy.

20. ☐ What is the independent variable in this part of the experiment?

The independent variable is temperature.

Collect Data

21. ☐ Complete the following steps to measure the change in pressure as magnesium reacts with the "hot" HCl.

a. Remove the test tube from the warm water bath.

b. Start recording data. �(6.2)

c. Add three 1-cm pieces (0.05 g) of magnesium ribbon to the test tube and quickly place the stopper attached to the absolute pressure sensor in the test tube.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

d. Stop recording data when the pressure reaches 125 kPa. �(6.2)

e. Rename this data run "hot". �(8.2)

22. ☐ Complete the following steps to measure the change in pressure as magnesium reacts with the "cold" HCl.

a. Remove the test tube from the cold water bath.

b. Start recording data. �(6.2)

c. Add three 1-cm pieces (0.05 g) of magnesium ribbon to the test tube and quickly place the stopper attached to the absolute pressure sensor in the test tube.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

d. Stop recording data when the pressure reaches 125 kPa. �(6.2)

e. Rename this data run "cold". �(8.2)

23. ☐ Why are you testing "hot" and "cold" HCl but not HCl at room temperature?

The same reaction at room temperature was already performed in the previous section as the baseline to which the "hot" and "cold" reactions can be compared.

24. ☐ Dispose of the contents of the test tubes according to the teacher's instructions and then rinse the test tubes with water so they can be re-used in the next part of the lab.

Analyze Data

25. ☐ Find the slope (reaction rate) of the initial linear portion of the data run for both the "hot" and "cold" HCl reacting with magnesium.

a. Display the run of data you want to analyze. �(7.1.7)

b. Apply a linear fit to the first 10 to 20 seconds of the data run after the magnesium was added. �(9.5)

c. Determine the slope of the linear fit line. �(9.6)

d. Record the slopes in Table 1 below.

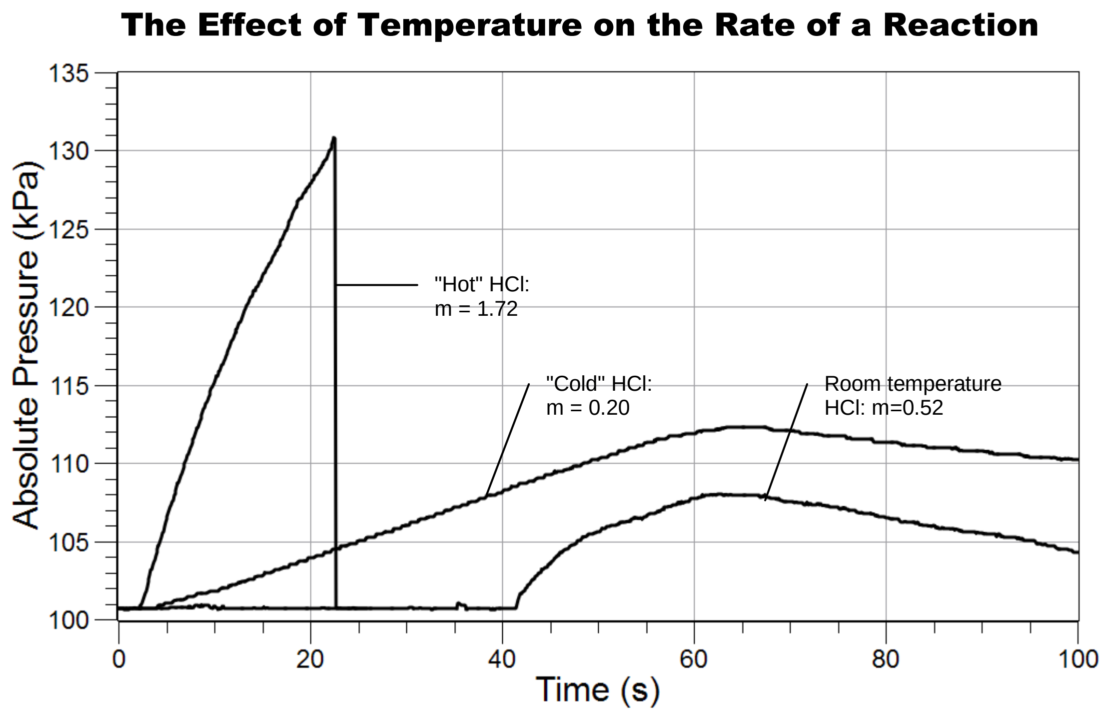
Table 1: Slopes of data runs for Mg reacting with HCl at different temperatures

Reaction Conditions	Slope (kPa/s)
Room temperature 1.0 M HCl + Mg ribbon (baseline reaction from Part 1 above)	0.52
"Hot" 1.0 M HCl + Mg ribbon	1.72
"Cold" 1.0 M HCl + Mg ribbon	0.20

26. ☐ Create a graph with all three runs of data displayed on your data collection system. �(7.1.7)

Note: Not all data collection systems will display all three runs of data on one set of axes.

27. ☐ Sketch or print a copy of the Absolute Pressure (kPa) versus Time (s) graph displaying the data collected when 1.0 M HCl and magnesium ribbon were reacted at three different temperatures (hot, cold, and room temperature). Label each data run as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Part 3 – The effect of surface area on the rate of reaction

Set Up

28. ☐ Which has a greater surface area, powdered magnesium or 1-cm pieces of magnesium ribbon?

Powdered magnesium has much more surface area than 1-cm pieces of magnesium ribbon.

29. ☐ Predict the effect of surface area on the rate of reaction. Explain your prediction using what you know is happening at the molecular level.

The powdered magnesium reacts faster than the magnesium ribbon because it has a greater surface area. The magnesium atoms and HCl atoms must collide for a reaction to occur. These two particles can only collide at the surface of the magnesium. Therefore, the more magnesium there is that can come in contact with HCl molecules, the faster the reaction occurs.

30. ☐ Add 5.0 mL of 1.0 M HCl into a test tube labeled “Mg powder”.

Collect Data

31. ☐ Start recording data. �(6.2)

32. ☐ Add ~0.05 g of magnesium powder to the HCl and quickly place the stopper attached to the absolute pressure senor in the test tube.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

33. ☐ Stop recording data when the pressure reaches 125 kPa. �(6.2)

34. ☐ Rename this data run "Mg powder". �(8.2)

35. ☐ Dispose of the contents of the test tube according to the teacher's instructions and then rinse the test tube with water so it can be re-used in the next part of the lab.

Data Analysis

36. ☐ Find the slope (reaction rate) of the initial linear portion of the data run for the reaction between 1.0 M HCl and powdered magnesium.

a. Display the run of data you want to analyze. �(7.1.7)

b. Apply a linear fit to the first 10 to 20 seconds of the data run after the magnesium was added. �(9.5)

c. Determine the slope of the linear fit line. �(9.6)

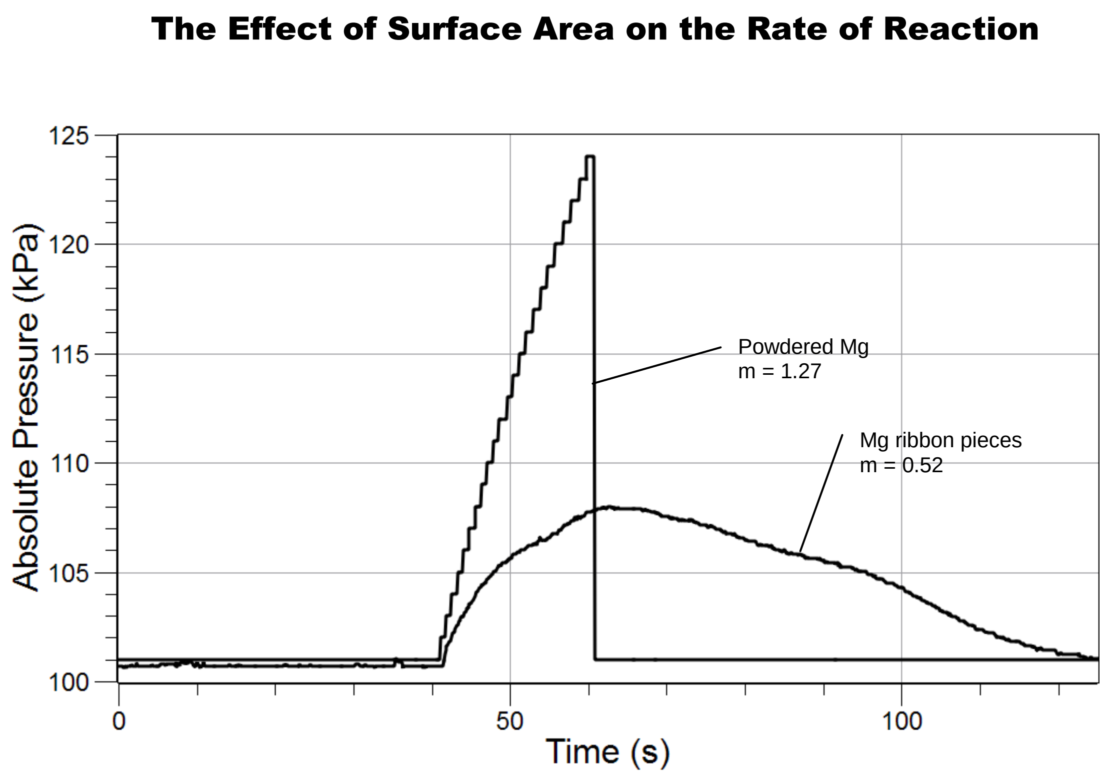
d. Record the slope in Table 2 below.

Table 2: Slopes of data runs for Mg ribbon and Mg powder reacting with HCl

Reaction Conditions	**Slope (kPa/s)**
1.0 M HCl + Mg ribbon (baseline reaction from Part 1 above)	0.52
1.0 M HCl + Mg powder	1.27

37. ☐ Create a graph with both runs of data (Mg ribbon and Mg powder) displayed on your data collection system. �(7.1.7)

38. ☐ Sketch or print a copy of the Absolute Pressure (kPa) versus Time (s) graph displaying the reaction between 1.0 M HCl and magnesium powder and the reaction between 1.0 M HCl and magnesium ribbon. Label each data run as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Part 4 – The effect of concentration on the rate of reaction

Set Up

39. ☐ How are 1 M and 2 M HCl solutions different? Use the term "concentration" in your explanation.

A 1 M HCl solution has 1 mole of HCl molecules per liter of solution and a 2 M HCl solution has 2 moles of HCl molecules per liter of solution; the 2 M solution is twice the concentration of the 1 M solution. The 2 M HCl solution contains more molecules of HCl, therefore it is more concentrated than the 1 M HCl solution.

40. ☐ Predict the effect of concentration on the rate of reaction. Explain your prediction using what you know is happening at the molecular level during a chemical reaction.

The more concentrated the reactants are, the faster the reaction occurs. For a reaction to occur, the reactant particles must collide with enough energy. The greater the number of particles there are, the greater the probability that the particles will collide and the faster the reaction occurs.

41. ☐ Measure 5 mL of 0.1 M, 2.0 M, and 4.0 M HCl into three different test tubes. Label each test tube with the corresponding concentration.

Collect Data

42. ☐ Complete the following steps to measure the change in pressure as magnesium reacts with the 0.1 M HCl.

a. Start recording data. �(6.2)

b. Add three 1-cm pieces (0.05 g) of magnesium ribbon into the test tube and quickly place the stopper attached to the absolute pressure sensor in the test tube.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

c. Stop recording data when the pressure reaches 125 kPa. �(6.2)

d. Rename this data run “0.1 M HCl”. �(8.2)

43. ☐ Complete the following steps to measure the change in pressure as magnesium reacts with the 2.0 M HCl.

a. Start recording data. �(6.2)

b. Add three 1-cm pieces (0.05 g) of magnesium ribbon into the test tube and quickly place the stopper attached to the absolute pressure sensor in the test tube.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

c. Stop recording data when the pressure reaches 125 kPa. �(6.2)

d. Rename this data run “2.0 M HCl”. �(8.2)

44. ☐ Complete the following steps to measure the change in pressure as magnesium reacts with the 4.0 M HCl.

a. Start recording data. �(6.2)

b. Add three 1-cm pieces (0.05 g) of magnesium ribbon into the test tube and quickly place the stopper attached to the absolute pressure sensor in the test tube.

Note: You may have to adjust the scale of the graph to observe any changes taking place. �(7.1.2)

c. Stop recording data when the pressure reaches 125 kPa. �(6.2)

d. Rename this data run “4.0 M HCl”. �(8.2)

45. ☐ Save the data file and clean up according to the teacher’s instructions. �(11.1)

Analyze Data

46. ☐ Find the slope of the initial linear portion of the data run for the reactions between magnesium ribbon and 0.1 M, 2.0 M, and 4.0 M HCl.

a. Display the run of data you want to analyze. �(7.1.7)

b. Apply a linear fit to the first 10 to 20 seconds of the data run after the magnesium was added. �(9.5)

c. Determine the slope of the linear fit line. �(9.6)

d. Record the slopes in Table 3 below.

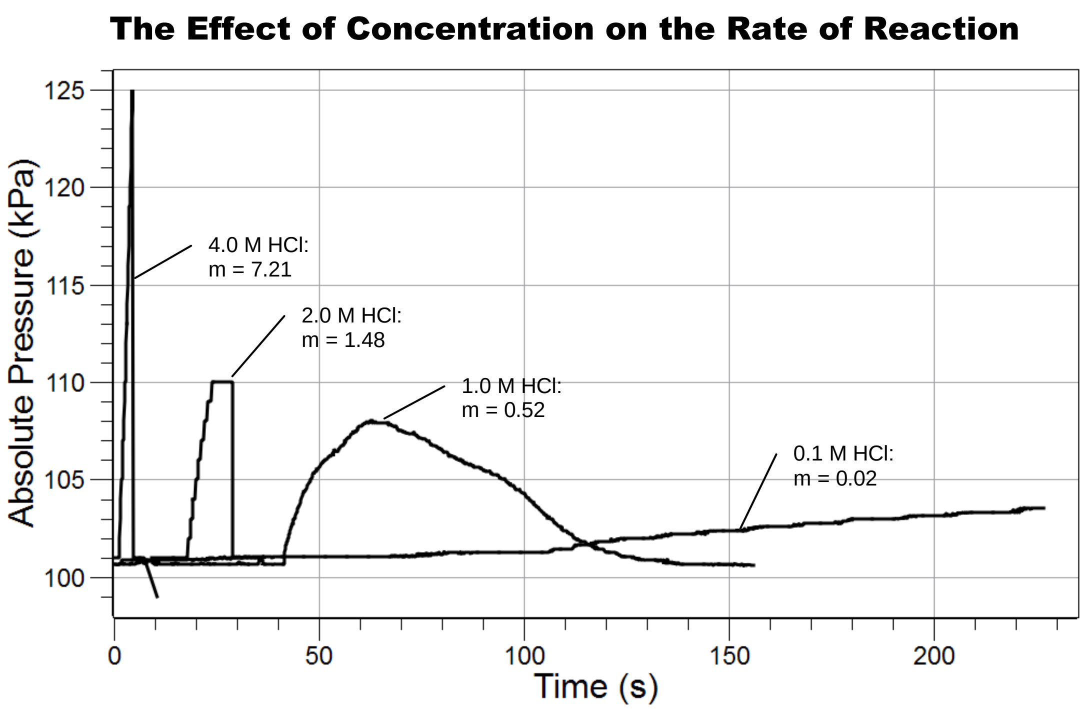
Table 3: Slopes of data runs for Mg reacting with different concentrations of HCl

Reaction Conditions	Slope (kPa/s)
0.1 M HCl + Mg ribbon	0.02
1.0 M HCl + Mg ribbon (baseline reaction from Part 1 above)	0.52
2.0 M HCl + Mg ribbon	1.48
4.0 M HCl + Mg ribbon	7.21

47. ☐ Create a graph with all four runs of data displayed on your data collection system. �(7.1.7)

Note: Not all data collection systems will display all four runs of data on one set of axes.

48. ☐ Sketch or print a copy of the Absolute Pressure (kPa) versus Time (s) graph displaying the data collected when magnesium ribbon and 0.1 M, 2.0 M, and 4.0 M HCl reacted. Label the data runs as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Data Analysis

1. ☐ Explain any differences between the rate of the baseline reaction and the rate of the reaction with the hydrochloric acid solutions at different temperatures.

As the temperature increased, the rates of the reactions (slopes of pressure versus time) increased. The cold HCl reacted the slowest with the least slope, and the warm HCl reacted the fastest with the greatest slope. This is because at higher temperatures, the hydrogen ions from the HCl move faster and collide more frequently with the magnesium; a greater collision frequency results in a greater rate of reaction.

2. ☐ Explain the differences between the rate of the baseline reaction and the rate of the reaction with the powdered magnesium.

The data should indicate that the powdered magnesium reacted faster than the magnesium ribbon. This is because the powdered magnesium is very finely divided, and, therefore, has a much greater surface area than the magnesium ribbon. Greater surface area available for reaction means more hydrogen ions can collide with magnesium in the same period of time compared to the magnesium ribbon. Again, higher collision frequency yields a greater rate.

3. ☐ Explain the differences between the rate of the baseline reaction and the rates of the reactions with different concentrations of HCl.

As the concentration increased, the rates of the reactions increased. The 4.0 M of HCl reacted the fastest (had the greatest slope), followed by 2.0 M and 1.0 M of HCl; 0.1 M of HCl reacted the slowest (had the least slope). Greater HCl concentration means there are more hydrogen ions in solution; this leads to a greater number of collisions with the magnesium per unit time compared to the more dilute solutions. A higher collision frequency, again, yields a faster rate of reaction.

Analysis Questions

1. Why is the absolute pressure sensor used in this experiment?

One of the products (hydrogen) is a gas, so the relative amount of hydrogen gas present is easily determined by measuring the pressure inside the test tube. None of the other reactants or products are as easily measured in real time.

2. Why is it important to establish a baseline reaction rate?

A baseline reaction rate is necessary for comparison purposes to determine which reactions react faster and which react slower. All scales, of whatever type, need a baseline value or “zero point” present to make comparisons. For example, measures of altitude have sea level as their “zero point.” To determine differences, it does not matter where the zero point is defined, only that there is one present and it is well defined.

3. Explain why the slope of the pressure versus time plot can be used to describe the rate of the reaction.

The greater the rate of the reaction, the faster hydrogen gas is produced. The faster hydrogen gas is produced, the faster the pressure increases in the test tube. The faster the pressure increases, the steeper the slope (change in y divided by the change in x) of the linear best–fit line for pressure (y-axis) versus time (x-axis) will be.

4. What combination of treatments available in the laboratory would create the fastest reaction rate between magnesium and hydrochloric acid?

A combination of powdered magnesium reacting with warm 4.0 M of HCl would lead to the fasted reaction rate possible using the supplied resources in the laboratory.

Synthesis Questions

Use available resources to help you answer the following questions.

1. List other ways in which the amount of reactants or products present in reactions can be determined realistically and in real time.

Depending on the type of reaction that is performed, measuring pH, absorbance, temperature, or conductivity could all be used to determine a change in the amount of products or reactants as a reaction occurs.

2. Why is it important to study rates of reactions?

Knowing the rate of a reaction allows for the determination of how long it will take for a given amount of product to form. This is important in the synthesis of chemicals such as medicine as well as in giving consumers an idea of how long a product, such as food, will last, or predicting dangerous combinations such as warm 4.0 M of HCl with powdered magnesium.

3. A catalyst is a substance that allows chemical reactions to occur at a lower energy than is normally expected. How do catalysts affect reaction rates? Why?

Catalysts cause reaction rates to increase. By lowering the energy needed for the products to form, a larger proportion of the collisions have sufficient energy, and, therefore, the reaction occurs faster.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. In general, the ____________ the concentration of reactants in a reaction, the ________ the rate of the reaction will be.

- **A.** Greater; slower
- **B. Greater; faster**
- **C.** Lower; faster
- **D.** No relationship exists

2. In general, the ____________ the surface area available for reaction is, the ________ the rate of the reaction will be.

- **A. Greater; faster**
- **B.** Greater; slower
- **C.** No relationship exists
- **D.** Smaller; faster

3. In general, the ____________ the temperature of a reaction is, the ________ the rate of the reaction will be.

- **A.** No relationship exists
- **B.** Lower; faster
- **C.** Greater; slower
- **D. Greater; faster**

4. At the molecular level, what needs to happen for reactant particles to form products?

- **A.** The reactant particles need to be in an excited state
- **B.** The reactant particles must have a molar mass greater than 10 g/mol
- **C.** The reactant particles need to be in the gaseous state
- **D. The reactant particles need to collide with sufficient energy**

5. What form of iron will form rust the fastest?

- **A.** A galvanized nail
- **B.** A solid block of iron
- **C. Iron filings**
- **D.** All forms of iron form rust at the same rate

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The part of chemistry that deals with the rates of chemical reactions is called **kinetics**. Explosions have very **large** rates of reactions; while the formation of rust has a very **small** reaction rate. **Collision theory** describes how chemicals react at the molecular level. Collision theory states that the frequency with which reacting particles run into each other determines the rate of the reaction. The greater the number of collisions, the **faster** the rate of the reaction. The number of collisions can be increased by **warming** the reactants, **increasing** the concentration of the reactants, or by increasing the particles' **surface area**. In contrast, **decreasing** any or all of these factors results in a lower reaction rate.

Extended Inquiry Suggestions

Investigate factors that slow the rate at which a cut apple browns.

Investigate factors that slow the rate of rust formation.

Investigate how the rates of reactions for extremely fast reactions (explosives) and extremely slow reactions (decomposition) are determined.

Extrapolate the data from initial rates to calculate the amount of time required to consume 1.0 g of magnesium ribbon in various concentrations of room temperature hydrochloric acid. Test the results and compare the results to the predictions.

Investigate the affects of temperature on the reaction of an effervescent tablet and water.

Construct a curve of reaction rates at various concentrations for the Landolt (iodine clock reaction). Use this to further discuss zero, first, and second order reactions.

23. Ideal Gas Law

Objectives

Determine the number of moles of carbon dioxide gas generated during a reaction between hydrochloric acid and sodium bicarbonate. Through this investigation, students:

- Apply the ideal gas law ($PV = nRT$) to experimentally determine the number of moles of carbon dioxide gas generated
- Use the balanced chemical equation and stoichiometry to calculate the theoretical number of moles of carbon dioxide gas generated

Procedural Overview

Students conduct the following procedures:

- Measure temperature and pressure throughout the reaction between hydrochloric acid and sodium bicarbonate
- Compare the experimentally determined number of moles of carbon dioxide generated to the number of moles that are theoretically possible and explain the discrepancies

Time Requirement

Preparation time	20 minutes
Pre-lab discussion and activity	30 minutes
Lab activity	30 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Absolute pressure sensor
- Stainless steel temperature sensor
- Blue plastic tubing for the temperature sensor[1]
- Sensor extension cable
- Balance, centigram
- Graduated cylinder or volumetric pipet, 10-mL
- Graduated cylinder, 1000-mL
- Test tube, 15-mm x 100-mm
- Plastic bottle, 300- to 500-mL
- Two-hole stopper that fits the plastic bottle
- Quick-release connector[2]
- Tubing, 1- to 2-cm[2]
- Tubing connector[2]
- 1.0 M Hydrochloric acid (HCl), 10 mL[3]
- Sodium bicarbonate ($NaHCO_3$), 0.80 g
- Glycerin, 2 drops
- Paper towels

[1]Included with PASCO's stainless steel temperature sensor.

[2]Included with PASCO pressure sensors.

[3]To prepare using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- ♦ Individual gas laws
- ♦ Pressure
- ♦ Chemical Reactions
- ♦ Limiting Reactants
- ♦ Mole calculations
- ♦ Stoichiometric calculations

Related Labs in This Guide

Labs conceptually related to this one include:

- ♦ Percent Oxygen in Air
- ♦ Boyle's Law
- ♦ Gay-Lussac's Law and Absolute Zero
- ♦ Stoichiometry
- ♦ Evidence of a Chemical Reaction

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- ♦ Starting a new experiment on the data collection system �(1.2)
- ♦ Connecting multiple sensors to the data collection system �(2.2)
- ♦ Starting and stopping data recording �(6.2)
- ♦ Adjusting the scale of a graph �(7.1.2)
- ♦ Displaying multiple graphs simultaneously �(7.1.11)
- ♦ Finding the values of a point in a graph �(9.1)
- ♦ Saving your experiment �(11.1)
- ♦ Printing �(11.2)

Background

The ideal gas law combines the four variables that describe a gas–volume (*V*), absolute pressure (*P*), temperature (*T*) and number of moles (*n*)–into one equation.

$$PV = nRT$$

The universal gas constant, *R*, is the same for all ideal gases. The value and units of *R*, however, depend on the specific units for pressure used in the equation. The most common forms are: 0.0821 (atm·L)/(mol·K), when pressure is in units of atmospheres, and 8.31 (kPa·L)/(mol·K), when pressure is in units of kilopascals. As with the other gas laws, the temperature values must be measured in Kelvin.

The universal gas constant can be calculated using standard conditions. Standard temperature is 273.15 K (0 °C) and standard pressure is 101.325 kPa (1 atm). The volume of one mole of gas at standard temperature and pressure (STP) is 22.4 L.

$$\frac{PV}{nT} = R$$

$$\frac{(1\,\text{atm})(22.414\ \text{L})}{(1\,\text{mol})(273.15\ \text{K})} = \frac{0.082057\ \text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}} \approx 0.0821(\text{atm}\cdot\text{L})/(\text{mol}\cdot\text{K})$$

$$\frac{(101.325\ \text{kPa})(22.414\text{L})}{(1\,\text{mol})(273.15\ \text{K})} = \frac{8.3145\ \text{kPa}\cdot\text{L}}{\text{mol}\cdot\text{K}} \approx 8.31(\text{kPa}\cdot\text{L})/(\text{mol}\cdot\text{K})$$

In this investigation, the ideal gas law will be used to determine the number of moles of carbon dioxide formed when hydrochloric acid (HCl) reacts with sodium bicarbonate ($NaHCO_3$). The students will react HCl and $NaHCO_3$ in a plastic bottle that is closed with a stopper.

$$HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + NaCl(aq) + CO_2(g)$$

The temperature and absolute pressure of the carbon dioxide (CO_2) produced in the reaction will be measured using sensors connected to the data collection system. The volume of the plastic bottle will be determined by measuring the amount of water it can contain.

In the analysis section, the students will calculate the theoretical number of moles of carbon dioxide gas released based on the amount of limiting reactant ($NaHCO_3$) in the reaction. They will then compare their experimental results to their theoretical calculations, reporting their calculated percent error.

Pre-Lab Discussion and Activity

Magnesium Metal in Acid

Engage the students by demonstrating the reaction between magnesium metal and acid. Add a 2-cm strip of magnesium ribbon to a small beaker containing dilute hydrochloric acid. Have the students suggest ways for determining the amount of hydrogen gas released. In the discussion, review evidence for a reaction, limiting reactants, excess reactants, the behavior of gases, and introduce the ideal gas law.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

1. Is the mixing of magnesium and hydrochloric acid a physical change or a chemical reaction? How do you know?

Magnesium and hydrochloric acid react to form magnesium chloride and hydrogen gas. The students should recognize this as a chemical reaction because a gas is released.

2. How can the amount of hydrogen gas generated be increased? Why will that work?

Add more magnesium. This will work because magnesium is the limiting reactant; when the magnesium is completely consumed, the reaction stops.

3. What is the difference between the limiting reactant and the excess reactant?

The limiting reactant is the reactant that gets used up, when it is used up the reaction stops. The excess reactant is still present when the reaction stops.

4. How can the amount (number of moles) of hydrogen gas released be determined in the laboratory?

Ask as a rhetorical question. Lead the students through the next series of questions to find the answer. The answer is that volume, pressure, and temperature can be measured and then used to calculate the number of moles of gas formed using the ideal gas law.

5. What properties of gases make them difficult to measure?

Gases take the shape and volume of their containers, and if they are not in a container, they disperse. This makes them difficult to measure.

6. What variables are used to describe the behavior of a gas?

Volume (V), absolute pressure (P), temperature (T) and number of moles (n) are the four variables used to describe gases.

7. How can the variables of a gas be measured in the lab?

Volume, V, is the volume of the container the gas occupies.

Pressure, P, can be measured using a pressure sensor.

Temperature, T, can be measured using a temperature sensor or thermometer.

Moles, n, is difficult to measure.

8. How do we measure moles of a solid? A liquid?

The number of moles of pure solids and liquids can be calculated from the mass of the sample using their formula weight.

$$\text{mass of sample} \times \frac{1\,\text{mole}}{\text{formula weight}} = \text{moles of sample}$$

$$30\text{ g C} \times \frac{1\text{ mole C}}{12.01\text{ g}} = 2.5\text{ moles of C}$$

9. How can the number of moles of hydrogen gas released be determined?

First, the hydrogen gas needs to be trapped inside a container. The volume, pressure, and temperature of the gas can then be measured. The number of moles can be calculated by substituting these measured values into an equation that relates all of these variables to the number of moles of gas present. This equation is the ideal gas law.

Ideal Gas Law

Introduce the students to the ideal gas law and explain the ideal gas constant, *R*. Have the students discuss how pressure, temperature, and volume can be measured.

10. What is the difference between an ideal gas and a real gas?

The particles in an ideal gas have no volume, allowing them to be compressed infinitely; they never touch one another. The molecules of an ideal gas have no attractions for each other; they never condense into a liquid. The collisions between molecules in an ideal gas are perfectly elastic; the system does not lose energy. Real gases, however, obey the constraints of reality; the particles do take up space, have intermolecular attractions, and lose energy in collisions.

11. Under what conditions do gases behave ideally?

Gases behave ideally at high temperatures and low pressures.

12. What is the ideal gas law? What do the symbols stand for?

The ideal gas law relates the absolute pressure of the system (P), the volume of the container (V), the number of moles of gas present (n), the temperature in Kelvin (T), and a constant (the universal gas constant, R) in a single equation: $PV = nRT$.

13. How can the ideal gas law be used to determine the number of moles of gas generated during a chemical reaction?

Confine the reaction in a closed container and measure the temperature, pressure, and volume of the gas. Substitute these values, as well as the universal gas constant, into the ideal gas law and solve for moles.

14. Dalton's law of partial pressure states that the pressure inside a container is due to all the different gases present pushing on the walls of the container. Knowing air is a mixture of gases that present in the bottle before the reaction starts, how can the pressure of just the product be determined?

The pressure before the reaction begins is from the gases in the air that occupy the container. Any additional pressure created during the experiment is from the gas that was produced by the chemical reaction. Subtracting the initial pressure from the final pressure will leave the pressure produced by only the product of the reaction.

15. The temperature may change during the chemical reaction. Which temperature reading should be used in the ideal gas law equation?

The highest temperature that occurs when the pressure is the greatest should be used in the ideal gas law. This is the temperature of the system when the reaction was fully complete and the greatest number of moles of product was inside the container.

16. How can the volume of the container be measured? Can you just read the label?

Simply reading the markings on the container may not be accurate. The best way to determine the volume is to completely fill the container with water and then measure the amount by transferring the water to a graduated cylinder.

17. How can the value for the universal gas constant, R, be calculated?

The universal gas constant can be calculated by substituting the values for *P*, *V*, *n*, and *T* at standard conditions (one mole of gas occupies 22.414 L of volume at 1 atm (101.325 kPa) pressure and 273.15 K) into the ideal gas law equation.

18. Are the units for the universal gas constant important?

Yes. Be sure to include the units in all parts of your calculation. If the units do not cancel, then you have made an error. The value for *R* depends on the units of pressure being measured either in atmospheres or kilopascals. As with all gas law calculations, the temperature must always be in Kelvin. Traditionally, volume is in liters.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid or base solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids and bases create heat; be extra careful when handling freshly prepared solutions and glassware as they may be very hot.
- ♦ Always add acids and bases to water, not the other way around, as the solutions may boil vigorously.
- ♦ Handle concentrated acids and bases in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Prepare 100 mL of 1.0 M hydrochloric acid from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 10 lab groups.

Starting with concentrated (12 M) HCl

a. Add approximately 50 mL of distilled water to a 100 mL beaker with a stir bar.

b. Slowly add 8.3 mL of 12 M HCl to the beaker with continuous stirring.

c. Allow the solution to cool, then carefully pour into a 100-mL volumetric flask and dilute to the mark with distilled water.

d. Cap and invert three times to ensure complete mixing.

Starting with dilute (6 M) HCl

a. Add approximately 50 mL of distilled water to a 100 mL volumetric flask.

b. Add 16.7 mL of 6 M HCl to the water and dilute to the mark with distilled water.

c. Cap and invert three times to ensure complete mixing.

Teacher Tip: To save time in the lab, you can prepare stoppers by inserting one sensor into each hole. Alternatively, you can have the first period class prepare the stoppers, but leave them assembled for additional classes.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Hydrochloric acid is a strong acid. Avoid contact with the skin and eyes.
- ♦ Be sure that all acids and bases are neutralized before disposal down the drain.
- ♦ Be aware that the gas being generated causes an increase in pressure which may expel the stopper from the bottle. Because of this, eye protection should be worn during this experiment to prevent injury due to flying objects as well as splashed chemical. Reducing the amount of the limiting reactant ($NaHCO_3$) will reduce the resulting pressure of the products. It is better to error on the side of too little of the reactants than too much.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

4	2	5	1	3
Tilt the bottle slightly to allow the HCl to spill out of the test tube and mix with the sodium bicarbonate.	Put the sodium bicarbonate and a test tube containing 10 mL of 1.0 M HCl into a plastic bottle and insert the stopper.	Analyze the data collected to determine the initial and final pressure inside the bottle as well as the final temperature.	Determine the mass of the sodium bicarbonate.	Begin recording temperature and pressure data.

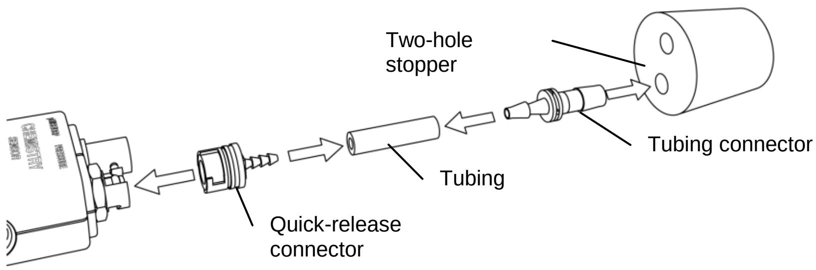
Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Connect the two-hole stopper to the absolute pressure sensor.



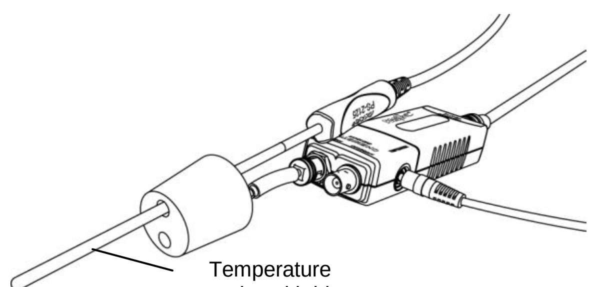
a. Insert the thicker end of the tubing connector into one of the holes in the stopper. If this is difficult, add a drop of glycerin.

b. Connect the 1- to 2-cm piece of tubing to the other, thinner end of the tubing connector.

c. Insert the barbed end of the quick-release connector into the open end of the 1- to 2-cm piece of tubing. If this is difficult, add a drop of glycerin.

d. Insert the quick-release connector into the port of the absolute pressure sensor and then turn the connector clockwise until the fitting clicks (about one-eighth turn).

2. ☐ Put a drop of glycerin into the open hole in the stopper and then insert the stainless steel temperature sensor, covered in blue plastic tubing, through the hole.



3. ☐ Remove any excess glycerin from the temperature sensor with a paper towel.

4. ☐ What are two reasons that glycerin is used in the setup above?

The glycerin acts as a lubricant and also creates and airtight seal so that the gas that will be generated will not escape.

5. ☐ Start a new experiment on the data collection system. �(1.2)

6. ☐ Connect the absolute pressure sensor (using an extension cable) and the stainless steel temperature sensor to the data collection system. �(2.2)

7. ☐ Display two graphs simultaneously. On one graph display Pressure (kPa) versus Time (s) and in the second graph, display Temperature (K) versus Time (s). �(7.1.11)

8. ☐ Weigh between 0.50 and 0.75 grams of sodium bicarbonate ($NaHCO_3$) and place it into an empty plastic bottle. Record the exact mass added below.

Mass of $NaHCO_3$ (g): 0.60 g

Teacher Tip: The amount of sodium bicarbonate must not exceed 0.80 g, otherwise the HCl will become the limiting reactant.

9. ☐ In this experiment, the sodium bicarbonate is the limiting reactant. What is a limiting reactant?

It is the reactant that is completely consumed in a reaction which in turn causes the reaction to cease.

10. ☐ Fill a small test tube with 10.0 mL of 1.0 M hydrochloric acid (HCl) and place it in a test tube rack until needed later.

11. ☐ In this experiment, the hydrochloric acid is the excess reactant. What is an excess reactant?

It is the reactant that is left over and is not completely consumed in a reaction. It ensures that the limiting reactant is completely consumed.

12. ☐ Tilt the plastic bottle and carefully slide the small test tube down inside so the HCl stays in the test tube and does not mix with the sodium bicarbonate.

13. ☐ Taking care not to spill the HCl from the test tube, place the stopper in the plastic bottle and *make sure it is secure*.

14. ☐ What will happen if the stopper is not secure in the plastic bottle?

The gas that is generated during the reaction will be able to escape from the bottle, which will cause incorrect pressure data.

Collect Data

15. ☐ Start recording data. �(6.2)

Note: During data collection you may need to adjust the scale of the graph to see the changes taking place. �(7.1.2)

16. ☐ While holding the stopper in place, slowly rotate the bottle until the HCl spills from the test tube and mixes with the $NaHCO_3$.

17. ☐ What do you observe? Do these observations suggest that a chemical reaction is taking place? Explain.

Bubbles are forming, the temperature is increasing, and the pressure is increasing. The formation of a gas indicates that a chemical reaction is occurring, because it is a new substance that has formed from the two reactants.

18. ☐ Is the chemical reaction exothermic or endothermic? How do you know?

Exothermic. The temperature of the system increased, which means that energy is being released.

19. ☐ When the pressure and temperature have reached a maximum value and bubbles are no longer forming within the plastic bottle, stop recording data. �(6.2)

20. ☐ Dispose of the contents of the plastic bottle according to your teacher's instructions and then rinse the plastic bottle with water.

21. ☐ Determine the volume of your container by filling the bottle completely with water and then pouring it into a 1000-mL graduated cylinder. Record the volume below.

Volume of plastic bottle (g): __________358 mL__________

22. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

1. ☐ Determine the measured quantities in Table 1 below from the graph of Pressure (kPa) versus Time (s) and Temperature (K) versus Time (s) that you collected. �(9.1)

Note: The amount of $NaHCO_3$ and volume data are the values recorded during the experiment in the Procedure section, above.

Table 1: Pressure, temperature, mass, and volume data

Measured Quantity	Value
Initial Pressure (kPa)	101
Final Pressure (kPa)	146
Final Temperature (K)	299.1
Amount of $NaHCO_3$ (g)	0.60
Volume (mL)	358

2. ☐ Find the change in pressure that occurred as a result of the chemical reaction.

Final Pressure – Initial Pressure = Change in Pressure

146 kPa – 101 kPa = 45 kPa

3. ☐ Based on the units for pressure, which value for the universal gas constant will be used in the ideal gas law? Be sure to include the proper units.

Because pressure is in units of kilopascals, the universal gas constant will be 8.31 (kPa·L)/(mol·L).

4. ☐ Calculate the experimental number of moles of carbon dioxide gas generated from the reaction. Use the change in pressure for P, the final temperature in Kelvin for T, the volume of the bottle in liters for V, and the proper value for the universal gas constant, R.

$$PV = nRT$$

$$\frac{PV}{RT} = n$$

$$\frac{(45\text{ kPa})(0.358\text{ L})}{\left(\frac{8.31\text{ L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}}\right)(299.1\text{ K})} = n$$

$$0.0065\text{ mol} = n$$

5. ☐ Write the chemical equation that represents the reaction happening in the bottle.

$HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + NaCl(aq) + CO_2(g)$

6. ☐ Calculate the amount of $CO_2(g)$ expected to have been generated based on the reactants used.

Since the sodium bicarbonate is identified as the limiting reactant, and the balanced equation shows a 1:1 ratio between $NaHCO_3$ and CO_2, then the number of moles of $NaHCO_3$ will be equal to the number of moles of CO_2 produced.

$$(0.60\text{ g NaHCO}_3)\left(\frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3}\right) = 0.0071\text{ mol NaHCO}_3$$

$$(0.0071\text{ mol NaHCO}_3)\left(\frac{1\text{ mol CO}_2}{1\text{ mol NaHCO}_3}\right) = 0.0071\text{ mol CO}_2$$

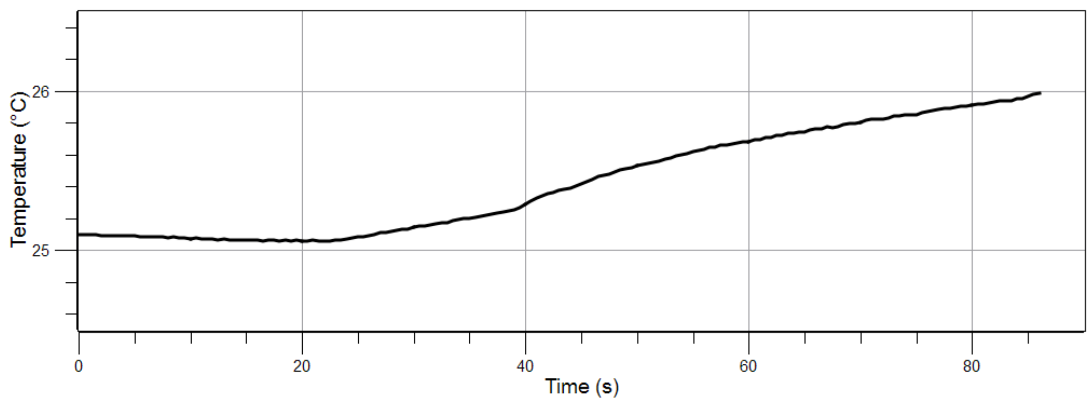
7. ☐ Calculate the percent error.

$$\text{Percent error} = \left|\frac{\text{accepted value} - \text{experimental value}}{\text{accepted value}}\right| \times 100$$

$$\text{Percent error} = \left|\frac{0.0071\text{ mol} - 0.0065\text{ mol}}{0.0071}\right| \times 100 = 8.5\%$$

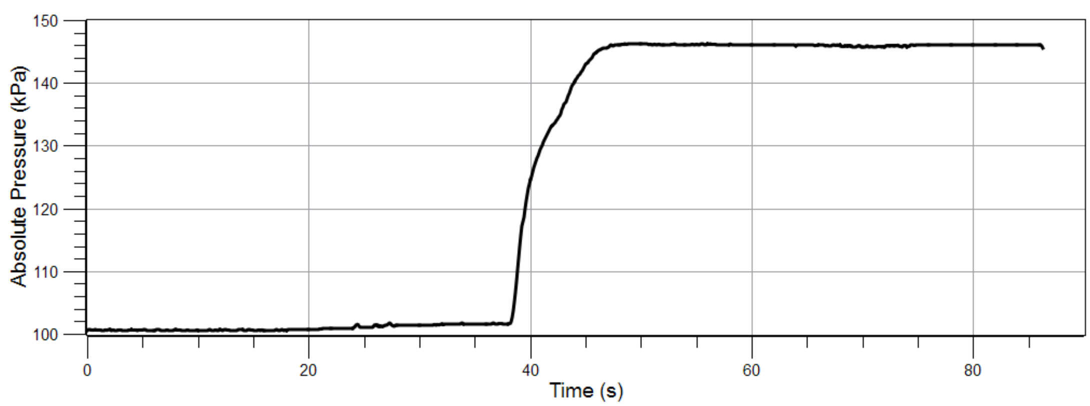
8. ☐ Sketch or print the graphs of Temperature (°C) versus Time (s). Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

Temperature Change as Hydrochloric Acid and Sodium Bicarbonate React



9. ☐ Sketch or print the graphs of Pressure (kPa) versus Time (s). Label the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)

Pressure Change as Hydrochloric Acid and Sodium Bicarbonate React



Analysis Questions

1. What happened to the pressure during the reaction? Why?

The pressure in the container increased as the reaction produced more and more carbon dioxide. As the reaction progressed, the number of moles of gaseous carbon dioxide produced increased resulting in a steady increase of the pressure in the container until the sodium bicarbonate (limiting reactant) was completely consumed, stopping the reaction. More gas particles were present at the end of the reaction than at the beginning, resulting in a greater number of molecular collisions with the walls of the container leading to an increase in observed pressure.

2. Why is the difference in pressure required in this experiment and not the difference in temperature?

At the beginning of the experiment, the bottle contained a mixture of gases (air). As the reaction progressed, these gases remained in the bottle and new molecules of carbon dioxide were added as a result of the reaction between hydrochloric acid and sodium bicarbonate. The increase in pressure was a direct result of the newly produced product. By using the difference in pressure, the ideal gas law is calculating only those variables directly related to the product produced. Temperature, however, is a measure of the average kinetic energy, which applies equally to all of the gas molecules in the container.

3. Would the same value for moles of carbon dioxide been calculated if the temperature substituted into the ideal gas law was in degrees Celsius? Explain.

No. All the gas laws, including the ideal gas law, require that the temperature be in Kelvin.

4. What could have contributed to your percent error?

The contents of the container were not entirely gas. This would change the actual volume available for the gas to occupy. There was some volume occupied by the mixture at the bottom of the bottle and the test tube. The seal made with the stopper may not have been leak-proof causing some of the product to escape reducing the pressure inside the container. Additionally, the conditions of the experiment were not conducive for this gas to behave as an ideal gas; the conditions may have allowed for the molecules to interact where their particle volumes, the attractions between each other, and the energy loss due to collisions were not negligible.

Synthesis Questions

Use available resources to help you answer the following questions.

1. What could be done to make the carbon dioxide gas behave more like an ideal gas?

Warm the carbon dioxide to a higher temperature or allow it to occupy a larger volume (container), thus reducing the overall pressure.

2. How many liters would the theoretical number of moles of carbon dioxide released from the 0.60 grams of sodium bicarbonate occupy at STP?

$$\left(0.60\text{ g NaHCO}_3\right)\left(\frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3}\right) = 0.0071\text{ mol NaHCO}_3$$

1 mole of any gas at standard temperature and pressure (STP: 1 atm, 0 °C) occupies 22.4 L of volume.

$$\left(0.0071\text{ mol CO}_2\right)\left(\frac{22.4\text{ L}}{1\text{ mol CO}_2}\right) = 0.16\text{ L}$$

3. How did the conditions in the plastic bottle compare to STP?

Standard temperature is 273.15 K (0 °C); the temperature in the bottle was warmer (299.1 K).

Standard pressure is 101.315 kPa; although the conditions began near standard pressure, the pressure was much greater at the end of the experiment (146 kPa).

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. At the same temperature and pressure, which sample contains the same number of moles of particles as 1 liter of O_2(g)?

- **A. 1 L Ne(g)**
- B. 2 L N_2(g)
- C. 0.5 L SO_2(g)
- D. 1 L H_2O(l)

2. A student calculated the number of moles of hydrogen gas produced from the reaction between magnesium and hydrochloric acid to be 0.142 moles. If the theoretical amount of hydrogen gas that should have been produced was 0.147 moles, what is the student's percent error for the experiment?

- A. $\frac{0.005}{0.142} \times 100$
- **B.** $\boldsymbol{\frac{0.005}{0.147} \times 100}$
- C. $\frac{0.147}{0.142} \times 100$
- D. $\frac{0.142}{0.147} \times 100$

3. Which temperature change would cause the volume of a sample of an ideal gas to double when the pressure of the sample remains the same?

- A. From 200 °C to 400 °C
- B. From 400 °C to 200 °C
- **C. From 200 K to 400 K**
- D. From 400 K to 200 K

4. How much pressure exists in a 10.00 L container filled with 2.50 moles of gas at 315.0 K?

- A. 0.000955 kPa
- B. 0.00153 kPa
- **C. 654 kPa**
- D. 1050 kPa

5. A gas sample occupies a container of 25.0 mL with a pressure of 125.1 kPa at room temperature (25.0 °C). The number of moles of this gas in the container is

- **A. 0.00126 moles**
- **B.** 0.0151 moles
- **C.** 1.26 moles
- **D.** 15.1 moles

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Answers section.

1. The ideal gas law combines four variables that describe a gas. These variables are pressure, volume, temperature, and number of moles. The ideal gas law is **PV = nRT**. The universal gas constant, *R*, is equal to 8.31 L·(kPa)/(K·mol). This constant is based on **standard** temperature and pressure (STP), where the temperature is 0.0 °C or **273.15 K** and the pressure is 1 atm or **101.315 kPa**. When using the ideal gas law, temperature must always be measured in **Kelvin**.

2. The **ideal gas law** can be used to experimentally determine the number of moles generated from a chemical reaction. The reaction must take place inside a **closed** container and the temperature, pressure, and **volume** of the container must be recorded. The ideal gas law is most accurate for gases that behave **ideally**. Real gases behave ideally at high **temperatures** and low **pressures**.

Extended Inquiry Suggestions

Generate a different gas (H_2) using magnesium ribbon and dilute hydrochloric acid and collect the same type of data. Compare the results and describe any differences between the two experiments.

Repeat the experiment at various temperatures (for example, outside on a cold/hot day or in a walk-in freezer or under heat lamps). Compare the results to those acquired at room temperature.

Find all the possible values for the universal gas constant. Using the appropriate form of the universal gas constant, explore the effects of collecting data in various units. Do different units for pressure affect the end result for the number of moles produced?

24. Heats of Reaction and Solution

Objectives

Determine the molar heat of solution for sodium hydroxide and ammonium chloride when they are dissolved in water, and the molar heat of reaction when magnesium reacts with hydrochloric acid. Through this investigation, students:

- Calculate the molar heat changes ΔH in physical and chemical processes
- Review exothermic and endothermic processes
- Write equations that show the molar heat changes in physical and chemical processes

Procedural Overview

Students will conduct the following procedures:

- Record temperature versus time data for ammonium chloride dissolving in water, for sodium hydroxide dissolving in water, and for magnesium metal reacting with hydrochloric acid
- Analyze the temperature versus time data to determine the change in temperature after correcting for heat loss
- Calculate molar heat (enthalpy) changes and compare them to accepted values

Time Requirement

♦ Preparation time	10 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	45 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Temperature sensor[1]
- Beaker, 250-mL
- Graduated cylinder, 50-mL
- Balance, centigram
- Polystyrene cup (2)
- Spatula
- Stir rod
- Paper towels
- Weighing paper
- Sand paper or steel wool, 1 piece
- Wash bottle and waste container
- Sodium hydroxide (NaOH) pellets, 1 g
- Ammonium chloride (NH_4Cl), 1 g
- Magnesium metal ribbon, 0.10 g
- 1.0 M Hydrochloric acid (HCl), 25 mL[2]
- Distilled (deionized) water, 50 mL

[1]Either the fast response or stainless steel temperature sensor is suitable.

[2]To prepare using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Writing chemical equations
- Exothermic and endothermic processes
- Calorimetry
- Heat capacity
- Enthalpy

Related Labs in This Guide

Labs conceptually related to this one include:

- Evidence of a Chemical Reaction
- Heat of Fusion
- Single Replacement Reactions
- Hess's Law

Using Your Data Collection System

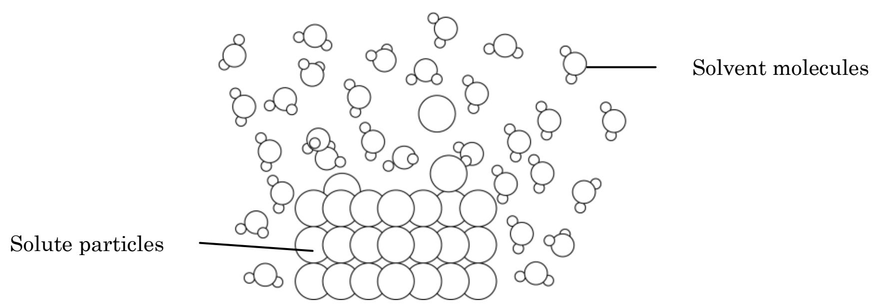
Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to your data collection system �(2.1)
- Recording a data run �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Selecting data points in a graph �(7.1.4)
- Showing and hiding data runs in a graph �(7.1.7)
- Naming a run of data �(8.2)
- Finding the values of a point in a graph �(9.1)
- Applying a curve fit �(9.5)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

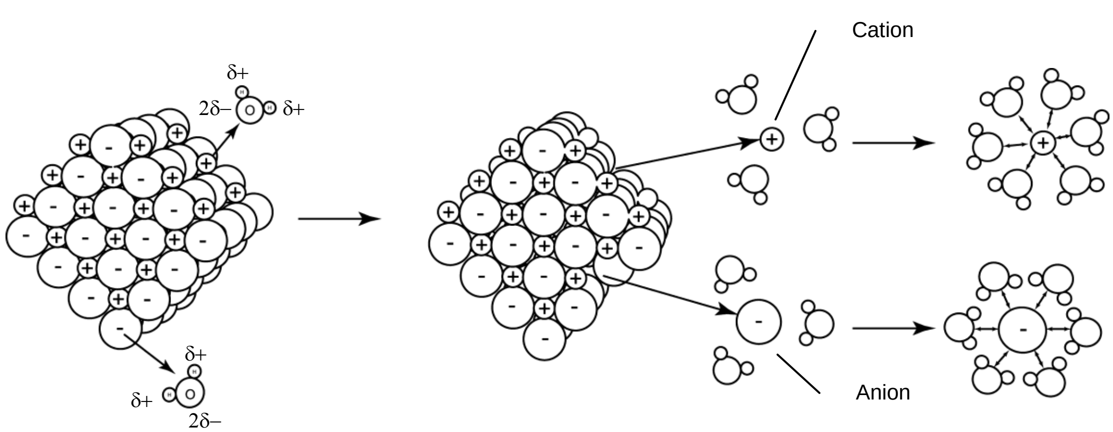
Physical changes and chemical reactions both may be accompanied by changes in energy, often in the form of heat. Chemical and physical processes that absorb heat from their surroundings are endothermic. Those that release heat into their surroundings are exothermic. Because most physical changes and chemical reactions occur at constant pressure, chemists can use the term enthalpy *H* for the heat energy released or absorbed by a system.

Phase changes require the addition or removal of heat energy. Another common physical change that involves changes in heat is dissolution, the process of dissolving a solid substance into a solvent to form a solution. When a substance dissolves, energy is required to “free” ions that are locked in their crystal lattice structure.



Solute dissolving in water.

Energy is then released when the “free” ions form attractions with molecules of the solvent. Several solvent molecules surround the free ion, creating a “solvent cage” around each ion in solution (solvation).



Formation of solvent cages.

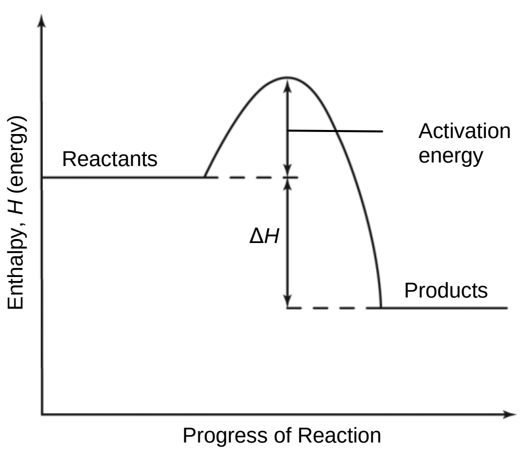
The heat absorbed by or released due to the process of dissolution is called the “heat of solution” (or enthalpy of solution). When the energy released from forming “solvent cages” is greater than the energy absorbed to break the attractions in the crystal lattice, the extra heat is released to the surroundings, and the process is exothermic. When the energy absorbed by the crystal lattice is greater than the energy released by forming "solvent cages" then the process absorbs the extra energy from the surroundings, and the process is endothermic.

Substances that spontaneously dissolve even though they require energy from their surroundings (endothermic, $\Delta H > 0$) do so because they are driven by the resulting increase in disorder (entropy, $\Delta S > 0$). You can see this when employing the equation for Gibbs free energy ($\Delta G = \Delta H - T\Delta S$, process is spontaneous if $\Delta G < 0$).

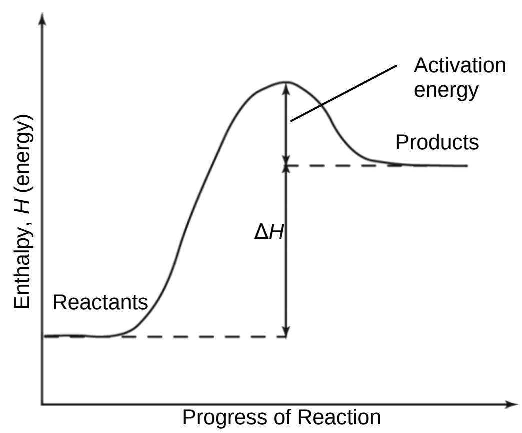
The heat that is exchanged per mole of substance during a chemical reaction is called the "molar heat of reaction" (or molar enthalpy of reaction) for that substance. The heat of reaction results from the difference in energy arising from the breaking of bonds (requires energy) in the reactants and the forming of bonds (releases energy) in the products.

Exothermic chemical reactions release energy to their surroundings because the products contain less internal energy (more stable bonds) than the reactants. On the other hand, endothermic chemical reactions absorb energy from their surroundings because the products contain more internal energy (less stable bonds) than the reactants. These differences are illustrated in the enthalpy diagrams below.

Enthalpy Diagram for an Exothermic Chemical Reaction



Enthalpy Diagram for an Endothermic Chemical Reaction



In addition to showing the energy differences between reactants and products, enthalpy diagrams also show the activation energy. The activation energy is the minimum amount of energy particles must have when they collide in order to react. If particles collide with energy less than the activation energy, bonds will not break, and products will not form.

In this experiment, students calculate the molar heat of solution for sodium hydroxide (NaOH), the molar heat of solution for ammonium chloride (NH_4Cl), and the molar heat of reaction for magnesium metal reacting with hydrochloric acid (HCl). In order to calculate these values, students will carefully weigh the reactants, and then collect temperature data as the reactants progress to products in a simple polystyrene cup calorimeter. The heat q of the process is then calculated by multiplying the mass m of the solution, the specific heat capacity c of water, and the change in temperature ΔT.

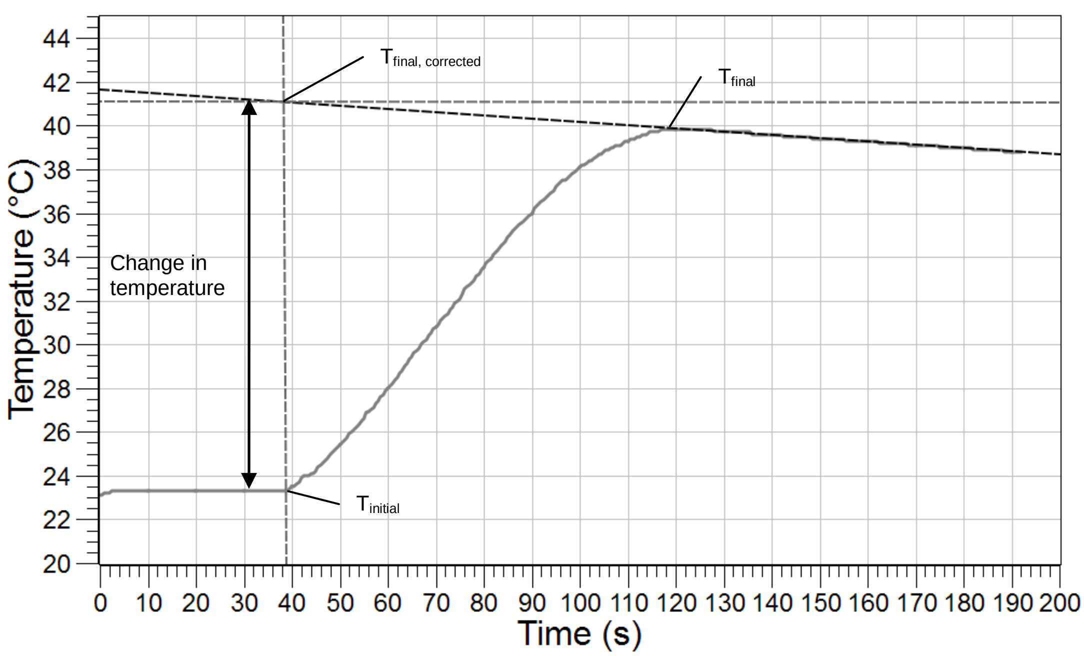
$$q = m \times c \times \Delta T$$

Because the pressure is constant, the heat released or absorbed can be replaced with the change in enthalpy ΔH. Because the solution in the calorimeter is actually part of the reaction's surroundings, a change in sign is required. Any *gain* in energy for the solution (seen as an increase in its temperature) is the result of a *loss* of energy from the reaction being studied. The equation for ΔH is given below.

$$\Delta H = -q$$

The polystyrene cup (as with all calorimeters) is not 100% efficient and will allow some heat to escape. Students will compensate for heat lost from the calorimeter by extrapolating at the same rate of cooling to determine the final temperature that would have been reached if no heat had been able to escape from the calorimeter. The value for ΔT will be the temperature that would have been reached $T_{\text{final, corrected}}$ minus the initial temperature T_{initial}.

Extrapolation at the Same Rate of Cooling to Find the Final Corrected Temperature



Pre-Lab Discussion and Activity

Heat of Solution Demonstration – Dissolve Potassium Nitrate in Water

Have your students predict whether or not there will be a temperature change when potassium nitrate (KNO_3) dissolves in water. Project a temperature versus time graph as you mix 5 g of potassium nitrate (KNO_3) in 200 mL of water. Discuss the results, and compare this physical process to other physical processes that release or absorb energy. Write the physical process that occurred as an equation using heat as either a product or a reactant.

1. What do you think will happen to the temperature as the salt, potassium nitrate (KNO_3), dissolves in water? Why?

Answers will vary, but the temperature will decrease.

2. Is dissolving a physical change or a chemical reaction? How do you know?

Dissolving is a physical change because no new substance is formed.

3. Was dissolving KNO_3 in water an exothermic or endothermic process? How do you know?

It is an endothermic process because the temperature decreased. Energy was absorbed from the surroundings.

4. Is the heat flowing from the KNO_3 to the water or from the water to the KNO_3?

The KNO_3 is absorbing the heat from the water. The heat flows from the water to the KNO_3.

5. Is the enthalpy change (ΔH) positive or negative for this reaction?

The enthalpy change is positive because this is an endothermic reaction.

6. How can you summarize the heat change involved in the dissolving of KNO_3 in a chemical equation? Is heat a reactant or a product?

$KNO_3(s) + \text{heat} \rightarrow K^+(aq) + NO3^-(aq)$

Heat is a reactant because it is absorbed.

7. What are some other physical processes that release or absorb heat? Give an example of how you can symbolize the heat change involved in the physical process in a chemical equation.

Phase changes are all accompanied by changes in heat. Melting and boiling are endothermic while freezing and condensing are exothermic.

For example, melting ice: $H_2O(s) + \text{heat} \rightarrow H_2O(l)$

For example, condensation of water: $H_2O(g) \rightarrow H_2O(l) + \text{heat}$

Dissolving at the Molecular Level

Engage your students in a discussion about why heat changes when a substance dissolves. Review with your students that crystal lattice structures must be broken apart (requiring energy), and that new forces of attraction form between the "free" particles and the solvent molecules (releasing energy). Use the pictures provided in the Background section to help students visualize what occurs at the molecular level.

8. What forces of attraction are holding KNO_3 in its solid, crystal lattice?

Ionic bonds between the potassium ions and nitrate ions hold the solid salt together.

9. What happens at the molecular level when you place the solid KNO_3 in water?

The water molecules surround the KNO_3 crystal lattice structure and form intermolecular attractions between the polar water molecules and the ions that are on the surface of the KNO_3 crystal lattice. The surrounded ions are then carried into the solution, and the process continues until the entire crystal has dissolved or the solution becomes saturated.

10. Is energy released or absorbed when intermolecular attractions are formed?

Energy is released when intermolecular attractions are formed.

11. Is energy released or absorbed in order for the ionic bonds in KNO_3 to break?

Energy needs to be absorbed to break the bonds.

12. What is the source of the energy to break the ionic bonds in KNO_3?

The energy needed to break the ionic bonds comes from the surroundings (the water making up the solution).

13. Once an ion is completely freed from its crystal lattice, what happens to it?

Solvent molecules surround the ion, forming a solvent cage.

14. Why did the temperature decrease when KNO_3 dissolved in water? What happens at the molecular level?

The amount of energy released by the formation of solvent cages was less than the amount of energy needed to break the ionic bonds in the crystal lattice. The heat (energy) was absorbed from the environment to make up the difference, thus cooling the solution.

Heat of Reaction Demonstration – Zinc Powder and Copper Sulfate Solution

Have your students predict whether or not a temperature change occurs when you add zinc powder to a copper sulfate ($CuSO_4$) solution. Project a temperature versus time graph as you mix 2 g of zinc powder in 100 mL of 0.5 M $CuSO_4$ solution. Be sure to record the exact mass of the zinc and copper sulfate solution so that you can calculate the heat of reaction. As the reaction occurs (5 to 10 minutes), discuss what is happening at the molecular level. When the reaction finishes, discuss the results. Compare the similarities and differences between the heat changes associated with chemical reactions, and those of physical changes.

15. What do you think will happen to the copper sulfate solution when you add zinc powder? Why?

Answers will vary. The blue color of the copper sulfate solution fades as the copper ions are reduced to solid copper. The solid metal in the solution will show a slight color change as the zinc powder is replaced by copper.

16. What products will form when you add zinc powder to copper sulfate solution?

This is a single replacement reaction. Zinc is higher on the activity series and will displace the copper.

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

17. What type of bonds hold the zinc atoms together in the solid state?

Metallic bonds hold the zinc atoms together.

18. What needs to be added in order to break the metallic bonds?

Energy needs to be added.

19. Where does the solid copper come from?

The copper ions in the solution get reduced to solid copper metal.

20. Is energy released or absorbed when solid copper forms?

Energy is released when copper forms.

21. **Is the overall reaction endothermic or exothermic? Explain your reasoning at the molecular level.**

The reaction is exothermic because the energy required to break the bonds is less than the energy released when the new bonds form.

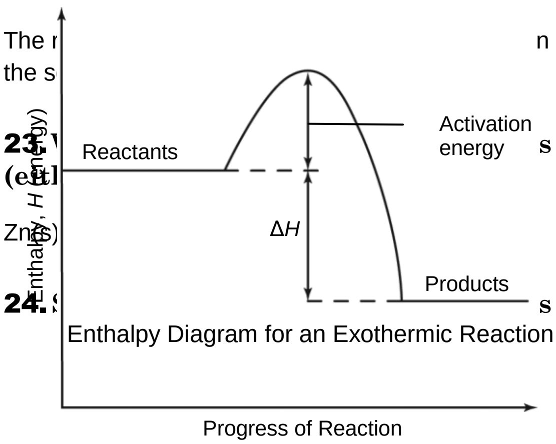
22. **Do the reactants or products contain more energy? How do you know?**

The r… n occurs, some energy is released into the surroundings and the s…

23. **W… s reaction showing heat in its proper location (ent…**

Zn(s)…

24. **S… s reaction.**



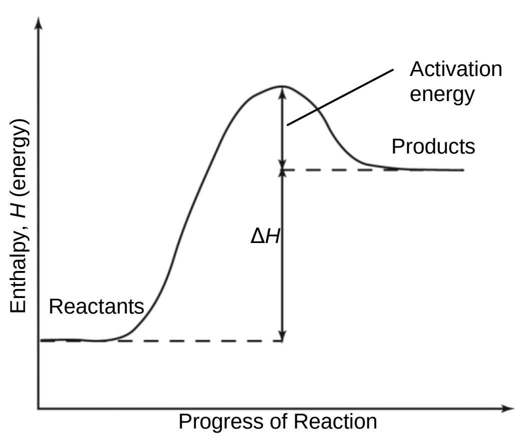
Enthalpy Diagram for an Exothermic Reaction

25. **In an endothermic reaction, do the reactants or products contain more energy? Explain your reasoning at the molecular level.**

The products contain more energy in an endothermic reaction. The energy needed to break the bonds in the reactants is greater than the energy released when forming the products. The extra energy required is drawn from the surroundings, which causes a decrease in the temperature of the solution.

26. **Sketch a generalized enthalpy diagram for an endothermic reaction.**

Enthalpy Diagram for an Endothermic Reaction



Determining the Experimental Molar Heat of Reaction

Engage your students in a discussion about the degree to which the simple polystyrene calorimeter stops heat transfer between the system and the surroundings. Discuss how the heat lost to the environment will affect the quantity of heat calculated. Explain how to use extrapolation of the cooling curve to correct for heat lost. Use data from the zinc/copper sulfate demonstration as an example of how to experimentally determine the molar heat of reaction for zinc, adding in the correction for heat lost through the calorimeter.

27. Why is a calorimeter used when making heat calculations?

Calorimeters help minimize the amount of heat that transfers between the system and the surroundings.

28. Do you think this simple polystyrene calorimeter completely eliminates the exchange of heat between the inside and the outside of the calorimeter? Explain your reasoning.

No, the calorimeter does not stop the exchange of heat. Without a lid, heat can easily be exchanged through the top of the calorimeter. Heat can also be lost through the sides of the container, which you can feel when you hold a hot beverage in a polystyrene cup.

29. Does the graph of Temperature (°C) versus Time (s) show any evidence that heat was being lost? What does this mean?

After achieving the highest temperature, the temperature began to decrease. If heat could not escape through the calorimeter, then the temperature would have remained constant at the highest temperature obtained. This means that the highest temperature observed is actually lower than it should be because some heat was lost to the surroundings outside of the calorimeter.

30. How can you use the slope of the cooling rate to correct for the heat lost?

If you assume that the heat lost from the calorimeter occurs at a constant rate, then you can extrapolate the cooling curve back to when the reaction began to find the actual temperature that would have been achieved if no heat was lost (corrected temperature).

31. Use this corrected temperature to find the temperature change, calculate the molar heat of the reaction (enthalpy of reaction), and find the percent error.

Step 1: Find q using the equation: $q = m \times c \times \Delta T$

Q = heat lost or gained by the solution

m = mass of the solvent (copper sulfate solution)

c = the specific heat of the solution (use the specific heat of water, 4.18 J/(g·°C))

$\Delta T = T_{\text{final, corrected}} - T_{\text{initial}}$

Step 2: Find ΔH using the equation, $\Delta H = -q$

Step 3: Find the number of moles of zinc

Step 4: Find the molar enthalpy of zinc: ΔH/mol Zn

Step 5: Find the percent error using the literature value for the molar enthalpy of zinc, –218 kJ/mol

$$\textbf{percent error} = \left|\frac{(\textbf{accepted value} - \textbf{experimental value})}{\textbf{accepted value}}\right| \times 100$$

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid or base solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids and bases creates heat; be extra careful when handling freshly prepared solutions and glassware as they may be very hot.
- ♦ Always add acids and bases to water, not the other way around, as the solutions may boil vigorously.
- ♦ Handle concentrated acids and bases in a fume hood; the fumes are caustic and toxic.

Make the following solution:

Make 1000 mL of 1.0 M hydrochloric acid from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 100 lab groups.

Starting with concentrated (12 M) HCl:

1. Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.
2. Slowly add 83.3 mL of 12 M HCl to the beaker with continuous stirring.
3. Allow the solution to cool, and then carefully pour it into a 1000-mL volumetric flask, and dilute to the mark with distilled water.
4. Cap and invert three times carefully to ensure complete mixing.

Starting with dilute (6 M) HCl:

1. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.
2. Add 166.7 mL of 6 M HCl to the water, and dilute to the mark with distilled water.
3. Cap and invert three times carefully to ensure complete mixing.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Sodium hydroxide and hydrochloric acid are corrosive irritants. Avoid contact with the skin and eyes.
- ♦ Be sure that all acids and bases are neutralized before disposing of them down the drain.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	5	3	4
After recording the exact mass of liquid in the calorimeter, put a temperature sensor in the liquid and start recording data.	Write down the exact mass of each reactant (both the solid and the water).	Correct for heat loss and then calculate the heat of reaction or solution.	After the temperature stabilizes, add the solid to the liquid in the calorimeter and mix thoroughly.	Stop recording temperature data after the temperature has reversed for at least 1 minute.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Part 1 – Dissolving Sodium Hydroxide (NaOH) and Ammonium Chloride (NH_4Cl)

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a temperature sensor to the data collection system. �(2.1)

3. ☐ Display Temperature (°C) versus Time (s) on a graph. �(7.1.1)

4. ☐ Measure approximately 1.0 g of sodium hydroxide (NaOH) pellets.

5. ☐ Record the exact mass of NaOH in the data table below.

Table 1: Data collected for dissolving sodium hydroxide in water

Mass of NaOH (g)	0.97
Volume of ~25 mL of Water (mL)	24.5
Mass of Water (g)	25.07

6. ☐ Measure approximately 25 mL of water using a graduated cylinder, and write the exact volume in the data table above.

7. ☐ Place a clean, dry polystyrene cup into another polystyrene cup to create a simple calorimeter. Place the calorimeter on the balance and tare the balance.

8. ☐ Pour the approximately 25 mL of water into the calorimeter.

9. ☐ Write the mass of the water in the data table above.

10. ☐ What does it mean to "tare" a balance? How would you find the mass of the water if you did not "tare" the balance?

"Tare" means to set the mass reading to zero while the object is on the balance. If you did not "tare" the balance, you would first have to find the mass of the empty cup and then the mass of the cup filled with the water. To find the mass of the water alone, subtract the mass of the empty cup from the mass of the cup and water.

cup + water
– cup
water

11. ☐ Compare the volume of the water with its mass. Is this what you would expect? Explain.

The numeric value of the volume and mass of the water should be the same. We expect this because the density of water is 1.0 g/mL at room temperature.

12. ☐ Place the calorimeter containing the approximately 25 mL of water into a 250-mL beaker.

Collect Data

13. ☐ Place the temperature sensor in the water that is in the calorimeter.

14. ☐ While viewing the graph display, start recording data. �(6.2)

Note: Allow the temperature to stabilize (remains constant for at least 30 seconds).

15. ☐ Add the solid NaOH to the water, and continuously stir the mixture until the NaOH completely dissolves. This may take several minutes.

Note: You may need to adjust the scale of the axes to see the changes taking place. �(7.1.2)

16. ☐ Stop recording data once the temperature has decreased for at least one minute after reaching its maximum temperature. �(6.2)

17. ☐ Name the data run "NaOH". �(8.2)

18. ☐ Dispose of the NaOH solution according to the teacher's instructions.

19. ☐ Clean and thoroughly dry the calorimeter, temperature sensor, and stir rod.

20. ☐ Why must the calorimeter, temperature sensor, and stir rod be dried before using them in the next part of the experiment?

You need dry equipment because any residual water will add an unknown amount to the quantity of water measured in the next part of the experiment. This additional, unaccounted water will absorb or release heat, affecting the results of the experiment (by giving an unreliable value for m in the equation, $q = m \times C \times \Delta T$), thus reducing the accuracy of the final calculated result.

21. ☐ Repeat the steps in the Set Up and Collect Data sections, this time substituting ammonium chloride (NH_4Cl). Take into account the following differences when you repeat the steps:

- Use approximately 1.0 g of ammonium chloride and approximately 25 mL of water.
- Record the data collected in the table below.
- Stop recording data when the temperature has increased for one minute or longer after reaching its minimum temperature. �(6.2)
- Name the data run "NH_4Cl". �(8.2)

Table 2: Data collected for dissolving ammonium chloride in water

Mass of NH_4Cl (g)	1.08
Volume of ~25 mL of Water (mL)	24.9
Mass of Water (g)	25.17

22. ☐ Dispose of the NH_4Cl solution according to the teacher's instructions.

23. ☐ Thoroughly clean and dry the calorimeter, temperature sensor, and stir rod.

Part 2 – Reacting Magnesium Metal and Hydrochloric Acid

Set Up

24. ☐ Cut a piece of magnesium ribbon 7- to 8-cm long (approximately 0.1 g). Use sandpaper or steel wool to remove any magnesium oxide that has formed on the magnesium ribbon.

25. ☐ Cut the cleaned magnesium ribbon into approximately 1-cm pieces.

26. ☐ Measure the mass of all the pieces together and record the mass in the data table below.

Table 3: Data collected for reacting magnesium metal and hydrochloric acid

Mass of Magnesium Metal Pieces (g)	0.10
Volume of ~25 mL of 1.0 M HCl (mL)	25.6
Mass of ~25 mL of 1.0 M HCl (g)	26.19

27. ☐ Why is the magnesium ribbon cut into pieces?

The magnesium ribbon is cut into pieces to make the reaction occur faster.

28. ☐ Measure approximately 25 mL of 1.0 M hydrochloric acid (HCl) using a graduated cylinder. Record the exact volume in the data table above.

29. ☐ Place a clean, dry calorimeter on the balance and tare the balance.

30. ☐ Pour the approximately 25 mL of 1.0 M HCl into the calorimeter.

31. ☐ Record the mass of the HCl in the data table above.

32. ☐ Why is a polystyrene cup used as the calorimeter instead of a beaker? Why is the polystyrene cup placed inside the beaker?

Polystyrene is a better insulator than glass; the polystyrene is better at keeping the heat trapped inside the calorimeter. You place the polystyrene cup inside the beaker because the air trapped between the walls of the cup and the beaker serves as additional insulation. The beaker will also provide more stability so the cup will be less likely to tip over.

Collect Data

33. ☐ While viewing the graph display, start recording data. �(6.2)

Note: Allow the temperature to stabilize (remain constant for at least 30 seconds).

34. ☐ Add the magnesium ribbon pieces, and stir until the magnesium has completely reacted. The reaction may take several minutes to occur.

Note: You may need to adjust the scale of the axes to see the changes taking place. �(7.1.2)

35. ☐ Stop recording data once the temperature has decreased for at least one minute after reaching its maximum temperature. �(6.2)

36. ☐ Record any observations you witnessed that suggest a chemical reaction occurred.

Heat was released to the surroundings, and a gas (hydrogen, H_2) was produced (seen as bubbles) as a result of the single replacement reaction:

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

37. ☐ Name the data run “Mg”. �(8.2)

38. ☐ Save the data file and clean up the lab station according to the teacher’s instructions. �(11.1)

Data Analysis

1. ☐ Determine the initial temperature ($T_{initial}$), the highest temperature actually attained ($T_{final, actual}$), and the temperature that would have been reached if there was no heat lost to the surroundings ($T_{final, corrected}$) for each run of data collected. Follow the steps below to do this on your data collection system.

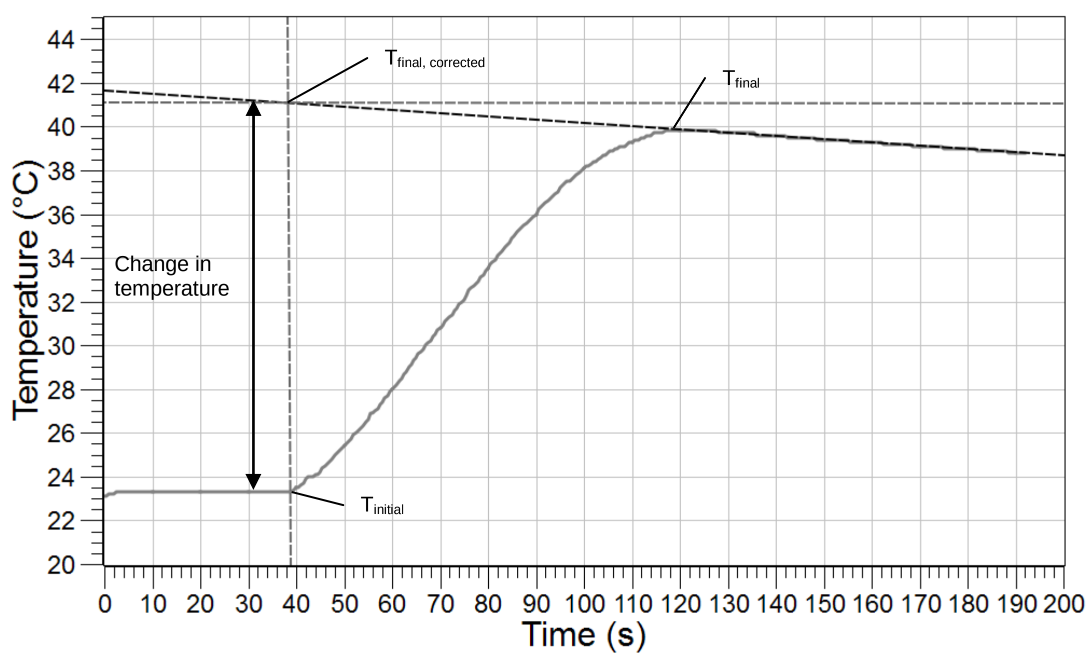
- **a.** Display the run of data you want to analyze. �(7.1.7)
- **b.** Find the initial temperature and the highest temperature actually attained by finding the coordinates at each of these points on the graph. �(9.1)
- **c.** Record the initial and final (actual) temperatures in Table 4 below.
- **d.** Select all the data points that were collected after the highest (including the highest temperature). �(7.1.4)
- **e.** Apply a linear fit to these selected data points. �(9.5)
- **f.** Adjust the scale of the graph so that you can find the point where the linear fit line crosses an imaginary vertical line extending up from the initial temperature. �(7.1.2)

Note: This is illustrated in the sample graph below.

- **g.** Record the final corrected temperature in Table 4 below.

<u>Sample Graph</u>

Extrapolation at the Same Rate of Cooling to Find the Final Corrected Temperature



Note: The difference between the final temperature and the corrected final temperature is because the solution was losing heat as the reaction occurred. The longer the reaction takes to occur the greater the extent of heat loss.

Table 4: Initial, final actual, and final corrected temperature values

Dissolving Sodium Hydroxide		Dissolving Ammonium Chloride		Reacting Magnesium Metal with Hydrochloric Acid	
T_{initial} (˚C)	22.4	T_{initial} (˚C)	23.1	T_{initial} (˚C)	23.3
$T_{\text{final, actual}}$ (˚C)	29.0	$T_{\text{final, actual}}$ (˚C)	20.1	$T_{\text{final, actual}}$ (˚C)	39.8
$T_{\text{final, corrected}}$ (˚C)	29.8	$T_{\text{final, corrected}}$ (˚C)	20.1	$T_{\text{final, corrected}}$ (˚C)	41.2

2. ☐ Determine the change in temperature for each process by subtracting the initial temperature from the final corrected temperature. Show your work and record you answers in Table 5 below.

Table 5: Change in temperature

Process	Show Your Work $T_{\text{final, corrected}} - T_{\text{initial}}$ (°C)	Change in Temperature ΔT (°C)
Dissolving sodium hydroxide	29.8 – 22.4	7.4
Dissolving ammonium chloride	20.1– 23.1	–3.0
Reacting magnesium metal with hydrochloric acid	41.2– 23.3	17.9

3. ☐ Calculate the heat absorbed by the solution in each process, q, by using the formula given below. Convert joules to kilojoules in your final answer. Show all of your work, including the units, and record your answers in Table 6 below.

$q = m \times c \times \Delta T$, where:

q = heat lost or gained by the solution

m = mass of the solvent (water in the dissolving processes and HCl in the reaction)

c = the specific heat of the solution (use the specific heat of water, 4.18 J/(g·°C))

$\Delta T = \mathrm{T}_{\text{final, corrected}} - \mathrm{T}_{\text{initial}}$

Table 6: Heat absorbed by the solution in each process

Process	Show Your Work here:	Heat q (kJ)
Dissolving sodium hydroxide	$q = (25.07\ \text{g})\left(\frac{4.18\ \text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(7.4\ {}^\circ\text{C}) = 775.5\ \text{J}$	0.78
Dissolving ammonium chloride	$q = (25.17\ \text{g})\left(\frac{4.18\ \text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(-3.0\ {}^\circ\text{C}) = -315.6\ \text{J}$	–0.32
Reacting magnesium metal with hydrochloric acid	$q = (26.19\ \text{g})\left(\frac{4.18\ \text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(17.9\ {}^\circ\text{C}) = 1959.6\ \text{J}$	1.96

4. ☐ Find the molar enthalpy for each substance. Show your work and record your answers in Table 7 below.

$$\text{molar enthalpy of solution/reaction} = \frac{\Delta H}{\text{moles of substance}}$$

Note: The amount of heat absorbed or released by the solution is the opposite of the amount of heat absorbed or released by the process.

$$\Delta H = -q$$

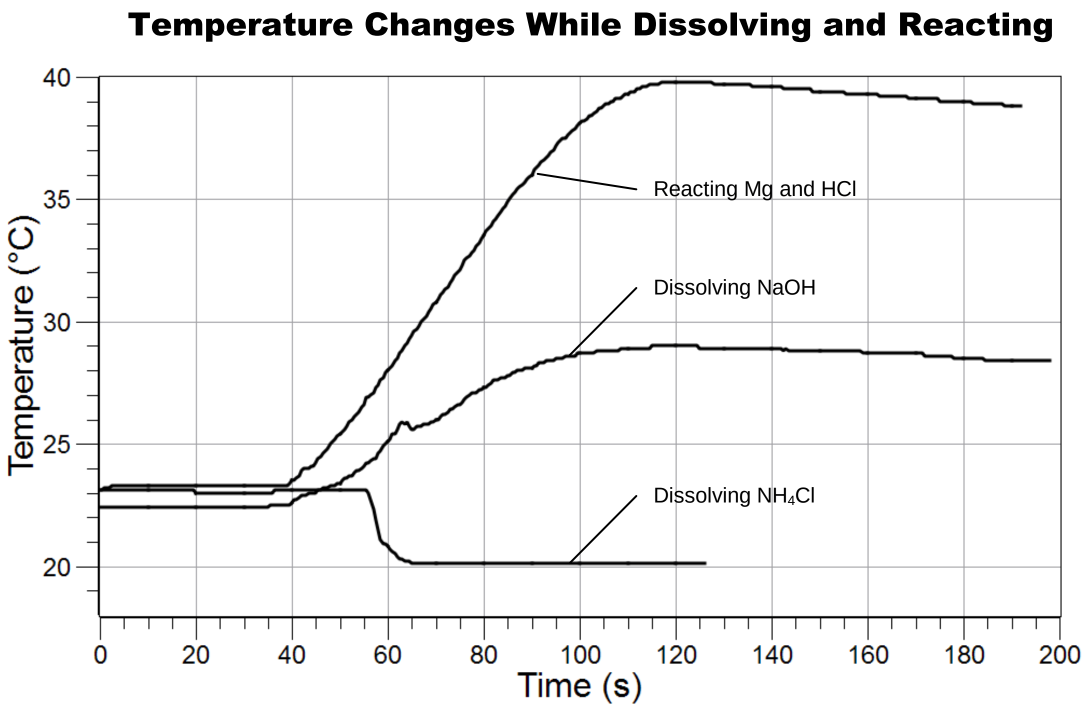
Table 7: Molar enthalpy for each substance

Process	Calculate Number of Moles (Show your work)	ΔH (kJ)	Calculate Molar Enthalpy, ΔH/mol (kJ/mol) (Show your work)
Sodium hydroxide	$(0.97\ \text{g NaOH})\left(\frac{1\ \text{mol NaOH}}{40.00\ \text{g NaOH}}\right) = 0.02425\ \text{mol NaOH}$ $= 0.024\ \text{mol NaOH}$	–0.78	$\frac{-0.78\ \text{kJ}}{0.02425\ \text{mol}} = -32$
Ammonium chloride	$(1.08\ \text{g}\ NH_4Cl)\left(\frac{1\ \text{mol}\ NH_4Cl}{53.49\ \text{g}\ NH_4Cl}\right) = 0.020191\ \text{mol}\ NH_4Cl$ $= 0.0202\ \text{mol}\ NH_4Cl$	0.32	$\frac{0.32\ \text{kJ}}{0.020191\ \text{mol}} = 16$
Magnesium	$(0.10\ \text{g Mg})\left(\frac{1\ \text{mol Mg}}{24.31\ \text{g Mg}}\right) = 0.0041135\ \text{mol Mg}$ $= 0.0041\ \text{mol Mg}$	–1.96	$\frac{-1.96\ \text{kJ}}{0.0041135\ \text{mol}} = -480$

5. ☐ Create a graph with all three runs of data displayed on your data collection system. �(7.1.3)

Note: Not all data collection system will display all three runs of data. If this is not possible create one graph for part one (dissolving) and a second graph for part two (reacting).

6. ☐ Sketch or print a copy of your Temperature (°C) versus Time (s) graph with all three data runs on one set of axes. Label each data run as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Analysis Questions

1. Determine the percent error for each process using the following known values:

$$\text{percent error} = \left|\frac{(\text{accepted value} - \text{experimental value})}{\text{accepted value}}\right| \times 100$$

Process	Accepted ΔH/mole (kJ/mol)	Percent Error (Show your work)
Molar heat of solution for sodium hydroxide	–44.5	$\left\|\frac{(-44.5) - (-32)}{-44.5}\right\| \times 100 = 28\%$
Molar heat of solution for ammonium chloride	14.8	$\left\|\frac{(14.8) - (-16)}{14.8}\right\| \times 100 = 8.1\%$
Molar heat of reaction for magnesium metal with hydrochloric acid	–462.0	$\left\|\frac{(-462.0) - (-480)}{-462}\right\| \times 100 = 3.9\%$

2. Suggest possible changes to the experimental procedure that could improve the accuracy of the results.

You can take steps to reduce the heat lost through the calorimeter, such as adding a lid, using a cup with a thicker polystyrene wall, or replacing the polystyrene cup with a vacuum flask (like those used to keep coffee or soup hot). Additionally, you could use a more precise balance to obtain better values for the masses of solids and solutions used.

3. Why was it necessary to correct the final temperature reached?

As the reaction occurred, heat was being exchanged with the surroundings. Correcting the final temperature is one way to account for this error.

4. Identify each process as exothermic or endothermic, and state your evidence in each case.

Process	Endothermic or Exothermic?	Evidence
Dissolving sodium hydroxide	exothermic	heat was released, temperature increased
Dissolving ammonium chloride	endothermic	heat was absorbed, temperature decreased
Reacting magnesium metal with hydrochloric acid	exothermic	heat was released, temperature increased

5. Write the chemical equation that illustrates the heat changes that occurred for each physical and chemical process you performed in this experiment. Use your experimentally determined values for molar heat of reaction or solution in the equations. Be sure to include state symbols for all reactants and products.

$NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq) + 32\ kJ$

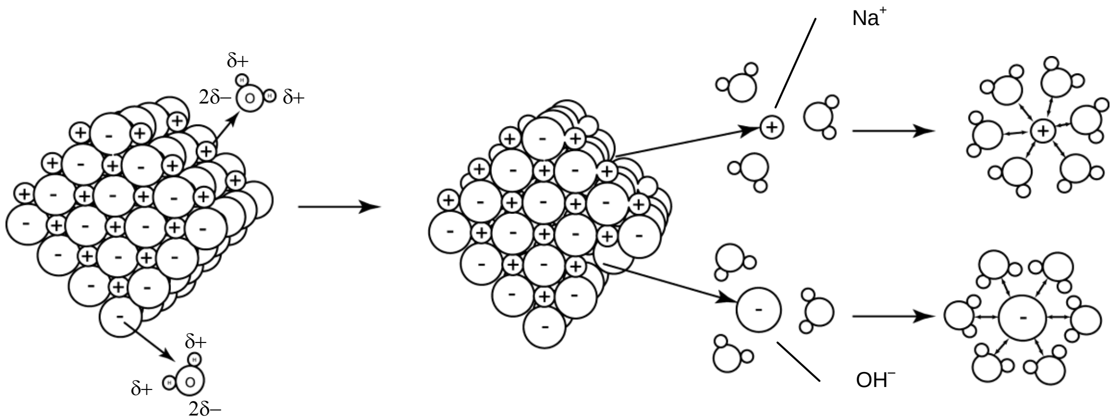
$NH_4Cl(s) + 16\ kJ \rightarrow NH_4^{+}(aq) + Cl^{-}(aq)$

$Mg(s) + 2HCl(aq) \rightarrow H_2(g) + MgCl_2(aq) + 480\ kJ$

Synthesis Questions

Use available resources to help you answer the following questions.

1. Draw a diagram that illustrates what happens at the molecular level to sodium hydroxide when it is dissolved in water.



2. Heat was either released or absorbed in all three processes, but only one of the processes was a chemical reaction. Explain how this is possible.

Energy is required in order to break attractions and released when attractions form. For the substances that were dissolving (NaOH and NH_4Cl), energy was required to separate the ions from one another and released when attractions formed between the ions and the solvent molecules. For the substances that were reacting (Mg with HCl), energy was required to separate the bonded atoms in the reactants and released when the atoms came together to form new bonds in the products.

3. Describe what would have happened if the physical and chemical processes were performed at 50 °C instead of room temperature?

The initial and final temperatures ($T_{initial}$, $T_{final,\ actual}$, and $T_{final,\ corrected}$) would all be greater, but the amount of heat absorbed or released (q and ΔH) would have been the same.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Heat of reaction refers to

- **A.** the heat released by a chemical reaction
- **B.** the heat absorbed by a chemical reaction
- **C.** The temperature of the solution after a chemical reaction occurs
- **D. Both A and B**

2. Using the chemical equation below, determine the energy released by burning 2 moles of propane (C_3H_8)?

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) + 2219.2 \text{ kJ}$$

- **A.** 2219.2 kJ
- **B.** **4438.4 kJ**
- **C.** 1109.6 kJ
- **D.** 11096 kJ

Use the paragraph and Table 8 below to answer Multiple Choice Questions 3 to 5 below.

Two grams of salt A were added to 50 mL of water, and the initial and final temperatures were recorded on the left side of Table 8 below. In a separate experiment, 2 g of salt B were added to 50 mL of water, and the initial and final temperatures were recorded on the right side of Table 8 below.

Table 8: Temperature data for Salt A and Salt B.

Salt A		**Salt B**	
Initial Temperature	**20 ˚C**	**Initial Temperature**	**20 ˚C**
Final Temperature	**35 ˚C**	**Final Temperature**	**10 ˚C**

3. What can be said about flow of energy in the two separate experiments?

- **A.** For salt A, the energy flows from the water to the salt. For salt B, the energy flows from the water to the salt
- **B.** For salt A, the energy flows from the salt to the water. For salt B, the energy flows from the salt to the water
- **C.** **For salt A, the energy flows from the salt to the water. For salt B, the energy flows from the water to the salt**
- **D.** For salt A, the energy flows from the water to the salt. For salt B, the energy flows from the salt to the water

4. What type of process occurred when salt A dissolved in water?

- **A.** Enthalpy
- **B.** Heat
- **C.** Endothermic
- **D.** **Exothermic**

5. Which equation illustrates the process that occurred when salt B dissolved in water?

- **A.** salt B(s) → salt B(aq) + heat
- **B.** salt B(aq) → salt B(aq) + heat
- **C.** **salt B(s) + heat → salt B(aq)**
- **D.** salt B(aq) + heat → salt B(aq)

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Physical changes and chemical reactions both may be accompanied by changes in **energy**, often in the form of heat or **enthalpy**. Chemical and physical processes that absorb heat from their surroundings are **endothermic**, while those that release heat into their surroundings are **exothermic**. The energy released or absorbed by different physical and chemical processes comes from the different internal energies of the starting substances and ending substances. In exothermic processes, the starting materials have **more** internal energy than the **products**. In endothermic processes, the **reactants** are lower in energy than the ending substances.

2. In chemical reactions, the amount of energy released per mole of material reacted is the molar **heat of reaction** of that material. Different physical processes have different names for the heat that is released based on the process that is taking place. For example, the amount of energy released per mole of material that is melting is called the molar heat of fusion. The amount of heat released per mole of substance being dissolved is called molar **heat of solution**. The molar heat of any physical or chemical process can be calculated by multiplying the change in **temperature**, the mass of the **solvent**, and the specific heat capacity of the solvent all divided by the **moles** of substance used. Exothermic processes have **negative** heats of reaction/solution while endothermic processes have **positive** heats of reaction/solution.

Extended Inquiry Suggestions

Improve the design of the calorimeter, and repeat the experiment.

Determine the effect of initial temperature on the heat of reaction/solution.

Determine the effects of the amount of reactants used and the change in temperatures observed as well as the resulting molar heats of reaction/solution.

Compare experimental heats of reactions with those calculated from heats of formation.

Lead a discussion on which types of fuels produce a greater heat of reaction when they are burned.

Determine the effect of initial temperature on the rate at which a substance dissolves or reacts by investigating endothermic and exothermic processes at both high and low temperatures.

25. Hess's Law

Objectives

Show that the change in enthalpy for the reaction between solid sodium hydroxide and aqueous hydrochloric acid can be determined using both a direct and an indirect method. Through this investigation, students:

- Measure initial and final temperatures and use the calculated temperature difference to determine the enthalpy change ΔH
- Confirm that the overall heat evolved or absorbed in a chemical process is the same whether the process takes place in one or several steps (Hess's law)
- Explain that directly measuring the heat evolved or absorbed in a chemical process is not always possible for a variety of reasons. Reasons include reactions that are too fast or too slow, are explosive, or involve toxic or expensive materials

Procedural Overview

Students conduct the following procedures:

- Measure temperature changes when chemicals are mixed
- Calculate molar enthalpies directly and indirectly using Hess's law
- Compare calculated molar enthalpies using direct and indirect measurements

Time Requirement

Preparation time	20 minutes
Pre-lab discussion and activity	20 minutes
Lab activity	45 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Temperature sensor[1]
- Beaker, 250-mL
- Graduated cylinder, 50-mL
- Spatula
- Polystyrene cup (2)
- Lid for polystyrene cup
- Weighing paper (2)
- Wash bottle and waste container
- 1.0 M Hydrochloric acid (HCl), 25 mL[2]
- 0.5 M Hydrochloric acid (HCl), 50 mL[3]
- 1.0 M Sodium hydroxide (NaOH), 25 mL[4]
- Sodium hydroxide (NaOH), pellets, 2.0 g
- Distilled (deionized) water, 50 mL

[1] Either the fast response or stainless steel temperature sensor is suitable.

[2] To formulate using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

[3] To formulate using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

[4] To formulate using solid sodium hydroxide (NaOH) pellets, refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Endothermic and exothermic reactions
- Heat calculations ($q = m \cdot c \cdot \Delta T$)
- Mole calculations

Related Labs in This Guide

Labs conceptually related to this one include:

- Heats of Reaction and Solution
- Heat of Fusion
- Specific Heat

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Naming a data run �(8.2)
- Viewing statistics of data, such as maximum and minimum �(9.4)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

Chemical and physical processes often involve changes in heat q. Processes that produce heat are exothermic and those that absorb heat are endothermic. Changes in temperature can be measured by performing the chemical reaction or physical process in an insulated container called a calorimeter.

Using an insulated container reduces the heat transfer to or from the surroundings, ensuring that the temperature change observed is solely the result of the process taking place inside the calorimeter. The heat of the process q is the product of the mass m, specific heat capacity c, and change in temperature ΔT.

$$q = m \times c \times \Delta T$$

Enthalpy H is the heat produced by a system at constant pressure. Since most experiments are carried out at constant pressure (atmospheric pressure), the enthalpy can be calculated using the same equation, but with a change in sign. The change in sign is required because the solution in the calorimeter is actually part of the reactions surroundings, but the enthalpy is produced by the reaction itself which is the system.

$$q = m \times c \times \Delta T$$
$$\Delta H = -q$$

Enthalpy changes resulting from chemical reactions can be determined through both direct and indirect measurements. Because heat is a state function, the enthalpy change of the overall reaction is equal to the sum of the enthalpy changes in the individual steps that make up the overall reaction; this is known as Hess’s law. A state function is one in which the amount of change depends only upon the starting and final positions. The processes and steps in between are irrelevant to the overall change.

Hess's law is useful because there are many circumstances when measuring the heat change of a chemical process is impossible. A direct measurement of heat may be impossible if the reactants involved are toxic, expensive, or rare. Reaction rates that are extremely fast (such as explosions) or very slow are difficult to measure accurately for the change in enthalpy.

The application of Hess's law allows alternate reactions to be studied. By selecting these alternate reactions as intermediate steps such that the summation of the individual steps gives the desired overall reaction, the enthalpy can be determined. The chemical equations can be manipulated following some guidelines:

- Reversing a reaction changes the sign on the enthalpy. For example, the endothermic reaction which separates two moles of water into its elements requires 571.6 kJ. The exothermic reaction which reverses the reaction produces 571.6 kJ.

Endothermic reaction:	$2H_2O(l) + \text{energy} \rightarrow 2H_2(g) + O_2(g)$	$\Delta H = +571.6$ kJ
Exothermic reaction:	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) + \text{energy}$	$\Delta H = -571.6$ kJ

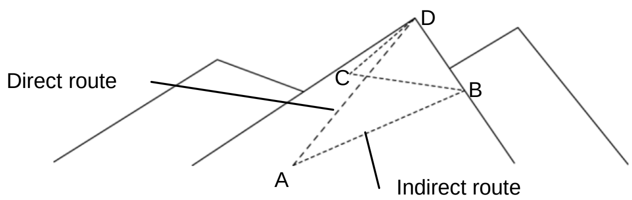
- Multiplying or dividing the chemical equation by a stoichiometric amount affects the enthalpy in a similar fashion. For example, doubling the formation of water produces twice the amount of energy, and halving the reaction produces half the amount of energy.

Doubling the reaction:	$2 \times (2H_2(g) + O_2(g) \rightarrow 2H_2O(l))$	$\Delta H = 2 \times (-571.6 \text{ kJ})$
	$4H_2(g) + 2O_2(g) \rightarrow 4H_2O(l)$	$\Delta H = -1143.2$ kJ
Halving the reaction:	$\frac{1}{2} \times (2H_2(g) + O_2(g) \rightarrow 2H_2O(l))$	$\Delta H = \frac{1}{2} \times (-571.6 \text{ kJ})$
	$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$	$\Delta H = -285.8$ kJ

Pre-Lab Discussion and Activity

Hiker Analogy

For this analogy, introduce two different hikers. The first hiker chooses the most direct route to the top of a mountain. The second hiker takes a path that is less direct, stopping along the way to enjoy the view and rest at different spots on the mountain. Engage the students in a qualitative and quantitative discussion about what the two paths have in common.



Elevation gains for each segment:
A-B 1,000 m,
B-C 300 m,
C-D 700 m.

1. What do the two hikes have in common?

The students should understand that the change in altitude for both hikers is exactly the same.

2. How much elevation did the hiker that went straight up the mountain gain? How do you know?

The hiker gained 2000 meters of elevation. This can be calculated by adding the three elevation gains of the second hiker.

3. Does the elevation gain depend on which route the hiker takes? Could a third route have been taken?

No, the elevation gain does not depend on the route taken. A third route could be taken and as long as the hiker started at point A and ended at point D the elevation gain would be the same.

Enthalpy of Chemical Reactions

Enthalpy is heat at constant pressure. Explain that enthalpy changes involved in chemical reactions are much like the route one takes hiking up a mountain. Multiple steps or reactions can be taken to get from a set of reactants to a set of products just like many different routes can be taken to get to the top of a mountain. When forming final products, the enthalpy change in each step (reaction) in the process can be added together to get the enthalpy change when the reactants are mixed all at once. This phenomenon is known as Hess's law. Use the reaction between magnesium metal and oxygen to form magnesium oxide to illustrate Hess's law.

4. According to Hess's law, how should ΔH_{rxn} of the indirect method compare with the ΔH_{rxn} of the direct method?

They should be the same.

5. How would you determine the change in enthalpy for the combustion of magnesium metal using the following data?

Direct:	$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$	$\Delta H = -1204$ kJ
Indirect, step 1:	$2Mg(s) + 4HCl(aq) \rightarrow 2H_2(g) + 2MgCl_2(aq)$	$\Delta H = -924$ kJ
Indirect, step 2:	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$	$\Delta H = -572$ kJ
Indirect, step 3:	$2MgCl_2(aq) + 2H_2O(l) \rightarrow 2MgO(s) + 4HCl(aq)$	$\Delta H = 292$ kJ

Add the three indirect equations, canceling equal reactants on opposite sides of the reaction arrows:

$$
\begin{array}{rrcll}
 & 2Mg(s) + \cancel{4HCl(aq)} & \rightarrow & \cancel{2H_2(g)} + \cancel{2MgCl_2}(aq) & \Delta H = -924 \text{ kJ} \\
+ & \cancel{2H_2(g)} + O_2(g) & \rightarrow & \cancel{2H_2O(l)} & \Delta H = -572 \text{ kJ} \\
 & \cancel{2MgCl_2}(aq) + \cancel{2H_2O(l)} & \rightarrow & 2MgO(s) + \cancel{4HCl(aq)} & \Delta H = 292 \text{ kJ} \\
\hline
 & 2Mg(s) + O_2(g) & \rightarrow & 2MgO(s) & \Delta H = -1204 \text{ kJ}
\end{array}
$$

Reason to Use Hess's Law

Engage the students in a discussion about the combustion of magnesium to understand the usefulness of Hess's law. Discuss measuring the enthalpy of a chemical reaction indirectly when it is not practical or possible to measure it directly.

Demonstrate the combustion reaction. Hold a strip of magnesium ribbon, 3- to 5-cm long, in a pair of clean tongs. Ignite one end using a gas burner in a dark room. The reaction produces a very bright, white light and white "smoke." The smoke is the magnesium oxide product and collects on the tongs coating them in white powder.

Avoid looking directly at the light produced. The white residue of magnesium oxide can be easily removed by wiping the tongs with a paper towel. This reaction was used in early flash powder photography as well as in "very lights" to illuminate battlefields in the First World War.

6. Is this reaction endothermic or exothermic?

This reaction produces a great amount of heat and light and is, therefore, exothermic.

7. Can the enthalpy of this reaction be determined directly using a simple polystyrene calorimeter?

Direct measurements of enthalpy may be impossible if the chemical reaction occurs too quickly or too slowly, or if the reactants are toxic, expensive, or rare. For the combustion of magnesium, the enthalpy of this reaction cannot be determined directly. In this case, the reaction would be extinguished by the water in the calorimeter. Additionally, the heat produced would boil the water or melt the polystyrene, destroying the calorimeter. An indirect method using Hess's law is required.

8. Can the enthalpy of this reaction be determined indirectly?

Yes. Hess's law can be applied using the alternate path (various chemical reactions) to make MgO.

Indirect, step 1: $2Mg(s) + 4HCl(aq) \rightarrow 2H_2(g) + 2MgCl_2(aq)$

Indirect, step 2: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Indirect, step 3: $2MgCl_2(aq) + 2H_2O(l) \rightarrow 2MgO(s) + 4HCl(aq)$

By adding the three indirect equations and canceling equal reactants on opposite sides of the reaction arrows, the overall, direct reaction, $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$, can be produced.

	$2Mg(s) + \cancel{4HCl(aq)}$	$\rightarrow$	$\cancel{2H_2(g)} + \cancel{2MgCl_2}(aq)$	$\Delta H = -924$ kJ
+	$\cancel{2H_2(g)} + O_2(g)$	$\rightarrow$	$\cancel{2H_2O(l)}$	$\Delta H = -572$ kJ
	$\cancel{2MgCl_2}(aq) + \cancel{2H_2O(l)}$	$\rightarrow$	$2MgO(s) + \cancel{4HCl(aq)}$	$\Delta H = 292$ kJ
	$2Mg(s) + O_2(g)$	$\rightarrow$	$2MgO(s)$	$\Delta H = -1204$ kJ

Hess's Law Experiment

Provide the students with an overview of the reactions to be performed in this investigation and how they relate to Hess's law.

For the direct method, students react solid sodium hydroxide pellets directly with aqueous hydrochloric acid to form sodium chloride and water. The change in enthalpy ΔH_{rxn} is calculated using the data collected during the reaction.

For the indirect method, students complete the same overall reaction, but do so using two different steps:

- **Step 1: Dissolve solid sodium hydroxide pellets in water to form a sodium hydroxide solution.**
- **Step 2: Mix a sodium hydroxide solution with aqueous hydrochloric acid.**
- **Calculate the change in enthalpy for step 1 ΔH_{step1} and step 2 ΔH_{step2}.**
- **Use Hess's law to calculate the enthalpy of the overall reaction using the equation below.**

$$\Delta H_{rxn} = \Delta H_{step\ 1} + \Delta H_{step\ 2}$$

Teacher Tip: Caution your students to be very careful to choose the right concentration of HCl. When the procedure calls for 1.0 M HCl, the students must use the 1.0 M HCl and not the 0.5 M HCl. Show both solutions, pointing out the difference between the clearly labeled bottles.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids creates heat. Be extra careful when handling freshly prepared solutions and glassware because they might be very hot.
- ♦ Always add acids to water, not the other way around, because the solutions may boil vigorously.
- ♦ Handle concentrated acids in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Make 1000 mL of 1.0 M HCl from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 40 lab groups.

Starting with concentrated (12 M) HCl:

a. Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.

b. Slowly add 83.3 mL of concentrated (12 M) HCl to the beaker with stirring.

c. Allow the solution to cool completely, then carefully pour the solution into a 1000-mL volumetric flask and dilute to the mark with distilled water.

d. Cap and invert three times to ensure complete mixing.

Starting with dilute (6 M) HCl:

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 166.7 mL of 6 M HCl to the water, and dilute to the mark with distilled water.

c. Cap and invert three times carefully to ensure complete mixing.

2. The following instructions make 1000 mL of 0.5 M HCl from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 20 lab groups.

Starting with (12 M) HCl:

a. Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.

b. Slowly add 41.7 mL of concentrated (12M) HCl to the beaker with stirring.

c. Allow the solution to cool completely, then carefully pour the solution into a 1000-mL volumetric flask and dilute to the mark with distilled water.

d. Cap and invert three times to ensure complete mixing.

Starting with (6 M) HCl:

- **a.** Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.
- **b.** Add 83.3 mL of 6 M HCl to the water, and dilute to the mark with distilled water.
- **c.** Cap and invert three times carefully to ensure complete mixing.

3. The following instructions make 1000 mL of 1.0 M NaOH from solid sodium hydroxide pellets. This is enough for 40 lab groups.

- **a.** Add approximately 500 mL distilled water to a 1000-mL beaker with a stir bar.
- **b.** Slowly add 40.0 g of solid NaOH pellets to the beaker with stirring.
- **c.** Allow the solution to cool completely, then carefully pour the solution into a 1000-mL volumetric flask and dilute to the mark with distilled water.
- **d.** Cap and invert three times to ensure complete mixing.

Safety

Add these important safety precautions to you normal laboratory procedures:

♦ Do not touch solid sodium hydroxide pellets with your hands. They are corrosive.

♦ Handle acids with care; they are corrosive.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

5	4	1	2	3
Use the data collected to calculate the molar enthalpy of the reaction directly and indirectly.	Dispose of the contents in the calorimeter. Repeat the experiment for the two remaining chemical reactions.	Measure and record the mass and volume of the starting materials.	Place the indicated starting material in the polystyrene calorimeter and begin collecting temperature data.	Add the second reactant and continue collecting data until the temperature decreases for 4 seconds.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect the temperature sensor to the data collection system. �(2.1)

3. ☐ Create a graph display of Temperature (°C) versus Time (s). �(7.1.1)

4. ☐ Place the polystyrene calorimeter (two polystyrene cups nested together) in the 250-mL beaker.

5. ☐ Place the temperature sensor inside the calorimeter.

Collect Data

Part 1 – Direct Reaction, One Step

6. ☐ Measure 50.0 mL of 0.50 M HCl in a clean graduated cylinder and carefully pour it into the calorimeter.

7. ☐ Record the exact volume of HCl used in Table 1 below.

Table 1: Direct reaction, solid sodium hydroxide pellets + 0.50 M hydrochloric acid

Volume (mL) of 0.50 M hydrochloric acid (HCl)	50.0
Mass (g) of sodium hydroxide (NaOH)	1.0
Initial temperature (°C)	24.0
Final temperature (°C)	32.8

8. ☐ Measure 1.0 g of solid sodium hydroxide (NaOH) and set it aside.

CAUTION: Do not touch the NaOH pellets. They burn your skin.

9. ☐ Record the exact mass in Table 1 above.

10. ☐ Why is it necessary to record the exact mass of sodium hydroixde used?

The amount of sodium hydroixde used counts as part of the mass of the solution. It also applies to calculating the energy per mole of NaOH.

11. ☐ Start recording data. �(6.2)

12. ☐ After 60 seconds, add the solid sodium hydroxide to the calorimeter.

13. ☐ Attach the lid to the calorimeter and swirl the apparatus until all the sodium hydroxide has dissolved.

14. ☐ Adjust the scale of the graph as necessary to see any changes taking place. �(7.1.2)

15. ☐ Continue to collect data until a maximum temperature is reached and the temperature decreases for 4 seconds.

16. ☐ Stop recording data. �(6.2)

17. ☐ Name the data run "direct". �(8.2)

18. ☐ Determine the initial temperature (minimum temperature) and the final temperature (maximum temperature) and record the values in Table 1 above. � (9.4)

19. ☐ What is the chemical equation for the process that just occurred in the calorimeter?

$NaOH(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

20. ☐ Remove the temperature sensor from the calorimeter. Rinse and dry the sensor.

21. ☐ Dispose of the solution in the calorimeter according to the teacher's instructions.

22. ☐ Rinse and dry the calorimeter.

23. ☐ Clean the graduated cylinder by rinsing it several times with water.

Part 2 – Indirect Reaction, Step One of Two

24. ☐ Place the clean, dry calorimeter into the 250-mL beaker.

25. ☐ Place the temperature sensor inside the calorimeter.

26. ☐ Measure 50.0 mL of distilled water in a graduated cylinder and pour it into the calorimeter.

27. ☐ Let the water sit until it reaches room temperature.

28. ☐ Record the exact volume of water in Table 2 below.

Table 2: Indirect reaction, step 1, solid sodium hydroxide pellets + water

Volume (mL) of water (H_2O)	50.0
Mass (g) of sodium hydroxide (NaOH)	1.0
Initial temperature (°C)	24.0
Final temperature (°C)	27.6

29. ☐ Why does the volume of water need to be measured carefully?

The volume of water is used to calculate the mass of water, which is used as the mass of the solution in the equation. The volume of water can be converted to mass using the density of water (1.00 g/mL).

30. ☐ Measure 1.0 g of solid sodium hydroxide (NaOH) and set it aside.

31. ☐ Record the exact mass in Table 2 above.

32. ☐ Collect temperature data as the reactants are mixed by following the steps you used earlier in the lab:

a. Start recording data. �(6.2)

b. After 60 seconds add the NaOH pellets.

c. Attach the lid and swirl the apparatus until all the NaOH is dissolved.

d. Continue to record data until a maximum temperature is reached and the temperature decreases for 4 seconds.

e. Stop recording data.�(6.2)

33. ☐ Name the data run "indirect step 1". �(8.2)

34. ☐ Determine the initial temperature (minimum temperature) and the final temperature (maximum temperature) and record the values in Table 2 above. �(9.4)

35. ☐ What is the chemical equation for the process that just occurred in the calorimeter?

$NaOH(s) \rightarrow NaOH(aq)$

36. ☐ Is dissolving sodium hydroxide an exothermic or endothermic process? How do you know?

The process is exothermic because the temperature increased.

37. ☐ Transition to the next stage of the activity by carefully cleaning and drying the temperature sensor and calorimeter.

Part 2 – Indirect Reaction, Step Two of Two

38. ☐ Place the clean, dry calorimeter into the 250-mL beaker.

39. ☐ Place the clean, dry temperature sensor into the calorimeter.

40. ☐ Why is it necessary to have the temperature sensor and calorimeter clean and dry?

It needs to be clean so it does not contaminate the next reaction. It needs to be dry because wetness adds to the exact mass of the solutions.

41. ☐ Measure 25.0 mL of 1.0 M HCl in a graduated cylinder and carefully pour it into the calorimeter.

42. ☐ Allow the acid to stand until it reaches room temperature.

43. ☐ Record the exact volume of HCl used in Table 3 below.

Table 3: Indirect reaction, step 2, sodium hydroxide pellets + 1.0 M hydrochloric acid

Volume (mL) of 1.0 M hydrochloric acid (HCl)	25.0
Volume (mL) of 1.0 M sodium hydroxide (NaOH)	25.0
Initial temperature (°C)	24.6
Final temperature (°C)	31.5

44. ☐ Clean the graduated cylinder by rinsing it several times with water.

45. ☐ Measure 25.0 mL of 1.0 M sodium hydroxide. Set it aside.

46. ☐ Record the exact volume of sodium hydroxide used in Table 3 above.

47. ☐ Collect temperature data as the reactants are mixed by following the steps you used earlier in the lab.

a. Start recording data. �(6.2)

b. After 60 seconds add the NaOH solution.

c. Attach the lid.

d. Continue to record data until a maximum temperature is reached and the temperature decreases for 4 seconds.

e. Stop recording data. �(6.2)

48. ☐ Name the data run "indirect step 2". �(8.2)

49. ☐ Determine the initial temperature (minimum temperature) and the final temperature (maximum temperature) and record the values in Table 3 above. �(9.4)

50. ☐ What is the equation of the chemical reaction that just occurred in the calorimeter?

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

51. ☐ Was this an exothermic or an endothermic process? How do you know?

The process is exothermic because the temperature increased.

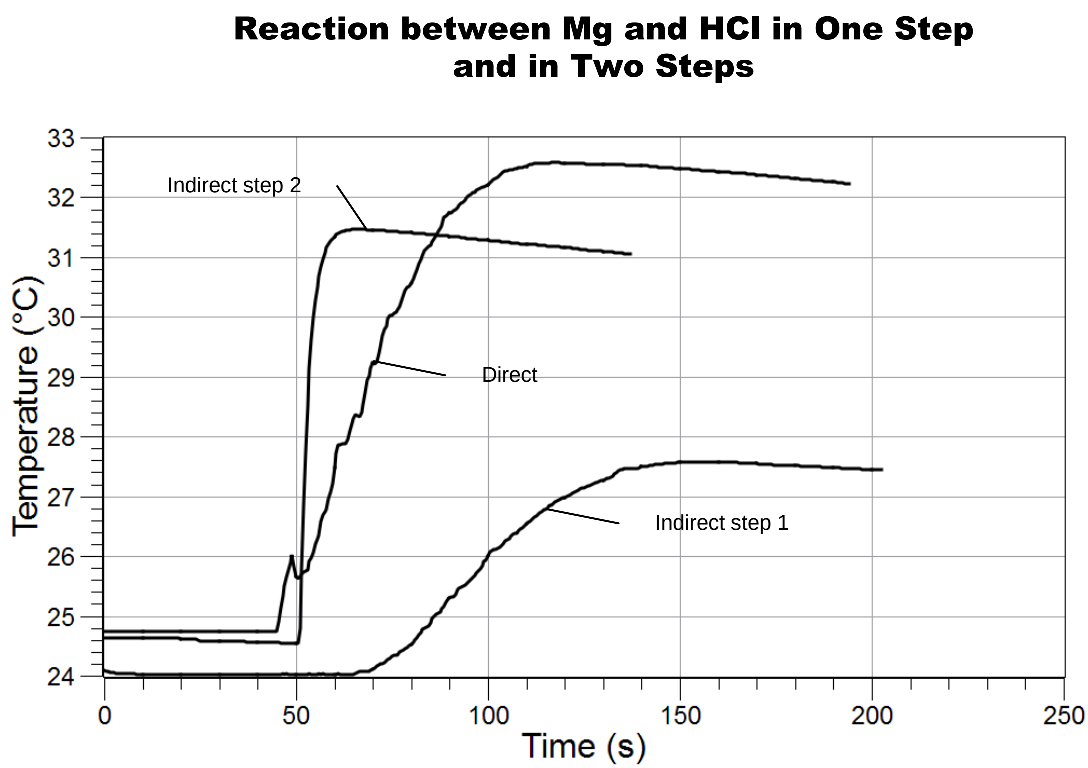
52. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

1. ☐ Create a graph with all three runs of data displayed on your data collection system. �(7.1.3)

Note: Not all data collection systems will display three runs of data. If this is not possible, create one graph for the direct method and a second graph for the indirect method.

2. ☐ Sketch or print a copy of your Temperature (°C) versus Time (s) graph with all three data runs on one set of axes. Label each data run as well as the overall graph, the x-axis, the y-axis, and include units on the axes. �(11.2)



Indirect Method

The steps below guide you through the process of calculating the molar enthalpy of reaction for the reaction between solid sodium hydroxide and aqueous hydrochloric acid. Through this indirect method, you calculate the molar enthalpy of each step and add the two steps together.

3. ☐ Calculate the temperature change for each step in the indirect method. Record your answers in Table 4 below.

Table 4: Indirect reactions, change in temperature

Temperature	**Step 1: Dissolving NaOH**	**Step 2: 1.0 M Sodium Hydroxide + 1.0 M Hydrochloric Acid**
Final temperature (°C)	27.6	31.5
Initial temperature (°C)	24.0	24.6
Change in temperature, ΔT (°C)	3.6	6.9

4. ☐ Calculate the mass of solution for each step. Use the density of water (1.00 g/mL) to convert between volume and mass. Assume solutions have a density equal to that of water.

Table 5: Indirect reactions, total mass of solutions

Step 1: Dissolving NaOH		**Step 2: 1.0 M Sodium Hydroxide + 1.0 M Hydrochloric Acid**	
Mass of H_2O (g)	$(50.0\text{ mL})\left(\frac{1.00\text{ g}}{\text{mL}}\right) = 50.0$	Mass of NaOH (g)	$(25.0\text{ mL})\left(\frac{1.00\text{ g}}{\text{mL}}\right) = 25.0$
Mass of NaOH (g)	1.0	Mass of HCl (g)	$(25.0\text{ mL})\left(\frac{1.00\text{ g}}{\text{mL}}\right) = 25.0$
Total mass of solution (g)	51.0	Total mass of solution (g)	50.0

5. ☐ Calculate the heat q absorbed by the solution in each step using the formula below. Show your work, including the units involved. Convert joules to kilojoules in your final answer.

$q = m \times c \times \Delta T$

q = heat lost or gained by the solution

m = mass of the solution

c = specific heat of water = 4.184 J/(g·°C)

ΔT = change in temperature of the solution

Table 6: Indirect reactions, heat absorbed by the solution

	Step 1: Dissolving NaOH	Step 2: 1.0 M Sodium Hydroxide + 1.0 M Hydrochloric Acid
$q = m \times c \times \Delta T$ Show your work	$(51.0\text{ g})\left(\dfrac{4.184\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)(3.6\,^\circ\text{C})$	$(50.0\text{ g})\left(\dfrac{4.184\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)(6.9\,^\circ\text{C})$
q (J)	770	1400
q (kJ)	0.77	1.4

6. ☐ Find the molar enthalpy of reaction ΔH for each step. The amount of heat absorbed by the solution is the opposite of the amount of heat generated by the change.

$\Delta H = -q$

moles = (mass)/(molar mass) or moles = (volume)(molarity)

$$\text{molar enthalpy} = \frac{\Delta H}{\text{moles of substance}}$$

Table 7: Indirect reactions, molar enthalpy of reaction

	Step 1: Dissolving NaOH	Step 2: 1.0 M Sodium Hydroxide + 1.0 M Hydrochloric Acid
ΔH (kJ)	–0.77 kJ	–1.4 kJ
Moles of NaOH	$(1.0\text{ g})\left(\dfrac{1\text{ mol}}{40.0\text{ g}}\right) = 0.025\text{ mol}$	$(25\text{ mL})\left(\dfrac{1.0\text{ L}}{1000\text{ mL}}\right)\left(\dfrac{1.0\text{ mol}}{1.0\text{ L}}\right) = 0.025\text{ mol}$
ΔH/mol (kJ/mol)	$\dfrac{-0.77\text{ kJ}}{0.025\text{ mol}} = -31$	$\dfrac{-1.4\text{ kJ}}{0.025\text{ mol}} = -56$

7. ☐ Calculate the molar enthalpy for the overall reaction between solid sodium hydroxide and aqueous hydrochloric acid by adding the molar enthalpy of reaction for each step.

$\Delta H_{rxn} = \Delta H_{step\ 1} + \Delta H_{step\ 2}$

$\Delta H_{rxn} = -31\text{ kJ/mol} + -56\text{ kJ/mol}$

$\Delta H_{rxn} = -87\text{ kJ/mol NaOH}$

Direct Method

8. ☐ Calculate the molar enthalpy for the reaction between solid sodium hydroxide and aqueous hydrochloric acid. Use the temperature data you collected by directly reacting the two reactants. Use the steps for the indirect method in this section to guide you if you get stuck. Record your answer to the correct number of significant figures.

Table 8: Direct reaction, molar enthalpy for the reaction between NaOH and HCl

Heat of Reaction Calculations	**Direct Reaction: Solid Sodium Hydroxide + 0.50 M Hydrochloric Acid**	
	Show your work in this column.	**Record your answer here.**
Change in Temperature, ΔT (°C)	32.8°C – 24.0°C	8.8
Total mass of solution, solvent + solute (g)	$(50.0\ \text{mL})\left(\frac{1.00\ \text{g}}{\text{mL}}\right) = 50.0\ \text{g}$ $50.0\ \text{g} + 1.0\ \text{g} = 51.0\ \text{g}$	51.0
q (J)	$(51.0\ \text{g})\left(\frac{4.184\ \text{J}}{\text{g} \cdot {}^\circ\text{C}}\right)(8.8\ {}^\circ\text{C}) = 1877.78\ \text{J}$	1900
q (kJ)	$(1877.78\ \text{J})\left(\frac{1\ \text{kJ}}{1000\ \text{J}}\right) = 1.87778\ \text{kJ}$	1.9
ΔH (kJ)	$\Delta H = -q = -1.87778$ kJ	–1.9
moles of NaOH	$(1.0\ \text{g})\left(\frac{1\ \text{mol}}{40.0\ \text{g}}\right) = 0.025\ \text{mol NaOH}$	0.025
ΔH mol (kJ/mol)	$\frac{-1.87778\ \text{kJ}}{0.025\ \text{mol}} = -75.1112\ \text{kJ/mol}$	–75

Analysis Questions

1. How do the enthalpies of reaction for the indirect method compare with the direct method? Are they the same?

Answers will vary. In the sample data, the enthalpy using the indirect method was determined to be –87 kJ/mol. Using the direct method it was –75 kJ/mol. For the level of precision of the method, it can be said the values were the same.

2. What is the percent error for the results produced in this experiment? (In this case, let the actual yield equal the enthalpy calculated indirectly and the theoretical yield equal the enthalpy calculated directly.)

$$\text{percent error} = \left|\frac{\text{accepted value} - \text{experimental value}}{\text{accepted value}}\right| \times 100$$

$$\text{percent error} = \left|\frac{(-75\ \text{kJ/mol}) - (-87\ \text{kJ/mol})}{-75\ \text{kJ/mol}}\right| \times 100 = 16\%$$

Synthesis Questions

Use available resources to help you answer the following questions.

1. In your own words, what is Hess's law?

Hess's law states that enthalpies of reactions are additive. Any set of reactions that ultimately begins and ends with the same products and reactants can be used indirectly to determine the enthalpy of a cumulative reaction.

2. In what circumstances would Hess's law be useful?

Hess's law is useful when the heat of a reaction cannot be determined directly: when the reaction is too slow, when the reaction is too fast (explosive), or when the materials are toxic, expensive, or simply not available.

3. What is the heat of reaction for the formation of "water gas" (a mixture of CO and H_2 gases)? Is this reaction endothermic or exothermic? Use the note and the data in Table 9 below to answer this question.

> **Note:** The following formula shows the formation of water gas, which is a mixture of CO and H_2 gases: $C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)$.

Table 9: Calculations for three reactions

Reaction 1:	**$C(s) + O_2(g) \rightarrow CO_2(g)$**	**$\Delta H = -394$ kJ/mol**
Reaction 2:	**$CO_2(g) \rightarrow CO(g) + ½\ O_2(g)$**	**$\Delta H = 283$ kJ/mol**
Reaction 3:	**$H_2O(g) \rightarrow H_2(g) + ½\ O_2(g)$**	**$\Delta H = 242$ kJ/mol**

$\Delta H_{rxn} = \Delta H_{rxn1} + \Delta H_{rxn2} + \Delta H_{rxn3}$

$\Delta H_{rxn} = (-394\ \text{kJ/mol}) + (283\ \text{kJ/mol}) + (242\ \text{kJ/mol}) = 131\ \text{kJ/mol}$

Because the heat of reaction, 131 kJ/mol, is positive, it is an endothermic reaction. It requires an input of energy to form "water gas" using this reaction.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. What data does NOT need to be collected in order to determine the enthalpy of a chemical reaction?

- **A.** Mass of the solution
- **B**. Change in temperature of the solution
- **C.** The specific heat capacity of the solution
- **D. Volume of the solution**

2. How does the enthalpy of the overall reaction compare with the sum of the enthalpies of the intermediate steps?

- **A.** Greater than
- **B**. Less than
- **C. Equal to**
- **D.** Either A or B depending on the reaction

3. How can the enthalpy of a particular reaction be determined?

- **A.** By measuring the change in temperature of the reaction directly
- **B**. By measuring the change in temperature of a series of reactions that, when added together, give the particular reaction
- **C. Both A and B are possible depending on the reaction**
- **D.** Enthalpies of reactions cannot be determined experimentally

4. What circumstances may require the use of Hess's law to find the enthalpy of a reaction?

- **A.** Reactions that take millions of years
- **B**. Reactions involving toxic chemicals
- **C.** Reactions that are explosive
- **D. All of the above**

5. What does Hess's law state?

- **A. The enthalpy change of the overall reaction is equal to the sum of the enthalpy changes in the individual steps**
- **B**. The enthalpy change of the overall reaction is equal to the product of the enthalpy changes in the individual steps
- **C.** The enthalpy change of the overall reaction is greater than the sum of the enthalpy changes in the individual steps
- **D.** The enthalpy change of the overall reaction is less than the product of the enthalpy changes in the individual steps

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Chemical and physical processes often involve changes in **heat**. Processes that produce heat are **exothermic** and those that absorb heat are **endothermic**. Changes in **temperature** can be measured by performing the chemical reaction or physical process in an insulated container called a **calorimeter**. The heat of the process is then calculated by multiplying the **mass**, the **specific heat capacity**, and change in temperature ΔT together ($q = m \times c \times \Delta T$).

2. **Enthalpy** is the term chemists use to express the heat content of a system at constant pressure. The enthalpy of some chemical reactions can be calculated directly by mixing the reactants and measuring the change in **temperature**. Some chemical reactions take too long (millions of years) or occur too **quickly** to get an accurate temperature reading. Other chemical reactions may involve toxic chemicals or chemicals that are too expensive or rare to be obtained. In these situations an **indirect method** using a series of chemical reactions that **add** together to give the overall reaction can be used in accordance with Hess's law. **Hess's law** states that the enthalpy change of the overall reaction is equal to the sum of the enthalpy changes in the **individual steps** that make up the **overall reaction**.

Extended Inquiry Suggestions

Use Hess's law to experimentally determine the enthalpy of combustion of magnesium.

Use thermodynamic tables to determine the theoretical value for the molar enthalpy of the reaction between sodium hydroxide and hydrochloric acid.

26. An Acid-Base Titration

Objectives

Use a titration to determine the concentration of a hydrochloric acid solution and the concentration of an acetic acid solution. Through this investigation, students:

- Differentiate between concentration and strength of acids and bases
- Perform neutralization reactions
- Describe and explain the shape of a titration curve

Procedural Overview

Students conduct the following procedures:

- Perform a strong acid–strong base titration
- Perform a weak acid–strong base titration
- Calculate the concentration of an unknown hydrochloric acid solution and an unknown acetic acid solution using collected data and stoichiometric calculations

Time Requirement

♦ Preparation time	20 minutes
♦ Pre-lab discussion and activity	45 minutes
♦ Lab activity	60 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Drop counter
- pH sensor
- Magnetic stirrer
- Micro stir bar
- Beaker (2), 250-mL
- Beaker (2), 50-mL
- Graduated cylinder, 100-mL
- Volumetric pipet or graduated cylinder, 10-mL
- Buret , 50-mL
- Ring stand
- Right-angle clamp
- Buret clamp
- Funnel
- Transfer pipet
- Waste container
- Wash bottle filled with distilled (deionized) water
- Buffer solution pH 4, 25 mL
- Buffer solution pH 10, 25 mL
- Distilled (deionized) water, 200 mL
- Hydrochloric acid (HCl) solution, 10 mL[1]
- Acetic acid ($HC_2H_3O_2$) solution, 10 mL[2]
- Standardized sodium hydroxide (NaOH) solution, 120 mL[3]

[1]The hydrochloric acid (HCl) solution should be approximately 0.1 M. To formulate this solution using concentrated (12 M) or dilute (6 M) HCl, refer to the Lab Preparation section.

[2]The acetic acid ($HC_2H_3O_2$) solution should be approximately 0.1 M. To formulate this solution using concentrated (glacial, 17.4 M) acetic acid or 6 M acetic acid refer to the Lab Preparation section.

[3]The standardized sodium hydroxide (NaOH) solution should be approximately 0.1 M. Refer to the Lab Preparation section on how to create and standardize the solution.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Balanced chemical equations
- Concentration (molarity)
- Molar ratios and stoichiometric calculations
- Properties of acids and bases
- pH

Related Labs in This Guide

Labs conceptually related to this one include:

- pH of Household Chemicals
- Diprotic Titration: Multi-Step Chemical Reactions
- Double Replacement Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connect a sensor to the data collection system. �(2.1)
- Connecting multiple sensors to the data collection system �(2.2)
- Calibrating a pH sensor �(3.6)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Showing and hiding data runs in a graph �(7.1.7)
- Changing the variable on the x-axis of a graph �(7.1.9)
- Naming a data run �(8.2)
- Finding the coordinates of a point in a graph �(9.1)
- Finding the slope at a point on a data plot �(9.3)
- Creating calculated data �(10.3)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

Titration is a common quantitative laboratory method used to determine the concentration of a reactant. A reagent of known concentration, called the titrant, is used to react with a measured volume of the other reactant, called the analyte. Because volume measurements of the analyte and titrant are key factors in this type of analysis, it is also known as a volumetric analysis.

When a basic solution is added to an acidic solution of unknown concentration, hydroxide ions from the basic solution react with hydrogen ions from the acidic solution to form neutral water and a salt. The type of salt formed depends on the acid and base used. This type of reaction is called a neutralization reaction.

acid + base → water + salt

$HY(aq) + XOH(aq) \rightarrow H_2O(l) + XY(aq)$

$HBr(aq) + KOH(aq) \rightarrow H_2O(l) + KBr(aq)$

Since pH is a measure of the concentration of hydronium ions (H_3O^+), the pH will change as more hydroxide ions (OH^-) are added. The point at which the number of moles of hydroxide ions added is equal to the number of moles of H_3O^+ ions present is called the equivalence point. The equivalence point can be determined from the point on a titration curve with the greatest slope.

In this lab, students will perform two titrations. One titration involves a strong acid and the other involves a weak acid. It is important to ensure the students understand that there is a difference between the strength of an acid and the concentration of an acid. The concentration of an acid is simply the number of moles of acid per liter of solution. The strength of an acid, however, refers to the amount of dissociation that occurs when the acid is placed in water. Two acids may have the same concentration, for example 1 M, but the strong acid will have a lower pH because more of its molecules dissociate into ions in the presence of water.

Students start by titrating a strong acid (hydrochloric acid) with a strong base (sodium hydroxide) forming water and sodium chloride (common table salt).

$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$

Second, students titrate a weak acid (acetic acid) with a strong base (sodium hydroxide) forming water and sodium acetate (a salt).

$HC_2H_3O_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$

The titration curves obtained (pH versus NaOH volume) show different initial pH values between the strong and the weak acids as well as different pH ranges at their equivalence points. Students use the volume of NaOH solution at the equivalence point to determine the number of moles of NaOH needed to neutralize the acid. Using the mole-to-mole ratio between NaOH and acid (in both reactions it is 1:1), students then determine the number of moles of acid. By dividing the moles of acid by the starting volume of acid, students are able to calculate the acid's original concentration.

Pre-Lab Discussion and Activity

Fish Tank and an Unknown Concentration of HCl

Hold up a flask partially filled with an unknown concentration of hydrochloric acid (HCl). The flask should be labeled HCl. Explain to the students that the biology teacher accidently poured half the contents of the flask into the classroom fish tank and wanted to know if the fish in the tank would survive. Using this scenario, discuss acids, bases, pH, and concentration. Ultimately the fate of the fish will depend on the concentration of the HCl added and the degree to which it lowers the pH in the fish tank. Have the students propose suggestions on ways to ensure the safety of the fish.

1. Will the fish survive even though hydrochloric acid was added to the fish tank?

The fish may or may not survive; it depends on the overall pH in the fish tank. The pH in the fish tank will depend on the concentration of the HCl that was added. The student discussion may involve acid rain and whether or not fish can survive in acidic aquatic conditions. The students should conclude that fish can live in slightly acidic conditions, but will die if the acid content surpasses a certain point.

2. What do you know about acids? Is HCl a strong acid or a weak acid? Does it matter?

Acids are substances that donate H^+ ions causing the H_3O^+ ion concentration to increase when dissolved in water. HCl is a strong acid. This is important because strong acids dissociate completely, which produces a greater number of H_3O^+ ions and therefore a lower pH.

3. What is pH and how does it relate to acids and fish?

pH is a measure of the H_3O^+ ion concentration. Low pH values mean there is a high concentration of H_3O^+ ions. Fish, like most living things, can only survive in a given range of pH. Most fish can survive when the pH of the water is between 6 and 8. Fish tend to die at extreme values (pH greater than 10 or less than 4).

4. Is there something we can do to counteract the effect of the acid and save the fish? Why would your suggestion work?

One option would be to add a base. Bases are substances that accept H^+ ions. They reduce the number of H_3O^+ ions in solution and cause the pH to increase. Adding a base would neutralize the acid and keep the pH within values that are viable for fish to survive. The amount of base added is important. If too much base is added, the pH could rise to a level just as unhealthy for the fish as an excessively low pH caused by the addition of an acid.

While adding base would neutralize the acid, every neutralization reaction always produces a salt. Both freshwater and saltwater fish have limits to the amount of salt they can withstand. Another option is to dilute the acid by adding freshwater to the tank. Adding water will decrease the concentration of the acid (same number of moles in a larger volume of water) to the level that the fish may survive, but it will never neutralize it completely.

5. How could we determine the H_3O^+ ion concentration in this (hold up the flask) HCl solution? How would knowing the concentration help us?

The concentration of ions could be determined from the pH of the solution or by performing a titration (the quantitative lab technique introduced in this lab). Knowing the concentration of acid would allow us to determine if we should add a base to neutralize the solution, and, if so, how much we need to add.

Neutralization Reaction

Engage students in a discussion of safety measures when handling and mixing strong acids and bases. Explain that the properties of acids and bases are canceled out, or neutralized, when the two are mixed. End the discussion by asking students to explain a neutralization reaction at the molecular level. Then have them propose ideas for methods for learning whether the H_3O^+ ions and OH^- ions are present in equal amounts, even though they cannot be seen.

6. What happens when concentrated HCl acid touches a person's skin? How about concentrated sodium hydroxide (NaOH)?

Both concentrated acids and bases are corrosive, causing severe burns to skin and damage to clothing.

7. What would happen if you mixed concentrated HCl and NaOH? What products would form? Would they be safe to touch?

$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$

Mixing concentrated solutions together will produce a great amount of heat, possibly enough to boil the solution. Reacting HCl with NaOH produces water and sodium chloride (common table salt). Table salt and water are both safe compounds that are consumed on a daily basis and are safe to touch.

8. What is a neutralization reaction?

A neutralization reaction occurs when an acid and a base react to form a salt and water.

9. Describe a neutral solution at the molecular level.

A neutral solution has equal amounts of H_3O^+ ions and OH^- ions.

10. When NaOH is added to HCl, how will we know when the solution is neutral?

Since we cannot see individual H_3O^+ and OH^- ions, we must add a colored substance (indicator) or use technology (pH sensor) to help us determine when the concentrations are equal.

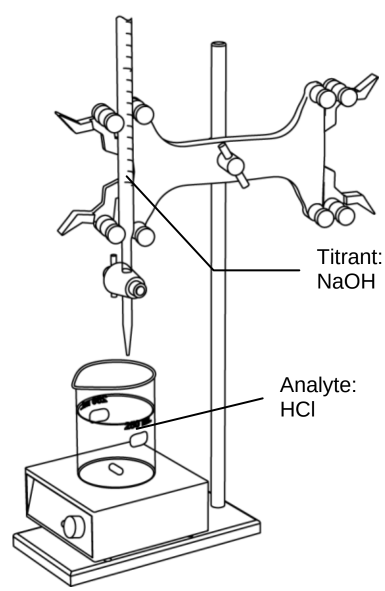
Titration Demonstration Using Phenolphthalein

To perform this demonstration using phenolphthalein, as well as the additional demonstration using the pH sensor, prepare a few solutions ahead of time. Be sure to follow the precautions listed in the Lab Preparation section below.

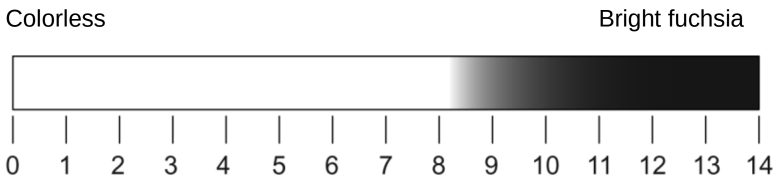
Prepare a solution of sodium hydroxide with an approximate concentration of 1 M: Dissolve 4.0 g of sodium hydroxide in 100 mL of distilled water.

Prepare a solution of hydrochloric acid with an approximate concentration of 0.5 M: Dilute either 4 mL of 12 M or 8 mL of 6 M HCl to a final volume of 100 mL.

Obtain a 1% solution of phenolphthalein in 95% ethanol from a supplier.



Demonstrate a traditional acid-base titration: Pour 50.0 mL of the HCl solution with the undisclosed concentration into a beaker. Add a stir bar and 2 to 3 drops of phenolphthalein indicator. Fill the buret with the standardized NaOH solution with the known concentration to the buret. Record the concentration of the standard NaOH solution as well as the initial volume. Open the stopcock and allow the liquid to start dripping into the acid analyte. Explain to the students that phenolphthalein is colorless in acidic conditions and turns bright fuchsia in basic conditions. This color change occurs near a pH of 8.



Traditional titration apparatus.

Color of phenolphthalein indicator in solutions with the corresponding pH values.

11. What is a titration and why is it used?

A titration is a common quantitative laboratory method used to determine the concentration of a reactant. A reagent of known concentration, called the titrant, is used to react with a measured volume of the reactant, called the analyte. The volumes of the titrant and analyte are carefully recorded and used with the concentration of the titrant to determine the unknown concentration of the analyte.

12. What is the equivalence point in a titration? Why is it useful?

The equivalence point of a titration is the point at which enough of the standard solution has been added so that the concentrations of the H_3O^+ and OH^- ions are equal. The volume of standard solution, the concentration of the standard solution, the mole ratio from the balanced chemical equation, and the volume of the analyte can then be used to calculate the concentration of the analyte.

13. Calculate the concentration of our unknown acid solution.

Sample data for a titration

Concentration of the Standard NaOH Solution (M)	1.00
Starting Volume of HCl Solution (mL)	50.00
Initial Volume of NaOH in the Buret (mL)	1.20
Final Volume of NaOH in the Buret (mL)	22.70

Sample calculations to determine concentration

Total Volume of NaOH Used (M)	Final volume NaOH – Initial volume NaOH = Total NaOH volume 22.70 mL – 1.20 mL = 21.50 mL
Moles of NaOH at the Equivalence Point (mol)	$\left(0.02150\ \text{L NaOH}\right)\left(\frac{1.00\ \text{mol NaOH}}{1\ \text{L NaOH}}\right) = 0.02150\ \text{mol NaOH}$
Balanced Chemical Equation	$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
Moles of Acid in Solution (mol)	$\left(0.02150\ \text{mol NaOH}\right)\left(\frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}}\right) = 0.02150\ \text{mol HCl}$
Concentration of Acid Solution (mol/L)	$\frac{0.02150\ \text{mol HCl}}{0.05000\ \text{L HCl}} = 0.4300\ \text{M HCl}$

14. What is an indicator and why is it used?

An indicator is a chemical substance that has a distinct color change at different pH values. Indicators are used to help chemists “see” changes taking place at the molecular level. In titrations, indicators are used to identify the equivalence point (end point) of the titration.

Note: In a traditional colorimetric titration, the term endpoint is sometimes used instead of equivalence point. The two terms are similar, but not the same. The endpoint is the point at which the solutions changes color and signals the end of the titration.

15. What are some limitations of using indicators in titrations?

It is hard to know exactly when the pink color has been achieved. It is easy to add too much of the standard solution and pass beyond the end point. When this happens, the entire titration must be repeated because the end point has been missed and the volume of standard solution added is incorrect. Indicators cannot be used with colored substances or with reactions that proceed slowly.

Titration Demonstration Using pH Sensor

Many of the limitations in a titration using an indicator can be eliminated by using a drop counter and a pH sensor. (Refer to the setup picture in the Procedure section.) Repeat the titration using a drop counter and pH sensor instead of pheonolphthalein. Display the data being collected so that all students can see it. Once the titration curve is complete, help students determine the equivalence point and analyze the shape of the titration curve.

16. What is a drop counter and why is it used in a titration? How does it overcome some of the limitations of indicator-based titrations?

A drop counter is exactly what its name implies; it counts the number of drops that pass through it. The volume of each drop can be determined by dividing the volume of titrant used by the total number of drops used to dispense that volume. A calculation can then be used to convert the number of drops at the equivalence point of a titration curve to the volume of titrant used.

This procedure eliminates the need for stopping the titration exactly at the equivalence point, which can be tricky when using an indicator. The procedure can also be completed for titrations that cannot use indicators.

17. Where is the equivalence point on a titration curve?

The equivalence point is where the slope of the titration curve is the steepest. For strong acid-strong base titrations, this occurs at a pH equal to 7. Remember, a pH of 7 means the H_3O^+ and OH^- concentrations are equal.

18. What does the slope of the titration curve represent?

The slope of the titration curve represents the change of pH per drop of NaOH (standard solution).

19. Describe how the slope of the titration curve changes as the number of drops (volume) is increased.

The slope of titration curve is very gradual at first. There is a very small change in pH with each drop of NaOH added. Quite suddenly, the slope increases drastically, indicating there is a large change in pH with each drop of NaOH added. The slope then decreases again and there is hardly any change in the pH with each drop of NaOH added.

20. Explain why there are such drastic changes in the slope of a titration curve.

The drastic changes in slope are due to the logarithmic nature of the pH scale. A logarithmic scale covers a very large range of values (in this case a large range of H_3O^+ ion concentration). At low pH values, there are a large number of H_3O^+ ions and so a drop containing a fixed number of OH^- ions has little affect on the total number of H_3O^+ ions. At higher pH values, there are fewer H_3O^+ ions. When the same size drop, containing the same fixed number of OH^- ions, is added, it now makes a much larger difference in pH.

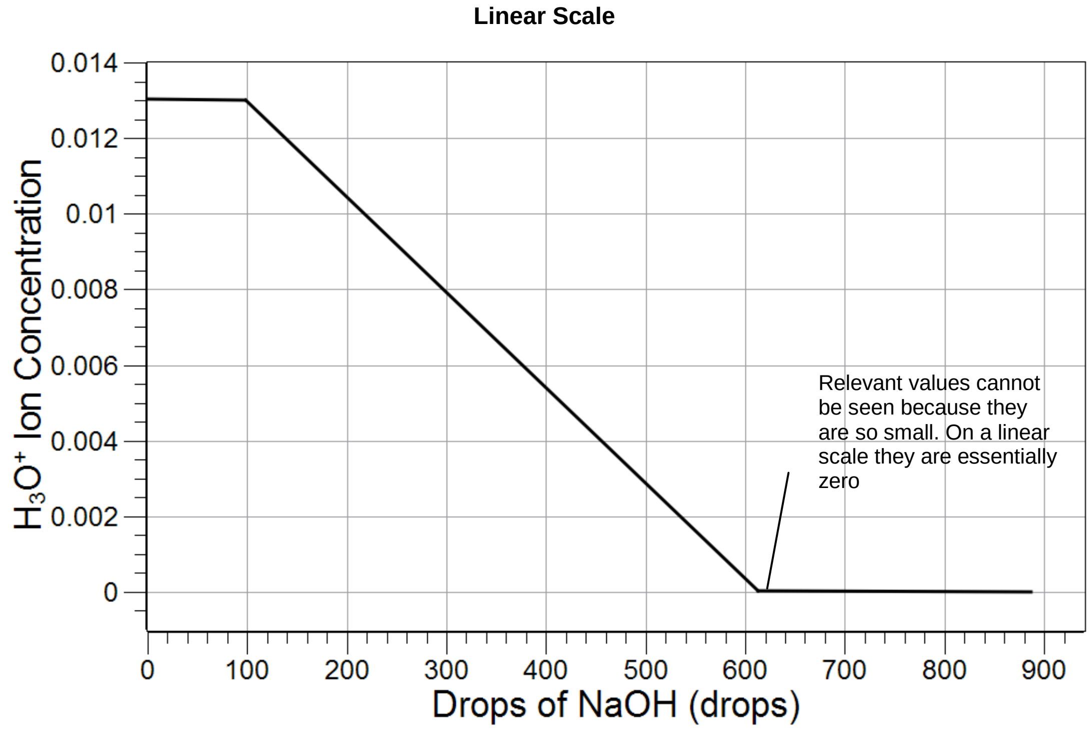
Cell Phone Minutes Analogy

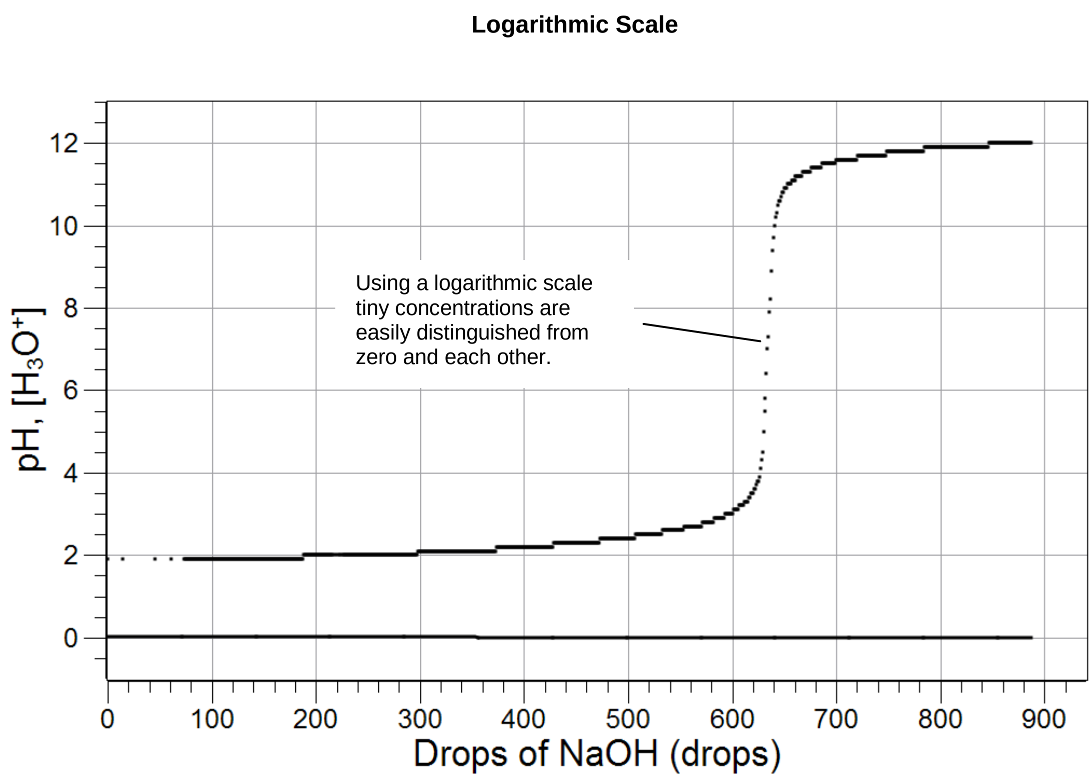
Use the following analogy to help your students understand the drastic changes in slope of a titration curve:

Think of the way many cell phone companies charge their customers. Typically the customer picks a program that includes a given number of minutes per month for a flat rate. If the customer talks for more minutes than their plan allows, they are charged extra for each minute they talk. For example, you decide to buy a cell phone plan that allows you to use 1000 minutes a month. For the sake of this analogy, assume each phone call you make lasts 15 minutes. At the beginning of the month making one 15-minute call hardly changes the number of minutes you have left because you have such a large number of minutes ($15 \div 1000 \times 100 = 1.5\%$). Toward the end of the month, however, one phone call makes a larger impact. If you have 25 minutes left on your plan and you make one 15-minute phone call, that suddenly cuts your minutes by more than half ($15 \div 25 \times 100 = 60\%$).

In a titration, the same thing is happening. Each added drop of NaOH solution uses up (neutralizes) the same amount of H_3O^+ just like one phone call uses up the same amount of minutes. The overall effect of one drop of NaOH depends on the "bank" of H_3O^+ ions it is being added to.

Use the following graphs to illustrate the need for the logarithmic pH scale. The first graph is a linear representation of H_3O^+ ion concentration versus drops of NaOH and the second graph is a logarithmic scale (pH versus drops of NaOH).





21. Explain what is happening to the H_3O^+ ion concentration as drops of NaOH are added.

The concentration of H_3O^+ ions decreases as drops of NaOH are added.

22. Why does the H_3O^+ ion concentration appear to level out as more and more drops of NaOH are added? Why is this misleading?

The H_3O^+ ion concentration is so small that it appears to be zero on this scale. This is misleading because, as drops are added, the concentration of H_3O^+ ions is not actually zero and continues to decrease.

23. Explain how a logarithmic scale makes it much easier to distinguish what is happening at very small concentrations of H_3O^+ ions.

The logarithmic scale allows for a greater range of values making it easy to compare very tiny numbers, such as when the H_3O^+ number of ions and OH^- ions become equal.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids creates heat. Be extra careful when handling freshly prepared solutions and glassware because they might be very hot.
- ♦ Always add acids to water, not the other way around, because the solutions may boil vigorously.
- ♦ Handle concentrated acids in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Prepare 1000 mL of 0.1 M hydrochloric acid from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 100 lab groups.

Starting with concentrated (12 M) HCl:

a. Place a stir bar in a 1000-mL beaker, and add approximately 500 mL of distilled water.

b. Slowly add 8.3 mL of 12 M HCl to the beaker, stirring continuously.

c. Allow the solution to cool. Then carefully pour the solution into a 1000-mL volumetric flask, and dilute to the mark with distilled water.

d. Cap and invert three times carefully to ensure complete mixing.

Starting with dilute (6 M) HCl

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 16.7 mL of 6 M HCl to the water, and dilute to the mark with distilled water.

c. Cap and invert three times carefully to ensure complete mixing.

2. Prepare 1000 mL of 0.1 M acetic acid from either concentrated (glacial, 17.4 M) or dilute (6 M) acetic acid. This is enough for 100 lab groups.

Starting with concentrated (glacial, 17.4 M) acetic acid:

- **a.** Place a stir bar in a 1000-mL beaker, and add approximately 500 mL of distilled water.
- **b.** Slowly add 5.7 mL of 17.4 M acetic acid to the beaker, stirring continuously.
- **c.** Allow the solution to cool. Then carefully pour the solution into a 1000-mL volumetric flask, and dilute to the mark with distilled water.
- **d.** Cap and invert three times carefully to ensure complete mixing.

Starting with dilute (6 M) acetic acid:

- **a.** Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.
- **b.** Add 16.7 mL of 6 M acetic acid to the water, and dilute to the mark with distilled water.
- **c.** Cap and invert three times carefully to ensure complete mixing.

3. Prepare 1000 mL of 0.1 M sodium hydroxide (NaOH). This is enough for 8 lab groups.

- **a.** Place a stir bar in a 1000-mL beaker and add approximately 500 mL of distilled water.
- **b.** Slowly add 4.0 g of solid NaOH to the beaker; allow it to dissolve completely while stirring continuously.
- **c.** Allow the solution to cool, then carefully pour the solution into a 1000-mL volumetric flask and dilute to the mark with distilled water.
- **d.** Cap and invert three times carefully to ensure complete mixing.
- **e.** The exact concentration of the NaOH solution will vary due to the hygroscopic (attracting water from the surrounding environment) nature of solid NaOH. The NaOH solution will absorb carbon dioxide from the air; this will produce a small amount of carbonic acid in the solution which will neutralize a small portion of the NaOH. For best results, accurately determine the concentration of (standardize) the 0.1 M NaOH solution by titrating it with potassium bitartrate (cream of tartar) as described in the next step. Report to the students the actual molarity of the NaOH solution.

4. To standardize NaOH using potassium bitartrate ($KHC_4H_4O_6$) follow the steps below.

- **a.** Using a 100-mL beaker, dissolve approximately 0.5 g of potassium bitartrate (record actual amount) in exactly 50 mL of distilled water.
- **b.** Using a calibrated pH sensor and a drop counter, find the volume of 0.1 M NaOH solution that is required to neutralize the potassium bitartrate solution.
- **c.** Use the following calculation to find the actual concentration of the NaOH solution:

$$\left(\text{grams KHC}_4\text{H}_4\text{O}_6\right)\left(\frac{1\ \text{mol}}{188.177\ \text{g KHC}_4\text{H}_4\text{O}_6}\right)\left(\frac{1\ \text{mol NaOH}}{1\ \text{mol KHC}_4\text{H}_4\text{O}_6}\right)\left(\frac{1}{\text{L NaOH added}}\right)$$

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Sodium hydroxide, hydrochloric acid, and acetic acid are corrosive irritants. Avoid contact with the eyes and wash hands after handling.
- ♦ Be sure that all acids and bases are neutralized before disposal down the drain.
- ♦ When mixing acids with water, always add the acid to the water, not the other way around, as the solutions may get hot enough to boil.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	3	5	4
Place 100 mL of distilled water and 10 mL of HCl into a 250-mL beaker and place it on the magnetic stirrer with the pH sensor in place.	Assemble the titration apparatus, rinse and then fill the buret with NaOH solution, and calibrate the pH sensor.	Write down the exact starting volumes of HCl and NaOH as well as the exact concentration of the NaOH solution. Start recording data.	Thoroughly clean the equipment and then repeat the procedure using the acetic acid solution.	Allow the NaOH to drip into the beaker until a pH of 12. Close the stopcock and stop recording data. Record the final volume of NaOH.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a pH sensor to the data collection system. �(2.1)

3. ☐ Place 25 mL of pH 4 buffer solution in one 50-mL beaker and 25 mL of pH 10 buffer solutions in a second 50-mL beaker. Use these solutions to calibrate the pH sensor. �(3.6)

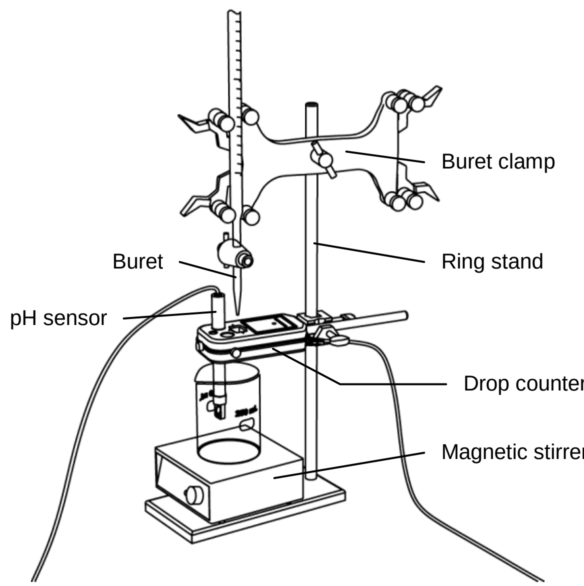
4. ☐ Using the terms "accuracy" and "precision," explain why is it necessary to calibrate the pH sensor?

A pH sensor needs to be calibrated to ensure accurate results. A pH sensor that is not calibrated will give precise results, but they may not be accurate.

5. ☐ Connect a drop counter to the data collection system. �(2.2)

6. ☐ Display pH versus Drop Count (drops) on a graph. �(7.1.1)

7. ☐ Assemble the titration apparatus, using the steps below and the illustration as a guide.



- **a.** Assemble the ring stand.
- **b.** Position the magnetic stirrer on (or next to) the base of the ring stand.
- **c.** Place a waste container on the magnetic stirrer.
- **d.** Use the buret clamp to attach the buret to the ring stand.
- **e.** Position the drop counter over the waste container and attach it to the ring stand using the right-angle clamp.
- **f.** Place the pH sensor through one of the slots in the drop counter.

8. ☐ Rinse the buret with several milliliters of the standardized NaOH solution. Follow the steps below to complete this step.

- **a.** Ensure that the stopcock is closed and use a transfer pipet to rinse the inside of the buret with several milliliters of the standardized NaOH solution.
- **b.** Open the stopcock on the buret and drain the rinse NaOH into the waste container.
- **c.** Repeat this process two more times.

9. ☐ Why is it necessary to rinse the buret with the standardized NaOH solution?

If there is any residual water or containment in the buret, it will dilute the NaOH and change its concentration. Rinsing eliminates any such contamination.

10. ☐ Make sure the stopcock on the buret is in the "off" position, and then use a funnel to fill the buret with about 50 mL of the standardized NaOH solution (titrant).

11. ☐ Drain a small amount of the titrant through the drop counter into the waste beaker to remove any air in the tip of the buret.

12. ☐ Why is it important to remove air from the tip of the buret?

Any air trapped in the buret tip is counted as volume of NaOH. If this happens the final amount of titrant used will be inaccurate.

13. ☐ Practice adjusting the stopcock on the buret so that the titrant goes through the drop counter in distinguishable drops that fall at about 2 to 3 drops per second.

Note: Good control of the stopcock is important. If you accidently open the stopcock too far and the NaOH flows out (as opposed to drops out), you will have to start over.

14. ☐ Why will it be necessary to start your titration over again if you accidently allow the titrant to flow out of the stopcock instead of drop?

The drop counter counts distinct drops. If the drops are not sufficiently distinct from one another, the drop counter will not function properly and the fluid volume will not be accurate.

15. ☐ Close the stopcock and then remove the waste container.

Data Collection

Table 1: Titration data

Measurement	Hydrochloric Acid (HCl) Trial	Acetic Acid ($HC_2H_3O_2$) Trial
Concentration of the standard NaOH solution (M)	0.098	0.098
Volume of acid used (mL)	10.0	10.0
Initial volume of NaOH in the buret (mL)	0.21	0.55
Final volume of NaOH in the buret (mL)	29.46	17.55

16. ☐ Record the concentration of the standardized NaOH solution in Table 1 above (HCl Trial).

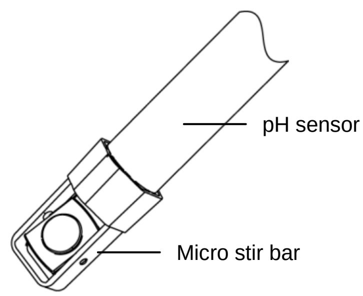
17. ☐ Using the 100-mL graduated cylinder, measure 100.0 mL of distilled water and add it to a 250-mL beaker.

18. ☐ Using the 10-mL graduated cylinder or a 10-mL volumetric pipet, measure 10.0 mL of HCl, and add it to the 100.0 mL of distilled water.

Caution: Always add acid to water.

19. ☐ Record the exact volume of acid used in Table 1 above (HCl Trial).

20. ☐ Add the micro stir bar to the end of the pH sensor.



21. ☐ Position the 250-mL beaker on the magnetic stirrer with the pH sensor submerged in the acidic solution.

22. ☐ Ensure the bulb of the pH sensor is fully submerged, and then turn on the magnetic stirrer and begin stirring at a slow-to-medium speed.

23. ☐ Why is it necessary to stir the solution during a titration?

Stirring thoroughly mixes the ions in the solution so that the recorded pH is for the entire solution.

24. ☐ Determine the initial volume of the titrant (NaOH solution) in the buret to a precision of 0.01 mL, and record this in the HCl column in Table 1 above.

25. ☐ Start recording data. �(6.2)

26. ☐ Carefully open the stopcock on the buret so that 2 to 3 drops per second are released.

27. ☐ Continue recording data until the pH value reaches 12. If needed, rescale the axes so that you can see the changes taking place. �(7.1.2)

28. ☐ What substances are being formed in the beaker? What type of reaction is occurring?

A salt (sodium chloride, NaCl) and water are being formed. The reaction is a neutralization reaction.

29. ☐ Write the chemical reaction that is occurring in the beaker.

$NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$

30. ☐ Close the stopcock when the pH of the solution reaches 12.

31. ☐ Stop recording data. �(6.2)

32. ☐ Name the data run "HCl". �(8.2)

33. ☐ Determine the final volume of the titrant in the buret and record the volume to 0.01 mL in the HCl column in Table 1 above.

34. ☐ Turn off the magnetic stirrer.

35. ☐ Remove the beaker and dispose of its contents according to the teacher's instructions.

36. ☐ Place the waste container under the pH sensor and use the wash bottle to thoroughly clean the micro stir bar and the pH sensor.

37. ☐ Dispose of this waste according to the teacher's instructions.

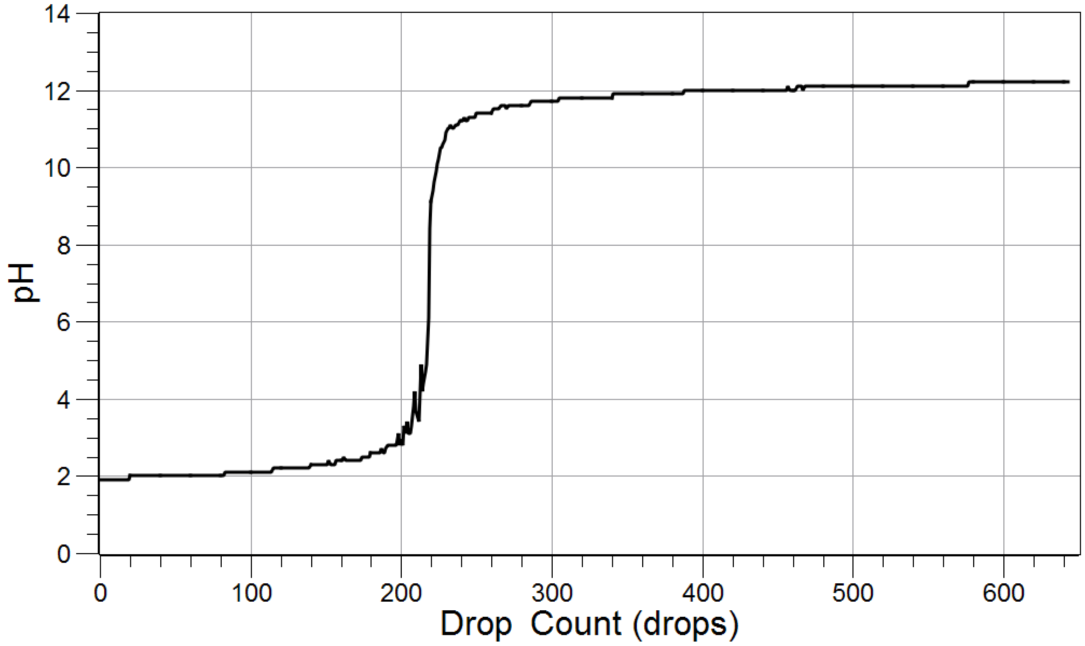
38. ☐ Perform a titration of acetic acid by repeating the steps in the Set Up and Collect Data sections above, this time substituting acetic acid. Take into account the following differences when you repeat the steps:

- Record the data collected in the acetic acid column in Table 1 above.
- Use 10.0 mL of acetic acid.
- Name the data run "Acetic acid".

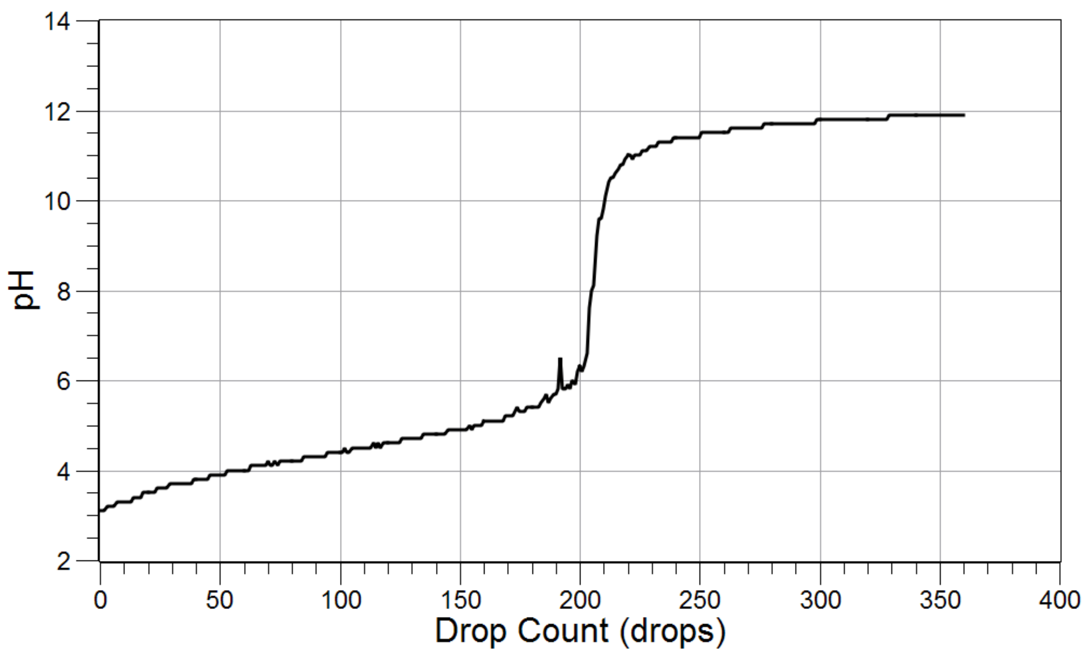
39. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Sample Data

Hydrochloric Acid Trial



Acetic Acid Trial



Data Analysis

1. ☐ Determine the total volume of NaOH used during each titration. Record the total volume used in table 2 below.

Total volume NaOH used = Final volume NaOH – Initial volume NaOH

Table 2: Volume of NaOH used in each titration

	HCl Trial (drops)	**$HC_2H_3O_2$ Trial (drops)**
Final volume of NaOH (mL)	29.46	17.55
Initial volume of NaOH (mL)	0.21	0.55
Total volume NaOH used (mL)	29.25	17.00

2. ☐ Use the pH versus Drop Count (drops) graph to determine the total number of drops used in each trial. Follow the steps below to complete this on your data collection system.

a. Display the run of data you want to analyze. �(7.1.7)

b. Find the final drop count by finding the coordinates of the last data point collected. �(9.1)

c. Record the final drop count for each trial in Table 3 below.

Table 3: Final drop count at the end of each titration

	HCl Trial	**$HC_2H_3O_2$ Trial**
Final drop count	643	360

3. ☐ Create a calculation to convert drop count to volume (mL) for each trial. Follow the steps below to do this on your data collection system.

a. Write the mathematical equation that can be used to convert drop count to volume. The general equation is given below, but you need to replace "total volume of titrant used" and "final drop count" with the numerical values determined above. Write the mathematical equation for each trial in Table 4 below.

$$\text{calcvolume} = [\text{Drop Count}] * (\text{total volume of titrant used} / \text{final drop count})$$

Note: In the equation above "calcvolume" stands for calculated volume. You will have a different calculation for each trial so the two calculated volumes need to have different names.

b. Enter the equations you determined above into the data collection system. �(10.3)

Table 4: Calculations for converting NaOH drop counts to volumes

Mathematical equation for the HCl trial: $$\text{calcvolumehcl} = [\text{Drop Count}] * (29.25 / 643)$$
Mathematical equation for the $HC_2H_3O_2$ trial: $$\text{calcvolumeacetic} = [\text{Drop Count}] * (17.00 / 360)$$

4. ☐ Determine the pH and volume of NaOH at the equivalence point for each trial. Follow the steps below to do this on your data collection system.

a. Change the units on the x-axis to the calculated volume for the run of data your want to analyze. �(7.1.9)

b. Display the run of data you want to analyze. �(7.1.7)

c. Find the coordinates of the equivalence point. The equivalence point is the data point with the greatest slope. �(9.3)

d. Record the pH and Volume of NaOH at the equivalence point in Table 5 below.

Table 5: pH at the equivalence point and volume of NaOH used to reach the equivalence point

	HCl Trial	**$HC_2H_3O_2$ Trial**
pH at the equivalence point	6.1	7.6
Volume of NaOH at the equivalence point (mL)	10.6	9.94

5. ☐ Calculate the molar concentration of each of the acid solutions. Use the following steps as a guide and record your work for each step in the tables provided.

a. Determine the number of moles of NaOH added using the volume of NaOH at the equivalence point and the molarity of the standardized NaOH solution.

b. Convert from moles of NaOH to moles of acid using the balanced chemical equation.

c Use the moles of acid and the starting volume of acid to determine the molarity of the acid.

Table 6: Calculating the molar concentration of the hydrochloric acid (HCl) solution

Name of Calculation	HCl Trial Show your work below
Moles of NaOH at the equivalence point (mol)	$(0.0106\ \text{L NaOH})\left(\frac{0.098\ \text{mol NaOH}}{1\ \text{L NaOH}}\right) = .0010\ \text{mol}$
Balanced chemical equation	$NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$
Moles of acid in solution (mol)	$(0.0010\ \text{mol NaOH})\left(\frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}}\right) = 0.0010\ \text{mol HCl}$
Concentration of acid solution (M)	$\left(\frac{0.0010\ \text{mol HCl}}{.0100\ \text{L HCl}}\right) = 0.10\ \text{M HCl}$

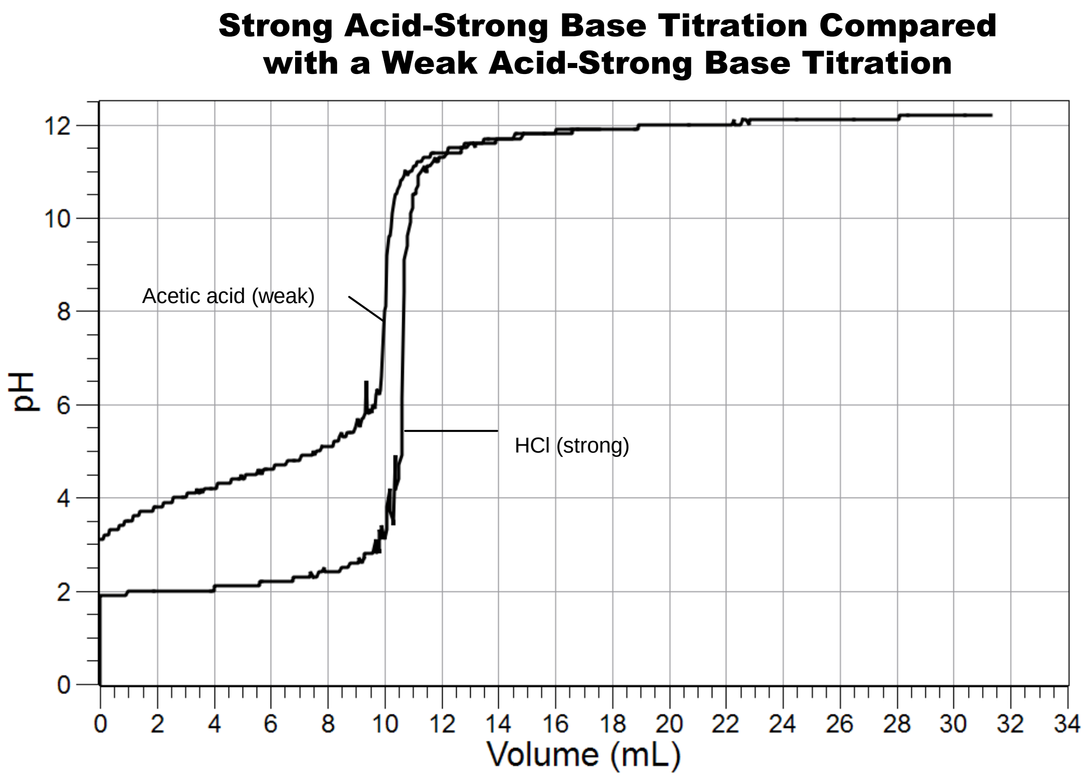
Table 7: Calculating the molar concentration of the acetic acid ($HC_2H_3O_2$) solution

Name of Calculation	Acetic Acid Trial Show your work below
Moles of NaOH at the equivalence point (mol)	$(0.00994\ \text{L NaOH})\left(\frac{0.098\ \text{mol NaOH}}{1\ \text{L NaOH}}\right) = 9.7 \times 10^{-4}\ \text{mol}$
Balanced chemical equation	$NaOH(aq) + HC_2H_3O_2(aq) \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$
Moles of acid in solution (mol)	$(9.7 \times 10^{-4}\ \text{mol NaOH})\left(\frac{1\ \text{mol } HC_2H_3O_2}{1\ \text{mol NaOH}}\right) = 9.7 \times 10^{-4}\ \text{mol } HC_2H_3O_2$
Concentration of acid solution (M)	$\left(\frac{9.7 \times 10^{-4}\ \text{mol } HC_2H_3O_2}{0.0100\ \text{L } HC_2H_3O_2}\right) = 0.097\ \text{M } HC_2H_3O_2$

6. ☐ On your data collection system, create a graph of pH versus Volume of NaOH (mL) with both runs of data displayed on the same set of axes. �(7.1.3)

Note: Use either of the calculated volumes on the y-axis. This graph is for comparison purposes only.

7. ☐ Sketch or print a copy of the graph of pH versus Volume of NaOH (mL) with both runs of data on one set of axes. Label each run of data as well as the overall graph, the x-axis, the y-axis, and include numbers on the axes. �(11.2)



Analysis Questions

1. What is the significance of the point on the titration curve where the slope is the steepest?

The point where the slope is the steepest marks the equivalence point. The equivalence point is significant because it is the point in the titration where, in this case, the amount of base added equals (or is very close to) the amount of acid that was originally present.

2. What trend did you notice, if any, in the slope of the titration curve between the start of the titration and the equivalence point? Propose an explanation for any trend you observed.

The slope started out very gently and became steeper as the titration got closer to the equivalence point. This was because each drop of titrant added supplied a roughly equal amount of NaOH to react with the acid in the beaker. At early stages in the titration, the amount of acid that reacted with the NaOH in the drop was small compared to the total amount of acid present, so the pH did not increase much. However, as the equivalence point was approached, each amount of acid that reacted with the NaOH in a drop was a much larger portion of the acid present, so the pH increased more quickly, leading to a steeper slope.

3. What trend did you notice, if any, in the slope of the titration curve between the equivalence point and the point where the titration was stopped? Propose an explanation for the trend you observed.

Again, each drop added (roughly) an equal amount of NaOH to the beaker. Nearer to the equivalence point, there was not much NaOH present, so each drop added increased the amount of NaOH present to a much greater extent compared to later stages in the titration. The pH, therefore, increased rapidly. As more NaOH was added to the beaker, each drop added less NaOH relative to what was present, so the pH did not increase as rapidly. This is why the slope of the curve leveled off as the titration proceeded.

4. What is the likelihood that the concentration of acid and base is exactly equal at the experimentally determined equivalence point? Explain your reasoning.

It is highly unlikely that an experimentally determined equivalence point has exactly the same amount of base as acid. This is because the smallest amount of base that can be added to the acid is one drop. All of the NaOH in the drop added to get to the point with the steepest slope may not have been necessary to react with the last amount of acid needed to have an equal amount of each.

5. What, if any, difference do you notice between the start of the titration and the equivalence point in the titration curves for the two different acids? Explain any difference that may exist. (You may assume the concentrations of the two acid solutions are the same.)

The pH of the acetic acid solution starts higher and remains higher throughout the titration until the equivalence point. This is because acetic acid is a weak acid and it does not fully dissociate in water. Therefore, the amount of hydronium ions present in solution is lower, and the pH is consequently higher.

6. What is the difference in pH at the equivalence points between the two acids? Explain.

The pH at the equivalence point for the HCl titration is very close to 7, and the equivalence point for the acetic acid titration is greater than 7. This is because when HCl is titrated with NaOH, the only species present at the equivalence point are water and sodium chloride, neither of which would cause the pH to be different from pure water.

At the equivalence point of acetic acid, however, the species present in solution are sodium acetate and water. The acetate ion itself is a base and reacts with water:

$C_2H_3O_2^-(aq) + H_2O(l) \rightarrow HC_2H_3O_2(aq) + OH^-(aq)$

The hydroxide ion that is produced causes the pH to be greater than 7.

Synthesis Questions

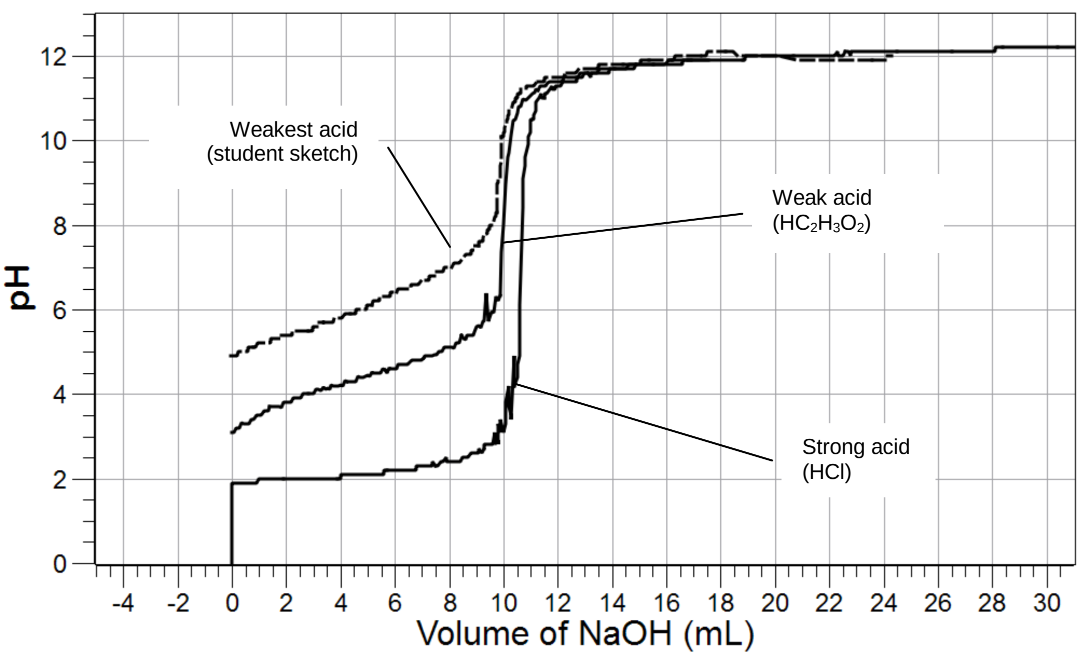
Use available resources to help you answer the following questions.

1. How can the concentration of an unknown sodium hydroxide solution be determined?

Perform a titration using a known concentration of a known acid, such as 0.1 M HCl, as the titrant.

2. Sketch the titration curves for a strong acid (HCl), a weak acid (acetic acid), and an acid that is weaker than acetic acid below. Label the curve for each acid as well as the overall graph, the x-axis, the y-axis, and include numbers on the axes.

Sketch of an Acid Weaker than Acetic Acid



3. Explain the difference between the strength and the concentration of an acidic solution.

The strength of an acidic solution refers to the degree to which the hydrogen ion dissociates. Concentration, however, refers to the amount of total acid molecules regardless of whether or not they dissociate.

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. For the titration of an acid with a base, the pH will start ________ and finish ________.

- **A.** low; low
- **B.** high; low
- **C.** **low; high**
- **D.** high; high

2. On a titration curve, the equivalence point is:

- **A.** The point with the smallest slope
- **B.** **The point with the greatest slope**
- **C.** The point with a slope of zero
- **D.** When the pH is equal to zero

3. If the hydronium ion concentration in an aqueous solution increased, what would happen to the hydroxide ion concentration and the pH of the solution?

- **A.** Hydroxide ion concentration would increase; pH would decrease
- **B.** Hydroxide ion concentration would decrease; pH would increase
- **C.** **Hydroxide ion concentration would decrease; pH would decrease**
- **D.** Hydroxide ion concentration would increase; pH would increase

4. What is a reaction between an acid and a base called?

- **A.** A pH reaction
- **B.** A titration reaction
- **C.** An s-shaped curve
- **D.** **A neutralization reaction**

5. Why are titrations performed?

- **A.** **To determine the concentration of known solutions**
- **B.** To determine the pH of known solutions
- **C.** To determine the type of molecules in a solution
- **D.** To differentiate an acid from a base

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. Acid-base reactions are extremely important and are present in many areas, from acid rain to the chemistry of the blood. An acid is a substance that when dissolved in water, causes the hydronium ion (H_3O^+) concentration to **increase** and the hydroxide ion (OH^-) concentration to **decrease**. A base does exactly the opposite. A solution with equal concentrations of H_3O^+ ions and OH^- ions is known as **neutral.** A solution containing more H_3O^+ ions than OH^- ions is considered **acidic** and has a pH **less** than 7. A solution containing fewer H_3O^+ ions than OH^- ions is considered **basic** and has a pH **greater** than 7. When an acid and base react, they form a **salt** and water. This is called a **neutralization** reaction.

2. A **titration** is a common quantitative laboratory method used to determine the concentration of a reactant. A reactant with a known concentration, called the **titrant**, is added to a known volume of a second reactant, called the **analyte**. In a titration setup, the titrant is placed in the **buret** and the analyte is placed in the **beaker**. As the titrant is added to the analyte, the pH of the solution in the beaker changes as the ratio of H_3O^+ ions and OH^- ions change. The point in the titration at which the amount of titrant added reacts exactly with the amount of analyte present is known as the **equivalence point**. On a titration curve, the equivalence point can be identified as the point with the **greatest** slope. The volume of the titrant added is multiplied by the **molarity** of the titrant to determine the number of moles in the titrant. The number of analyte moles can be determined using the **mole ratio** between the reactants as indicated in the balanced chemical equation. Finally, the number of analyte moles is **divided** by the starting volume of the analyte to give the analyte concentration.

Extended Inquiry Suggestions

Determine the percentage of acetic acid in a commercial sample of vinegar.

Have students sketch, or actually carry out, the titration of a polyprotic acid with a strong base or a polybasic base with a strong acid.

Investigate the causes and effects of acid rain.

Collect water samples from local rivers or lakes and test the acid/ base content.

Investigate the effectiveness of various antacids to neutralize stomach acid using the "back-titration" technique.

Investigate the acid content of various fruit juices.

27. Diprotic Titration: Multi-Step Chemical Reactions

Objectives

Determine the concentration of a sodium carbonate solution. Through this investigation, students:

- Learn that chemical reactions can be the sum of several individual reactions
- Learn that the conjugate base of a polyprotic acid generally accepts one proton at a time

Procedural Overview

Students conduct the following procedures:

- Perform a strong acid-strong base titration
- Calculate the concentration of an unknown sodium carbonate solution using collected data and stoichiometric calculations
- Calculate the number of moles and liters of carbon dioxide gas produced using collected data and stoichiometric calculations

Time Requirement

♦ Preparation time	20 minutes
♦ Pre-lab discussion and activity	20 minutes
♦ Lab activity	60 minutes

Materials and Equipment

For each student or group:

- Data collection system
- Drop counter
- pH sensor
- Micro stir bar
- Magnetic stirrer
- Beaker (2), 50-mL
- Beaker, 250-mL
- Graduated cylinder, 50-mL
- Graduated cylinder, 100-mL
- Transfer pipet
- Buret, 50-mL
- Buret clamp
- Ring stand
- Right-angle clamp
- Funnel
- Waste container
- Wash bottle filled with distilled (deionized) water
- Buffer solution, pH 4, 25 mL
- Buffer solution, pH 10, 25 mL
- Distilled (deionized) water, 200 mL
- Sodium carbonate (Na_2CO_3) solution, 40 mL[1]
- 1.0 M Hydrochloric acid (HCl), 110 mL[2]

[1]The sodium carbonate solution should be approximately 0.5 M. To formulate this solution using solid sodium carbonate (Na_2CO_3), refer to the Lab Preparation section.

[2]To formulate using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Balanced chemical equations
- Concentration (molarity)
- Molar ratios and stoichiometric calculations
- Limiting reactants
- Standard molar volume
- Properties of acids and bases
- pH
- Titration curves

Related Labs in This Guide

Labs conceptually related to this one include:

- pH of Household Chemicals
- An Acid-Base Titration
- Hess's Law
- Stoichiometry

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- Starting a new experiment on the data collection system �(1.2)
- Connecting a sensor to the data collection system �(2.1)
- Connecting multiple sensors to the data collection system �(2.2)
- Calibrating a pH sensor �(3.6)
- Recording a run of data �(6.2)
- Displaying data in a graph �(7.1.1)
- Adjusting the scale of a graph �(7.1.2)
- Displaying multiple data runs on a graph �(7.1.3)
- Showing and hiding data runs in a graph �(7.1.7)
- Changing the variable on the x-axis or y-axis �(7.1.9)
- Naming a data run �(8.2)
- Finding the values of a point in a graph �(9.1)
- Finding the slope at a point �(9.3)
- Creating calculated data �(10.3)
- Saving your experiment �(11.1)
- Printing �(11.2)

Background

An acid is a substance that is able to donate a hydrogen ion (Brønsted-Lowry acid). Acids are further classified as monoprotic, diprotic, or triprotic acids depending on the number of hydrogen ions they can donate. Monoprotic acids, such as hydrochloric acid (HCl) and nitric acid (HNO_3), are able to donate one hydrogen ion. Diprotic acids, such as sulfuric acid (H_2SO_4), are able to donate two hydrogen ions, while triprotic acids, such as phosphoric acid (H_3PO_4), are able to donate three hydrogen ions.

A base is the complement of an acid. A base is a substance that accepts a hydrogen ion (Brønsted-Lowry base). When an acid donates a hydrogen ion, the remaining group of atoms is called a conjugate base. For example, the conjugate base of HCl is Cl^- and the conjugate base of H_2SO_4 is HSO_4^-. A conjugate acid is formed when a base gains a hydrogen ion, so HCl is the conjugate acid of the base Cl^-. Conjugate acids and conjugate bases always occur in pairs with their complementary acid or base.

In this experiment, students perform a titration to determine the concentration of a sodium carbonate (Na_2CO_3) solution. The students add a standardized hydrochloric acid (HCl) solution to neutralize the sodium carbonate solution. The overall chemical reaction is:

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$$

The equation suggests that carbon dioxide bubbles should be visible as soon as the reactants are mixed. Students will notice, however, that carbon dioxide does not form immediately upon addition of acid, but does form later as more HCl is added. There is a delay because the reaction occurs in multiple steps to form carbonic acid, a diprotic acid. In the first step, the carbonate ion gains one hydrogen ion to form sodium hydrogen carbonate ($NaHCO_3$) and sodium chloride (NaCl).

$$HCl(aq) + Na_2CO_3(aq) \rightarrow NaHCO_3(aq) + NaCl(aq)$$

Once all of the carbonate ions have received one proton to form the bicarbonate ion (HCO_3^-), the reaction proceeds to the final products. In the second step, HCl reacts with sodium hydrogen carbonate to form carbonic acid. The carbonic acid immediately decomposes into carbon dioxide and water. It is during this step that the bubbles of carbon dioxide gas are released.

$$HCl\ (aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl\ (aq) \rightarrow CO_2(g) + H_2O(l) + NaCl\ (aq)$$

The overall reaction is the sum of the two individual reaction steps.

Step 1:		$HCl(aq)$	+	$Na_2CO_3(aq)$	$\rightarrow$	$\cancel{NaHCO_3(aq)}$	+	$NaCl(aq)$
Step 2:	+	$HCl\ (aq)$	+	$\cancel{NaHCO_3(aq)}$	$\rightarrow$	$CO_2(g)$	+	$NaCl(aq)$
Overall Reaction:		$2HCl\ (aq)$	+	$Na_2CO_3\ (aq)$	$\rightarrow$	$CO_2(g)$	+	$2NaCl\ (aq)$

On a titration curve, each step has its own equivalence point. The first equivalence point occurs when the original number of moles of hydrogen ions is equal to the number of moles of carbonate ions (CO_3^{2-}). At the first equivalence point, all of the carbonate ions have been converted to hydrogen carbonate ions. The second equivalence point occurs when the number of moles of hydrogen ions is equal to the number of moles of hydrogen carbonate ions (HCO_3^-). At the second equivalence point, all of the hydrogen carbonate ions have been converted to carbonic acid, which immediately decomposes into carbon dioxide and water.

The concentration of sodium carbonate can be determined using the equivalence point for either the first or second step. The volume of HCl used to get to the first equivalence point should be equal to the volume of HCl to get from the first equivalence point to the second.

This titration can also be performed twice using two different indicators. Phenolphthalein indicator will change color at the first equivalence point (going from magenta to colorless) and methyl orange indicator will change at the second equivalence point (going from yellow to red-orange).

Pre-Lab Discussion and Activity

A Puzzle to Solve

Hold up a solution of sodium carbonate (Na_2CO_3) and have your students predict what they will observe when hydrochloric acid (HCl) is added. Write the chemical reaction on the board for the students to use as a basis for their prediction.

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$$

Use a transfer pipet to add several drops of acid. Comment out loud that no bubbles are forming. Ask the students why there are no bubbles forming. Continue to add acid until bubbles start to form. Ask the students if they have ever seen the "vinegar and baking soda volcano" reaction. Mention that you will be substituting hydrochloric acid for the vinegar; this is valid, since both are acids. Demonstrate the reaction by adding dilute hydrochloric acid directly to a beaker containing baking soda. Allow the reaction to foam up and out of the beaker and into a tray. Ask them if there is a delay in the formation of bubbles. Write the reaction on the board.

$$HCl(aq) + NaHCO_3(aq) \rightarrow CO_2(g) + NaCl(aq) + H_2O(l)$$

Engage the students in a discussion about the similarities and differences between the two reactions and possible explanations for the delay in bubble formation by asking the following questions:

1. What do you think will happen when hydrochloric acid is added to a solution of sodium carbonate?

Bubbles of carbon dioxide will form, as shown in the chemical equation.

2. Why are bubbles not forming? Is the chemical equation incorrect?

At first, there are no bubbles. The equation is correct, but the reaction occurs in multiple steps. The bubbles do not form during the first step, but they do form after a few minutes when the second step is underway.

3. What happens when acid is added to sodium bicarbonate (vinegar added to baking soda)?

Bubbles form immediately.

4. What is the difference between the ions involved in each of the reactions?

In the first reaction, the polyatomic ion is a carbonate ion (CO_3^{2-}) and in the second reaction, it is a bicarbonate ion (HCO_3^-). The ions differ by one hydrogen ion. Their cations (sodium) are the same in both reactions.

5. What type of reactions are these? Are the products, CO_2, H_2O, and NaCl expected?

Since these are acid-base neutralization reactions, both a salt and water are predicted to form. Students should expect NaCl and H_2O, but they may question the carbon dioxide.

6. Which is the acid, and which is the base?

HCl is the acid in both reactions. The base is the sodium carbonate in the first reaction and sodium hydrogen carbonate in the second reaction.

7. Where does the carbon dioxide come from?

The hydrogen ion reacts with the carbonate ion or hydrogen carbonate ion of the base to form carbonic acid (H_2CO_3). Carbonic acid decomposes into CO_2 and H_2O.

8. Does the decomposition of carbonic acid cause the delay in the bubbles? Explain your reasoning.

No, because when acid is added to sodium hydrogen carbonate there is no delay in CO_2 formation.

9. How can we explain the delay in carbon dioxide formation in the reaction between hydrochloric acid and sodium carbonate?

The reaction must take place in two parts. In the first reaction, the carbonate ion gains a hydrogen ion to form the hydrogen carbonate ion.

$HCl(aq) + Na_2CO_3(aq) \rightarrow NaHCO_3(aq) + NaCl(aq)$

Once this reaction is complete, the decomposition of H2CO3 in the second reaction can occur. It is the second reaction that creates carbon dioxide. So, the delay is caused by the first step.

$HCl\ (aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl\ (aq) \rightarrow CO_2(g) + H_2O(l) + NaCl\ (aq)$

10. How can we combine these two steps into one, overall reaction?

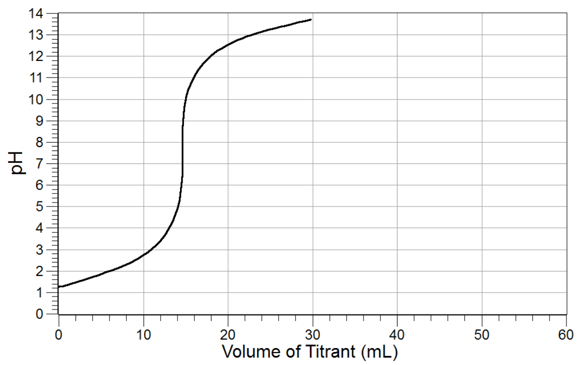
Add the reactions together.

Step 1:		$HCl(aq)$	+	$Na_2CO_3(aq)$	$\rightarrow$	$\cancel{NaHCO_3(aq)}$	+	$NaCl(aq)$
Step 2:	+	$HCl\ (aq)$	+	$\cancel{NaHCO_3(aq)}$	$\rightarrow$	$CO_2(g)$	+	$NaCl(aq)$
Overall Reaction:		$2HCl\ (aq)$	+	$Na_2CO_3\ (aq)$	$\rightarrow$	$CO_2(g)$	+	$2NaCl\ (aq)$

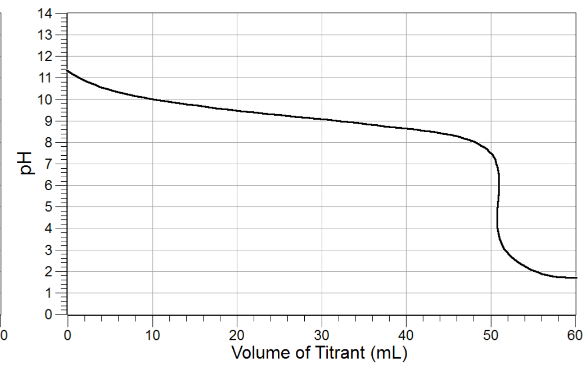
Titration Curves and Polyprotic Acids

Review with the students the terms conjugate acid, conjugate base, and polyprotic acids. Show the students four titration curves, like the ones below, and have them predict what the titration curve will look like when they add hydrochloric acid to their unknown concentration of sodium carbonate. Review the key parts of a titration curve.

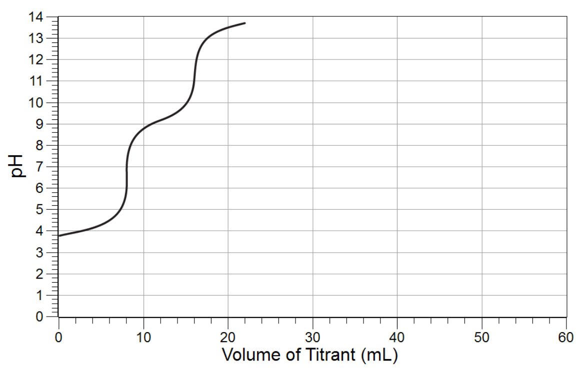
Titration curve A



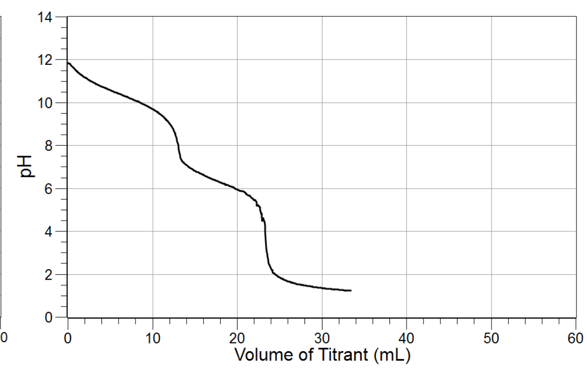
Titration curve B



Titration curve C



Titration curve D



11. **Where is the equivalence point on a titration curve? What does it indicate?**

The equivalence point is where the slope of the curve is the steepest. The equivalence point is when the moles of reactant and product are equal.

12. **Using the terms "base" and "conjugate acid," describe the relationship between CO_3^{2-} and HCO_3^-.**

In the first step, CO_3^{2-} is a base and the resulting HCO_3^- is its conjugate acid.

13. **What is the difference between monoprotic and polyprotic acids?**

Monoprotic acids, such as HCl, only have one hydrogen ion to donate. Polyprotic acids have two or more hydrogen ions to donate; a diprotic acid has two hydrogen ions, and a triprotic acid has three.

14. **In this lab, carbonic acid (H_2CO_3) is being formed. How many hydrogen ions can it produce? Is this acid classified as monoprotic, diprotic, or triprotic?**

Since carbonic acid has two hydrogen ions it can donate, it is a diprotic acid.

15. Do conjugate bases of polyprotic acids accept hydrogen ions one at a time or all at once? How do you know?

Conjugate bases of polyprotic acids tend to accept one proton at a time. If the protons were accepted all at once, bubbles of carbon dioxide gas would have formed immediately when hydrochloric acid was added to the solution of sodium carbonate.

16. Which of the titration curves represent a diprotic acid? How do you know?

Both C and D are from the titration of a diprotic acid. In titration curve C, the diprotic acid was titrated with a base, as evident by its initial and final pH values. In titration curve D, the diprotic acid's conjugate base was titrated with an acid, because its pH started at a high value and ended at a low value. They are known to be diprotic because they have two equivalence points. Each equivalence point represents equal amounts of products and reactants for each of the steps that make up the overall reaction.

17. Which of the titration curves resemble the one expected for the reaction between sodium carbonate and hydrochloric acid?

It will look like titration curve D. The analyte (sodium carbonate) is a base, so it will start at a high pH and will slowly decrease as the titrant (HCl) is added. Since the reaction involves two steps (formation of a diprotic acid), the graph will have two equivalence points.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- ♦ Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- ♦ If acid or base solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- ♦ Diluting acids and bases creates heat; be extra careful when handling freshly prepared solutions and glassware as they may be very hot.
- ♦ Always add acids and bases to water, not the other way around, as the solutions may boil vigorously.
- ♦ Handle concentrated acids and bases in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Prepare 1000 mL of 0.50 M sodium carbonate (Na_2CO_3). This is enough for 20 lab groups.

   a. Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.

   b. Slowly add 53.0 g of solid Na_2CO_3 to the beaker and allow it to dissolve completely with continuous stirring.

   c. Allow the solution to cool, then carefully pour the solution into a 1000-mL volumetric flask and dilute to the mark with distilled water.

   d. Cap and invert three times to ensure complete mixing.

2. Prepare 1000 mL of 1.0 M hydrochloric acid from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 10 lab groups.

Starting with concentrated (12 M) HCl:

a. Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.

b. Slowly add 83.3 mL of 12 M HCl to the beaker with continuous stirring.

c. Allow the solution to cool, then carefully pour the solution into a 1000-mL volumetric flask and dilute to the mark with distilled water.

d. Cap and invert three times to ensure complete mixing.

Starting with dilute (6 M) HCl:

a. Add approximately 500 mL of distilled water to a 1000-mL volumetric flask.

b. Add 166.7 mL of 6 M HCl to the water and dilute to the mark with distilled water.

c. Cap and invert three times to ensure complete mixing.

Safety

Add these important safety precautions to your normal laboratory procedures:

- Hydrochloric acid is a strong acid. Avoid contact with the skin and eyes.
- Be sure that all acids and bases are neutralized before disposal down the drain.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

2	1	3	5	4
Rinse the buret with 1.0 M hydrochloric acid and practice adjusting the stopcock. Collect the HCl in a waste container.	Set up the titration equipment and calibrate the pH sensor.	Place a 250-mL beaker containing diluted analyte (sodium carbonate) on the magnetic stirrer.	Repeat the titration to collect a second set of data.	Titrate the sodium carbonate solution with HCl until the pH stabilizes at a pH less than 2.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a pH sensor to the data collection system. �(2.1)

3. ☐ Place 25 mL of pH 4 buffer solution in a 50-mL beaker and 25 mL of pH 10 buffer solution in a second 50-mL beaker. Use these solutions to calibrate the pH sensor. �(3.6)

4. ☐ Will a pH sensor that is not calibrated give precise results? Why must the pH sensor be calibrated?

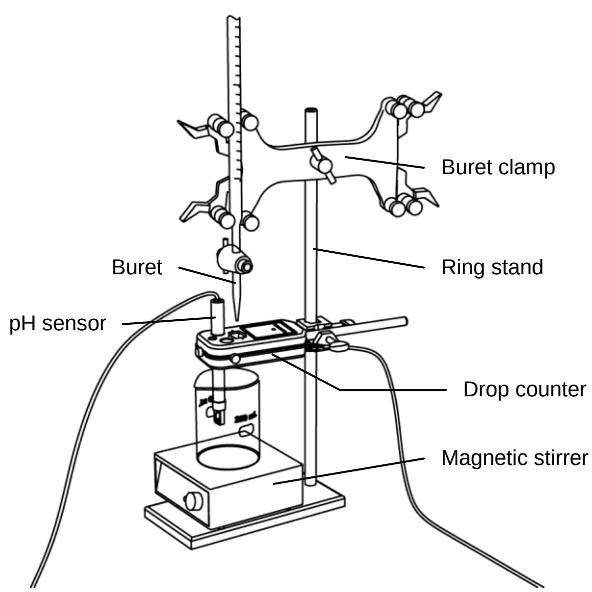
Yes, a pH sensor that is not calibrated will give precise results, but they may not be accurate. Calibrating the pH sensor ensures that the pH readings are accurate as well as precise.

5. ☐ Connect a drop counter to the data collection system. �(2.2)

6. ☐ Display pH versus Drop Count (drops) on a graph. �(7.1.1)

7. ☐ Assemble the titration apparatus, using the steps below and the illustration as a guide.

- **a.** Assemble the ring stand.
- **b.** Position the magnetic stirrer on (or next to) the base of the ring stand.
- **c.** Place a waste container on the magnetic stirrer.
- **d.** Use the buret clamp to attach the buret to the ring stand.
- **e.** Position the drop counter over the waste container and attach it to the ring stand using the right-angle clamp.
- **f.** Place the pH sensor through one of the slots in the drop counter.



8. ☐ Rinse the buret with several milliliters of the standardized HCl solution. Follow the steps below to complete this step.

a. Ensure that the stopcock is closed and use a transfer pipet to rinse the inside of the buret with several milliliters of the HCl solution.

b. Open the stopcock on the buret and drain the HCl rinse into the waste container.

c. Repeat this process two more times.

9. ☐ Make sure the stopcock is in the "off" position and then use a funnel to fill the buret with about 50 mL of the HCl solution (titrant).

10. ☐ Drain a small amount of the titrant through the drop counter into the waste beaker.

11. ☐ Why is it necessary to drain a small amount of titrant through the drop counter before you begin a titration?

Draining a small amount of titrant will remove any air that is trapped in the tip of the buret. Any air trapped in the buret is counted as volume of HCl. If this happens, the final amount of HCl used will be inaccurate.

12. ☐ Practice adjusting the stopcock on the buret so that the titrant goes through the drop counter in distinguishable drops that fall at about 2 to 3 drops per second.

Note: It is important that you have good control of adjusting the stopcock. If you accidently open the stopcock too far and the HCl flows out (as opposed to drops out), then you will have to start over.

13. ☐ How can the green light on the drop counter help you determine if the HCl is falling in distinguishable drops?

The green light blinks each time a drop falls through the drop counter. If there is a solid green light, the HCl is flowing too fast and the drop counter cannot distinguish the individual drops. If the green light is blinking, then the drop counter is functioning properly and the flow rate is correct.

14. ☐ Close the stopcock and then remove the waste container.

Data Collection

Table 1: Titration data

Measurement	Trial 1	Trial 2
Concentration of the HCl solution (M)	1.0	1.0
Volume of Na_2CO_3 used (mL)	20.05	18.7
Initial volume of HCl in the buret (mL)	1.60	0.10
Final volume of HCl in the buret (mL)	36.50	34.55

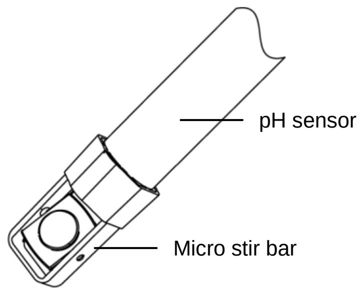
15. ☐ Record the exact concentration of the standardized HCl solution in Table 1 above.

16. ☐ Using the 100-mL graduated cylinder, measure 100.0 mL of distilled water and add it to a 250-mL beaker.

17. ☐ Using the 50-mL graduated cylinder, measure approximately 20.0 mL of the sodium carbonate (Na_2CO_3) solution and add to the 100 mL of distilled water in the beaker.

18. ☐ Record the exact volume of Na_2CO_3 used in Table 1 above.

19. ☐ Add the micro stir bar to the end of the pH sensor.



20. ☐ Position the 250-mL beaker containing the sodium carbonate on the magnetic stirrer with the pH sensor submerged in the solution.

21. ☐ Ensure that the bulb of the pH sensor is fully submerged in the solution and then turn on the magnetic stirrer and begin stirring at a slow-to-medium speed.

22. ☐ Determine the initial volume of the titrant (HCl) in the buret to a precision of 0.01 mL, and record this in Table 1 above.

23. ☐ Start recording data. �(6.2)

24. ☐ Carefully open the stopcock on the buret so that 2 to 3 drops per second are released. If needed, rescale the axes so that you can see the changes taking place. �(7.1.2)

25. ☐ Continue to record data until the pH of the solution stabilizes at a pH less than 2. Do not let the volume of the titrant fall below the last mark on the buret.

26. ☐ Why do you need to ensure that the volume of the titrant does not fall below the zero mark on the buret?

If the amount of titrant in the buret falls below the last mark, there is no way to measure the exact amount of titrant added to the analyte and volumetric analysis cannot be completed.

27. ☐ Which reaction step is taking place at the beginning of the titration between the initial pH (~12) and a pH value of 8? Describe the reaction that is occurring.

The first step is occurring, where the acid (HCl) is reacting with the base (Na_2CO_3) to produce a salt (NaCl) and a new base ($NaHCO_3$).

HCl(aq) + Na_2CO_3(aq) → NaCl(aq) + $NaHCO_3$(aq)

28. ☐ What begins to happen when the pH of the solution is lower than 8? How can these observations be explained?

Carbon dioxide bubbles begin to form. This is because the second step of the reaction is taking place, in which additional acid (HCl) is reacting with the base created in the first step ($NaHCO_3$) to produce more salt (NaCl) and carbonic acid (H_2CO_3). The carbonic acid immediately decomposes into water (H_2O) and carbon dioxide (CO_2), which produces the observed bubbles.

HCl(aq) + $NaHCO_3$(aq) → NaCl(aq) + H_2CO_3(aq) → NaCl(aq) + CO_2(g) + H_2O(l)

29. ☐ Close the stopcock on the buret when the pH of the solution stabilizes at a pH less than 2.

30. ☐ Stop data recording. �(6.2)

31. ☐ Name the data run "trial 1". �(8.2)

32. ☐ Determine the final volume of the titrant in the buret and record the volume to 0.01 mL in Table 1 above.

33. ☐ Turn off the magnetic stirrer.

34. ☐ Remove the beaker and dispose of its contents according to the teacher's instructions.

35. ☐ Place the waste container under the pH sensor and use the wash bottle to thoroughly clean the micro stir bar and the pH sensor.

36. ☐ Dispose of this waste according to the teacher's instructions.

37. ☐ Thoroughly clean the 250-mL beaker so that you can reuse it for trial 2.

38. ☐ Collect a second set of data. Refer to the steps above as needed. Record the data in the Trial 2 column of Table 1 above.

39. ☐ Name the second run of data "trial 2".

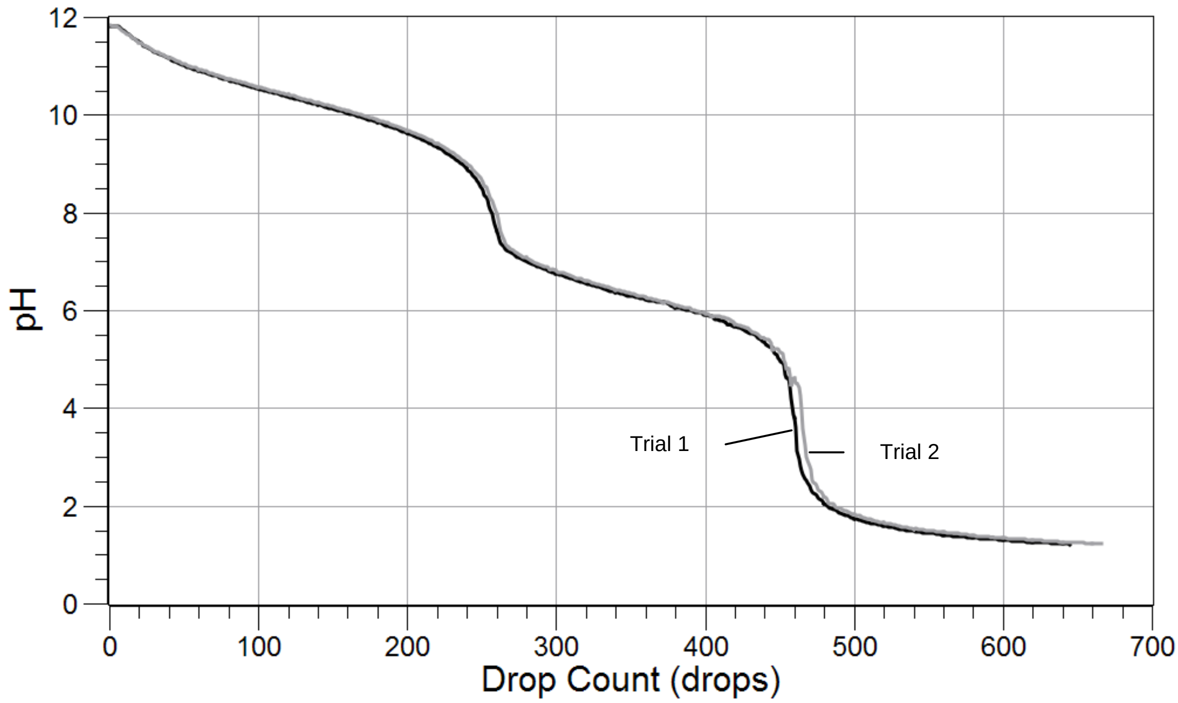
40. ☐ Why is it necessary to take multiple sets of data?

Collecting multiple sets of data ensures that you are able to reproduce the same results. Reproducibility is necessary to ensure reliable data.

41. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Sample Data

Titration curve of sodium carbonate with hydrochloric acid collected on a graph of pH versus Drop Count (drops).



Data Analysis

1. ☐ Determine the total volume of HCl titrant used in each trial. Record the total volume used in Table 2 below.

Total volume used = Final volume HCl – Initial volume HCl

Table 2: Total volumes of HCl titrant used in each trial

	Trial 1	Trial 2
Final volume of HCl (mL)	36.50	34.55
Initial volume of HCl (mL)	1.60	0.10
Total Volume HCl used (mL)	34.90	34.45

2. ☐ Use the pH versus Drop Count (drops) graph to determine the total number of drops used in each trial. Follow the steps below to complete this on your data collection system.

a. Display the run of data you want to analyze. �(7.1.7)

b. Find the final drop count by finding the coordinates of the last data point collected. �(9.1)

c. Record the final drop count for each trial in Table 3 below.

Table 3: Final drop count

	Trial 1	Trial 2
Final drop count	645	665

3. ☐ Create a calculation to convert drop count to volume (mL) for each trial. Follow the steps below to do this on your data collection system.

a. Write the mathematical equation that can be used to convert drop count to volume. The general equation is given below, but you need to replace "total volume of titrant used" and "final drop count" with the numerical values determined above. Write the mathematical equation for each trial in Table 4 below.

$$\text{calcvolume} = [\text{Drop Count}] * (\text{total volume of titrant used} / \text{final drop count})$$

Note: In the equation above "calcvolume" stands for calculated volume. You will have a different calculation for each trial so the two calculated volumes need to have different names.

b. Enter the equations you determined above into the data collection system. �(10.3)

Table 4: Calculations for converting HCl drop count to volume

Mathematical equation for trial 1: $\text{calcvolume1} = [\text{Drop Count}] * (34.90 / 645)$
Mathematical equation for trial 2: $\text{calcvolume2} = [\text{Drop Count}] * (34.45 / 665)$

4. ☐ Determine the volume of HCl at both equivalence points for each trial. Follow the steps below to do this on your data collection system.

a. Change the units on the x-axis to the calculated volume for that run of data you want to analyze. �(7.1.9)

b. Display the run of data you want to analyze. �(7.1.7)

c. Find the coordinates of each of the equivalence points. �(9.3)

d. Record the volume of HCl at each of the equivalence points in Table 5 below.

Table 5: Volume of titrant at equivalence points

	Trial 1	Trial 2
Volume at Equivalence Point 1 (mL)	12.93	12.23
Volume at Equivalence Point 2 (mL)	25.97	24.61

5. ☐ Calculate the molar concentration of the sodium carbonate solution using the first equivalence point. Use the following steps as a guide and record your work for each step in Table 6 below.

a. Determine the number of moles of HCl added at the first equivalence point using the volume at equivalence point #1 and the molarity of the HCl solution.

b. Convert from moles of HCl to moles of sodium carbonate using the balanced chemical equation for the first reaction step.

c. Use the moles of sodium carbonate calculated and the volume of sodium carbonate used as the analyte to determine the molarity of the base.

d. Repeat for Trial 2.

Table 6: Molar concentration of the sodium carbonate solution at the first equivalence point

	Trial 1 (show your work)	Trial 2
Moles of HCl at the first equivalence point (mol)	$(0.01293\,\text{L HCl})\left(\frac{1.00\ \text{mol HCl}}{1\,\text{L HCl}}\right) = 0.0129$	0.0122
Balanced chemical equation for the first reaction	$HCl(aq) + Na_2CO_3(aq) \rightarrow NaHCO_3(aq) + NaCl(aq)$	Same as Trial 1
Moles of sodium carbonate in solution (mol)	$(0.0129\ \text{mol HCl})\left(\frac{1\ \text{mol}\ Na_2CO_3}{1\ \text{mol HCl}}\right) = 0.0129$	0.0122
Concentration of Na_2CO_3 solution (M)	$\left(\frac{0.0129\ \text{mol}\ Na_2CO_3}{0.02005\ \text{L}\ Na_2CO_3}\right) = 0.643$	0.652

6. ☐ Calculate the molar concentration of the sodium carbonate solution using the second equivalence point. Use the following steps as a guide and record your work for each step in Table 7 below.

a. Determine the number of moles of HCl added at the second equivalence point using the volume at equivalence point 2 and the molarity of the HCl solution.

b. Convert from moles of HCl to moles of sodium carbonate using the overall balanced chemical equation.

c. Use the moles of sodium carbonate calculated and the volume of sodium carbonate used as the analyte to determine the molarity of the base.

d. Repeat for Trial 2.

Table 7: Molar concentration of the sodium carbonate solution at the second equivalence point

	Trial 1 (show your work)	**Trial 2**
Moles of HCl at the second equivalence point (mol)	$(0.02597\text{ L HCl})\left(\frac{1.00\text{ mol HCl}}{1\text{ L HCl}}\right) = 0.0260$	0.0246
Balanced chemical equation for the overall reaction	$2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$	Same as Trial 1
Moles of sodium carbonate in solution (mol)	$(0.0260\text{ mol HCl})\left(\frac{1\text{ mol }Na_2CO_3}{2\text{ mol HCl}}\right) = 0.0130$	0.0123
Concentration of Na_2CO_3 solution (M)	$\left(\frac{0.0130\text{ mol }Na_2CO_3}{0.02005\text{ L }Na_2CO_3}\right) = 0.648$	0.658

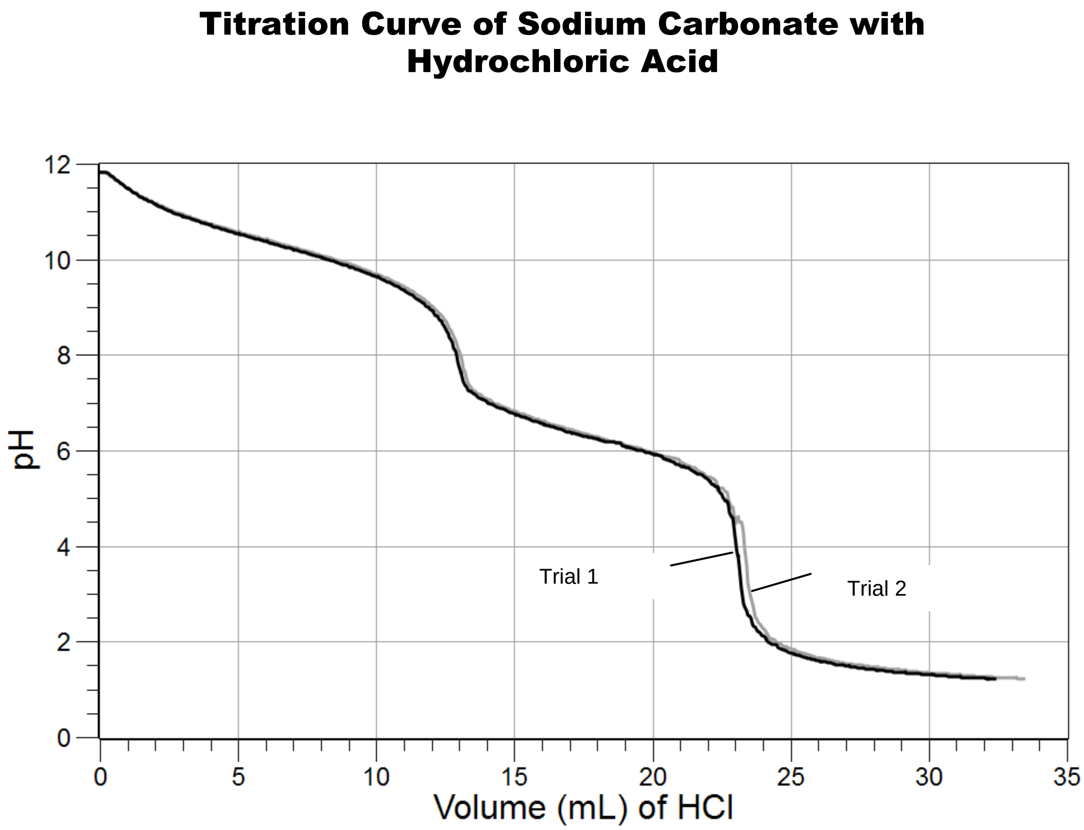
7. ☐ Calculate the average molarity of the sodium carbonate solution using all four calculated values.

$$\text{Average} = \frac{(0.643\text{ M} + 0.652\text{ M} + 0.648\text{ M} + 0.658\text{ M})}{4} = 0.650\text{ M }Na_2CO_3$$

8. ☐ On your data collection system, create a graph of pH versus Volume of HCl (mL) with both runs of data displayed on the same set of axes. �(7.1.3)

Note: Use either of the calculated volumes on the y-axis. This graph is for comparison purposes only.

9. ☐ Sketch or print a copy of the graph of pH versus Volume of HCl (mL) with both runs of data on one set of axes. Label each run of data as well as the overall graph, the x-axis, the y-axis, and include numbers on the axes. �(11.2)



Analysis Questions

1. When did the bubbles produced by the formation of carbon dioxide gas start to become visible?

Bubbles started to form after the first the equivalence point.

2. Write the two chemical equations as separate steps that add together to give the overall reaction: $2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$

Step 1: $HCl(aq) + Na_2CO_3(aq) \rightarrow NaHCO_3(aq) + NaCl(aq)$

Step 2: $HCl\ (aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl\ (aq) \rightarrow CO_2(g) + H_2O(l) + NaCl\ (aq)$

3. Before the first equivalence point, the beaker contained a mixture of both carbonate ions and bicarbonate ions. Which of these two ions accept hydrogen ions easier? How do you know?

The carbonate ion must accept H^+ easier, because no carbon dioxide was formed. If any of the bicarbonate ions accepted hydrogen ions, the bubbling would have started immediately.

4. After the first equivalence point, the carbon dioxide production was rapid. Were there any carbonate ions remaining in solution?

After the first equivalence point there are no longer any carbonate ions in solution.

5. What do you notice about the volume of acid needed to reach each equivalence point?

It required approximately the same amount of acid to reach the first equivalence point as it did to go from the first to the second equivalence point.

6. Why do you think the bubbling stopped after the second equivalence point?

At the second equivalence point, all of the original carbonate ions have gained two hydrogen ions to become carbonic acid which immediately decomposed into carbon dioxide and water. With no carbonate ions remaining, no new carbonic acid can be produced to decompose into carbon dioxide (stopping the bubbles from forming). The extra acid added beyond the second equivalence point has no carbonate ions to react with.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Often, the product of one chemical reaction can become the reactant in another. Give an example of how this statement is true using the reactions seen in this experiment.

The bicarbonate ions (HCO_3^-) produced in the first step of the titration became the reactant in the second step.

2. Calculate the number of moles of carbon dioxide produced in one of the trials. Identify the limiting reactant. Show your work, including the balanced chemical equation used.

Hydrochloric acid was in excess, and sodium carbonate was the limiting reactant.

The balanced chemical equation is:

$2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$

Trial #1: 0.0201 L Na_2CO_3 used

$$(0.0201\ \text{L}\ Na_2CO_3)\left(\frac{0.650\ \text{mol}\ Na_2CO_3}{1\ \text{L}\ Na_2CO_3}\right)\left(\frac{1\ \text{mol}\ CO_2}{1\ \text{mol}\ Na_2CO_3}\right) = 0.0131\ \text{mol}\ CO_2$$

3. Calculate the number of liters of carbon dioxide produced. Assume standard temperature and pressure.

$$(0.0131\ \text{mol}\ CO_2)\left(\frac{22.4\ \text{L}\ CO_2}{1\ \text{mol}\ CO_2}\right) = 0.293\ \text{L of}\ CO_2$$

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. In an acid-base titration, the equivalence point of the curve shows:

- **A. The point where there are the same number of moles of acid and base**
- **B.** The point where the buret has run out of acid
- **C.** The point where pH is a maximum
- **D.** The point where pH is a minimum

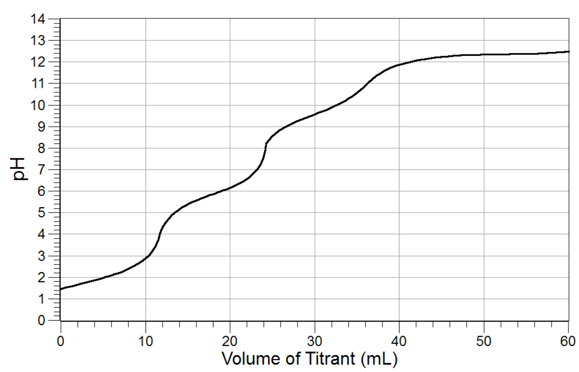
2. Carbonic acid is known to be a *diprotic* acid. Why?

- **A.** Because it reacts with a base
- **B. Because it has two hydrogen ions that can be donated**
- **C.** Because it makes carbon dioxide
- **D.** Because it can act as an acid or a base

3. How many equivalence points are there on a titration of a diprotic acid?

- **A.** 0
- **B.** 1
- **C. 2**
- **D.** 3

4. The titration curve below represents what type of acid?



- **A.** Monoprotic acid
- **B.** Diprotic acid
- **C. Triprotic acid**
- **D.** Tetraprotic acid

5. What gas is released when an acid reacts with sodium carbonate?

- **A.** Hydrogen gas
- **B.** Oxygen gas
- **C.** Methane gas
- **D. Carbon dioxide gas**

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The Brønsted-Lowry definition of acids and bases states that **acids** donate hydrogen ions and **bases** accept hydrogen ions. When an acid **donates** a hydrogen ion, the remaining anionic portion is called the **conjugate base**. The carbonate ion used in this experiment has a negative two charge. It can accept **two** hydrogen ions to become carbonic acid and is an example of a **diprotic** acid. When the carbonate ions react with acid in the first step, they initially become **bicarbonate (HCO_3^-)** ions (write both the name and formula). These ions can then react in the second step with additional acid to become **carbonic acid (H_2CO_3)** (write both the name and formula). Carbonic acid decomposes into **water** and bubbles of **carbon dioxide**. The bubbles begin to form after the **first** equivalence point. The bubbles will continue until the original **sodium carbonate (Na_2CO_3)** (write both the name and formula) is completely consumed.

Extended Inquiry Suggestions

Titrate amino acid solutions and compare the titration curves to those of other polyprotic acids.

Determine the number of hydrogen ions in the acid ingredient in cola drinks by constructing a titration curve.

Have the students use the half equivalence point to determine the two equilibrium constants in this reaction.

Provide the students with a sample of an unknown solid acid and have them identify the acid as monoprotic (potassium acid phthalate, potassium bitartrate, or benzoic acid), diprotic (oxalic acid or salicylic acid), or triprotic (citric acid). Have them further identify the acid by determining the molar mass of the acid and comparing it to a list of known acids.

28. Le Châtelier's Principle

Objectives

Determine the effect of concentration changes on the equilibrium of a system. Through this investigation, students:

- Use Le Châtelier's principle to explain observed changes and to make predictions
- Relate pH values with the acid-base indicator phenolphthalein

Procedural Overview

Students conduct the following procedures:

- Determine the pH of a variety of solutions
- Predict the shift in equilibrium when concentrations of specific reactants or products are changed

Time Requirement

♦ Preparation time	20 minutes
♦ Pre-lab discussion and activity	30 minutes
♦ Lab activity	40 minutes

Materials and Equipment

For each student or group:

- Data collection system
- pH sensor
- Beaker (2), 100-mL
- Beaker (2), 50-mL
- Graduated cylinder, 25-mL
- Graduated cylinder, 50-mL or 100-mL
- Transfer pipet (3)
- Waste container
- Wash bottle filled with distilled (deionized) water
- Buffer solution pH 4, 25 mL
- Buffer solution pH 10, 25 mL
- Distilled (deionized) water, 100 mL
- Phenolphthalein indicator, 4 drops
- 0.1 M Hydrochloric acid (HCl), 5 mL[1]
- 0.1 M Sodium hydroxide (NaOH), 5 mL[2]
- 0.5 M Acetic acid ($HC_2H_3O_2$), 50 mL[3]
- 0.5 M Sodium acetate ($NaC_2H_3O_2$), 10 mL[4]

[1]To formulate using concentrated (12 M) or dilute (6 M) hydrochloric acid (HCl), refer to the Lab Preparation section.

[2]To formulate using solid sodium hydroxide (NaOH) pellets or dilute (1 M) NaOH, refer to the Lab Preparation section.

[3]To formulate using concentrated (glacial, 17.4 M) acetic acid ($HC_2H_3O_2$) or household vinegar (5% acetic acid), refer to the Lab Preparation section.

[4]To formulate using solid sodium acetate trihydrate ($NaC_2H_3O_2 \cdot 3H_2O$), refer to the Lab Preparation section.

Concepts Students Should Already Know

Students should be familiar with the following concepts:

- Chemical equations
- Acids and bases
- pH
- Equilibrium

Related Labs in This Guide

Labs conceptually related to this one include:

- Evidence of a Chemical Reaction
- pH of Household Chemicals
- An Acid-Base Titration
- Diprotic Titration: Multi-Step Chemical Reactions

Using Your Data Collection System

Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

- ♦ Starting a new experiment on the data collection system �(1.2)
- ♦ Connecting a sensor to the data collection system �(2.1)
- ♦ Calibrating a pH sensor �(3.6)
- ♦ Monitoring live data in a digits display �(6.1)

Background

Equilibrium refers to a dynamic state in which opposing actions occur simultaneously and at the same rate. Both physical processes and chemical processes can reach a state of equilibrium.

An example of a physical process is the liquid-vapor equilibrium in a closed container. Molecules of the liquid evaporate, and at the same time, vapor molecules condense. Eventually equilibrium is reached, and the rate of evaporation equals the rate of condensation. The relative amounts of liquid and vapor inside the container remain constant even though both processes (evaporation and condensation) continue to take place.

The process is similar for a chemical reaction that has reached equilibrium. The rate of the forward reaction equals the rate of the reverse reaction, and the relative amounts of reactants and products remain steady. Because their relative amounts remain constant, the ratio of their concentrations is also constant. The value for this constant ratio can be calculated for the reaction. It is known as the equilibrium constant.

Outside stress can disrupt a system at equilibrium. Changes in temperature, pressure, or concentration of the reactants or products will disturb the equilibrium of the system. Henri Le Châtelier, a French chemist, observed that a system will respond to a stress in such a manner so as to relieve the stress. This phenomenon is known as Le Châtelier's principle.

When a chemical system at equilibrium is disrupted, either the forward or reverse reaction occurs at a greater rate until the system reestablishes a new equilibrium. Using the Haber process as an example, if ammonia gas is synthesized in a closed container, then the following equilibrium results:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{heat.}$$

If the system is stressed by adding more nitrogen gas, then the forward reaction will occur at a faster rate until equilibrium is reestablished. This is described by stating that the “forward reaction is favored” or that the reaction has “shifted to the right.”

Le Châtelier's principle allows chemists to predict how stress factors from outside the system will influence either a physical or chemical system at equilibrium. This information helps chemists manipulate experimental conditions to optimize product formation.

Pre-Lab Discussion and Activity

Equilibrium

Use a simple double pan balance. Alternatively, construct one using a ruler on a fulcrum. Place ten coins on each side, making sure the resulting masses balance the pans. Indicate that the balance is in equilibrium.

Demonstrate dynamic equilibrium by moving one coin from one side of the balance to the other while doing the same with another coin in the opposite direction.

Demonstrate Le Châtelier's principle by disrupting the system, then reestablishing it. First, add six coins to one side. Then move three coins from the heavy side to the light side. Disrupt the equilibrium again by moving four coins from one side to the other. the equilibrium by moving two coins from the heavy side to the light side.

1. What is meant by dynamic equilibrium?

A system in equilibrium is balanced. Dynamic means components of the system are constantly moving. A system in dynamic equilibrium is actively changing while maintaining a balanced outward appearance. This occurs when opposing actions are occurring simultaneously and at the same rate.

Chemical Equilibrium

Write two chemical reactions on the board which can illustrate the difference between a chemical reaction that goes to completion (or very nearly to completion) and a reaction that exists at equilibrium. Consider the following two equations:

Completion	**sodium sulfate + strontium chloride → sodium chloride + strontium sulfate** **$Na_2SO_4(aq) + SrCl_2(aq) \rightarrow 2NaCl(aq) + SrSO_4(s)$**
Equilibrium	**nitrogen + hydrogen ⇌ ammonia (in a closed system)** **$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + heat$ $\Delta H = -92$ kJ/mol**

2. What is chemical equilibrium? How does a chemical reaction that reaches equilibrium differ from a reaction that goes to completion?

Chemical equilibrium means the forward reaction occurs at the same rate as the reverse reaction. A chemical equation that goes to completion only has a forward reaction. A double arrow (⇌) indicates that both forward and reverse reactions occur.

Le Châtelier's Principle

Show students a beaker containing a saturated salt solution. Write the equilibrium equation of a saturated salt solution on the board.

$$Na^+(aq) + Cl^-(aq) \rightleftharpoons NaCl(s)$$

Explain that this system is in dynamic equilibrium. Disturb the equilibrium by adding 50 to 100 mL of 1 M HCl to the beaker of saturated sodium chloride. The added HCl increases the amount of Cl^- present. When a system at equilibrium is disturbed, the reaction shifts in such a manner as to relieve the stress. This observation was made by the French chemist, Henri Le Châtelier, and is called Le Châtelier's principle.

3. The beaker originally contained only a saturated solution of sodium chloride (salt). Macroscopically nothing appeared to be happening, but at the molecular level the particles were undergoing two processes at the same rate. What were these two processes?

The sodium and chloride ions were joining together to form solid sodium chloride crystals (recrystallizing). At the same time and at the same rate, the solid sodium chloride were dissociating to form sodium and chloride ions.

4. Adding hydrochloric acid disrupted the equilibrium of the system. How did the system compensate for the addition? How do you know?

The additional chloride ions from the hydrochloric acid caused the reaction to shift to the right (the forward reaction was favored). This produced more solid sodium chloride, as seen by the cloudiness of the solution.

Acid-Base Equilibrium

This experiment determines the effects of changing concentrations in two different equilibrium systems: the dissociation of phenolphthalein and the dissociation of acetic acid. Both of these substances are weak acids.

Weak acids partially dissociate in water. Explain that phenolphthalein is a weak acid used as an indicator for acid-base reactions. Demonstrate the colors by adding a few drops of phenolphthalein indicator to a two beakers: one with dilute hydrochloric acid and the other with dilute sodium hydroxide. The weak acid (Hphph) is colorless and it's ion ($phph^-$) is pink. Write the equilibrium equation for the dissociation of phenolphthalein on the board.

$$\underset{\text{(colorless)}}{\mathrm{Hphph}} + H_2O \rightleftharpoons H_3O^+ + \underset{\text{(pink)}}{phph^-}$$

Point out that hydrogen ions (H^+) from the acid combine with H_2O molecules to form hydronium ions (H_3O^+). Remind the students that pH is a measure of the H_3O^+ ion concentration and that low values for pH indicate higher concentrations of H_3O^+.

5. What does it mean to partially dissociate?

Dissociate means to break apart into its ions. Partially dissociate means that some of the particles dissociate, but not all of them.

6. What are the two ways to detect changes in the concentration of phenolphthalein (Hphph) compared to its ion ($phph^-$)?

Changes in both color and pH could be used to detect changes in concentration. A colorless solution will have a high pH and will indicate a higher concentration of phenolphthalein as Hphph. A pink solution will have a low pH and indicate a higher concentration of the phenolphthalein ion ($phph^-$) and the hydronium ion.

Lab Preparation

These are the materials and equipment to set up prior to the lab.

Follow these safety procedures as you begin your preparations:

- Wear eye protection, lab apron, and protective gloves when handling acids. Splash-proof goggles are recommended. Either latex or nitrile gloves are suitable.
- If acid solutions come in contact with skin or eyes, rinse immediately with a copious amount of running water for a minimum of 15 minutes.
- Diluting acids creates heat. Be extra careful when handling freshly prepared solutions and glassware because they might be very hot.
- Always add acids to water, not the other way around, because the solutions may boil vigorously.
- Handle concentrated acids in a fume hood; the fumes are caustic and toxic.

Prepare the following solutions:

1. Prepare 100 mL of 0.1 M hydrochloric acid (HCl) from either concentrated (12 M) or dilute (6 M) HCl. This is enough for 20 lab groups.

Starting with concentrated (12 M) HCl:

a. Add approximately 50 mL of distilled water to a 100-mL beaker with a stir bar.

b. Slowly with continuous stirring, add 0.83 mL (measured using a graduated pipet) of 12 M HCl to the water.

c. After the solution cools, carefully pour it into a 100-mL volumetric flask.

d. Dilute to the mark with distilled water.

e. Cap and invert three times to ensure complete mixing.

Starting with 6 M HCl:

a. Add approximately 50 mL of distilled water to a 100-mL beaker with a stir bar.

b. Slowly with continuous stirring, add 1.67 mL (measured using a graduated pipet) of 6 M HCl to the water.

c. After the solution cools, carefully pour it into a 100-mL volumetric flask.

d. Dilute to the mark with distilled water.

e. Cap and invert three times to ensure complete mixing.

2. Prepare 100 mL of 0.1 M sodium hydroxide (NaOH) solution from either solid or1 M NaOH. This is enough for 20 lab groups.

Starting with solid NaOH:

- **a.** Add approximately 50 mL of distilled water to a 100-mL beaker with a stir bar.
- **b.** With continuous stirring, add 0.40 g of NaOH to the water, and allow it to completely dissolve.
- **c.** After the solution cools, carefully pour it into a 100-mL volumetric flask.
- **d.** Dilute to the mark with distilled water.
- **e.** Cap and invert three times to ensure complete mixing.

Starting with 1 M NaOH:

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 10.0 mL of 1 M NaOH to the water
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert three times to ensure complete mixing.

3. Prepare 1000 mL of 0.5 M acetic acid ($HC_2H_3O_2$) solution from either concentrated (glacial, 17.4 M) acetic acid or from household vinegar (5% acetic acid). This is enough for 20 lab groups.

Starting with 17.4 M $HC_2H_3O_2$ (concentrated, also known as glacial acetic acid)

- **a.** Add approximately 500 mL of distilled water to a 1000-mL beaker with a stir bar.
- **b.** Slowly with continuous stirring, add 28.7 mL of 17.4 M $HC_2H_3O_2$ to the water.
- **c.** After the solution cools, carefully pour it into a 1000-mL volumetric flask.
- **d.** Dilute to the mark with distilled water.
- **e.** Cap and invert three times to ensure complete mixing.

Starting with household vinegar (5% $HC_2H_3O_2$)

- **a.** Add 602.4 mL of vinegar (5% $HC_2H_3O_2$) to a 1000-mL volumetric flask.
- **b.** Dilute to the mark with distilled water.
- **c.** Cap and invert three times to ensure complete mixing.

4. Prepare 100 mL of 0.5 M sodium acetate ($HC_2H_3O_2$) solution from solid sodium acetate trihydrate salt ($NaC_2H_3O_2 \cdot 3H_2O$). This is enough for 10 lab groups.

- **a.** Add approximately 50 mL of distilled water to a 100-mL volumetric flask.
- **b.** Add 6.80 g of $NaC_2H_3O_2 \cdot 3H_2O$ to the water and swirl to dissolve completely.
- **c.** Dilute to the mark with distilled water.
- **d.** Cap and invert three times to ensure complete mixing.

Safety

Add these important safety precautions to your normal laboratory procedures:

- ♦ Sodium hydroxide, hydrochloric acid, and acetic acid are corrosive irritants. Avoid contact with the eyes and wash hands after handling.
- ♦ Be sure that all acids and bases are neutralized before being disposed.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

3	1	2	4	5
Use the pH and color of the HCl and NaOH solutions to predict how to make the color of the current solution change.	Add 5 drops phenolphthalein to two separate beakers of distilled water. Record the pH and color of each solution.	Add 5 drops of HCl to one beaker and 5 drops of NaOH to the other beaker. Record the pH and color of each solution.	Test your prediction and use the results to predict how the pH of an acetic acid solution will change when sodium acetate is added.	Test your prediction by measuring the pH of an acetic acid solution before and after sodium acetate is added.

Procedure with Inquiry

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: Students use the following technical procedures in this activity. The instructions for them (identified by the number following the symbol: "�") are on the storage device that accompanies this manual. Choose the file that corresponds to your PASCO data collection system. Please make copies of these instructions available for your students.

Part 1 - Phenolphthalein

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a pH sensor to the data collection system. �(2.1)

3. ☐ Place 25 mL of pH 4 buffer solution in a 50-mL beaker and 25 mL of pH 10 buffer solutions in a second 50-mL beaker. Use these solutions to calibrate the pH sensor. �(3.6)

4. ☐ Label a 100-mL beaker as "A" and add approximately 50 mL of distilled water. Repeat for beaker "B".

5. ☐ Add 1 to 2 drops of phenolphthalein indicator to the distilled water in each beaker. Record the color of the solution in each beaker after adding the phenolphthalein in Table 1 below (in the Collect Data section).

Note: Phenolphthalein is a weak acid and is used as an indicator for acid-base reactions. The weak acid is colorless (Hphph) and its ion is pink ($phph^-$). In aqueous solutions, the two forms are in equilibrium according to the equation below.

$$\underset{\text{(colorless)}}{\text{Hphph}} + H_2O \rightleftharpoons H_3O^+ + \underset{\text{(pink)}}{phph^-}$$

6. ☐ Does the equilibrium of phenolphthalein in water favor the products or reactants? How do you know?

The equilibrium favors the reactants, because the solution is colorless.

7. ☐ Predict what will happen when hydronium ions (H_3O^+) are *removed* from the solution. Explain your prediction.

The solution will turn pink, because the weak acid will dissociate to replace the H_3O+ that has been removed. This creates a larger amount of the $phph^-$ which is pink.

8. ☐ How might hydronium ions be removed from the solution?

By adding a base to the solution, the reaction will neutralize the hydronium ions and produce water.

$H_3O^+ + OH^- \rightarrow 2H_2O$

9. ☐ Predict what will happen when hydronium ions are *added* to the solution. Explain your prediction.

The solution will remain colorless. This is because ions of phenolphthalein (phph–) react with H3O+ ions, which causes the amount of H3O+ ions in the solution to decrease. This creates more Hphph which is colorless.

10. ☐ How might hydronium ions be added?

Simply add an acid, such as HCl to the solution.

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Collect Data

11. ☐ Monitor live pH data in a digits display. �(6.1)

12. ☐ Place the pH sensor in beaker A, allow the pH reading to stabilize, and then record the pH of the solution in Table 1 below.

Table 1: Color and pH with phenolpthalein

Equilibrium Conditions	Color	pH
Beaker A: distilled water with phenolphthalein	Colorless	7.0
Beaker B: distilled water with phenolphthalein	Colorless	7.0
Beaker A: after adding 5 drops of HCl	Colorless	4.2
Beaker B: after adding 5 drops of NaOH	Pink	10.7

13. ☐ Rinse the pH sensor with distilled water.

14. ☐ Repeat for beaker B: place the pH sensor in the solution, allow the reading to stabilize, record in the pH value in Table 1; and rinse the pH sensor.

15. ☐ Add 5 drops of 0.1 M HCl to beaker A and record the color in Table 1 above.

16. ☐ Measure the pH in beaker A and record the pH value in Table 1 above.

17. ☐ Rinse the pH sensor with distilled water.

18. ☐ Add 5 drops of 0.1 M NaOH to beaker B.

19. ☐ Measure the pH in beaker B and record both the color and the pH in Table 1 above.

20. ☐ Remove the pH senor from solution and rinse the sensor with distilled water.

21. ☐ What color is the solution in beaker A right now? If the solution is pink, what can you do to make it turn clear? If the solution is clear what can you do to make it turn pink? Why will this work?

The solution in beaker A is clear. Adding NaOH will shift the equilibrium toward the products by removing one of the products (H_3O^+) and will cause the solution to turn pink.

22. ☐ What color is the solution in beaker B right now? If the solution is pink, what can you do to make it turn clear? If the solution is clear what can you do to make it turn pink? Why will this work?

The solution in beaker B is pink. Adding HCl will shift the equilibrium toward the reactants by increasing the concentration of one of the products (H_3O^+) and will cause the solution to turn clear.

23. ☐ Test the predictions you made. Record exactly what you added to each solution, whether or not the color changed, and the final pH of the solution after you made the change in Table 2 below.

Table 2: Color and pH after shifting equilibrium with NaOH and HCl

Equilibrium Conditions	Action Taken	Color Change?	pH
Beaker A with phenolphthalein + HCl	Added 5 drops NaOH	turned pink	9.6
Beaker B with phenolphthalein + NaOH	Added 22 drops HCl	turned colorless	3.6

24. ☐ Dispose of the solutions according to the teacher's instructions.

25. ☐ Wash the equipment that you used so that it can be reused in the next section.

Part 2 – Acetic Acid

Set Up

26. ☐ Add 50.0 mL of 0.5 M acetic acid solution to a 100-mL beaker.

27. ☐ As indicated by the following equilibrium system, will the pH of the 0.5 M acetic acid solution be greater than, less than, or equal to 7? Explain your reasoning.

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

Note: This represents the dissociation of acetic acid in water.

The pH should be less than 7 because the solution is acidic. The solution is acidic because it contains hydronium ions.

28. ☐ Measure 10.0 mL of 0.5 M sodium acetate in a graduated cylinder.

29. ☐ What will happen to the concentration of acetate ions when sodium acetate is added to the acetic acid solution? Hint: Sodium acetate dissociates completely in water according to the chemical equation below.

$$NaC_2H_3O_2(s) \rightarrow Na^+(aq) + C_2H_3O_2^-(aq)$$

When sodium acetate is added to the solution, the number of acetate ions will increase.

30. ☐ What will happen to the pH of the acetic acid solution when sodium acetate is added? Explain your prediction.

The reaction will favor the reactants (shift to the left) to reestablish equilibrium. To reach a new equilibrium, the number of H_3O^+ ions will decrease as they are used to consume added acetate ions. This results in an increase of the pH.

Collect Data

31. ☐ Place the pH sensor in the beaker containing acetic acid.

32. ☐ Allow the pH reading to stabilize, and then record the value in Table 3 below.

Table 3: pH for acetic acid and for sodium acetate

Equilibrium Conditions	pH
0.5 M acetic acid	3.3
0.5 M acetic acid + 0.5 M sodium acetate	4.7

33. ☐ Pour the 10 mL of 0.5 M sodium acetate into the beaker containing acetic acid and stir gently with the pH sensor.

34. ☐ Allow the pH reading to stabilize, and then record the value in Table 3 above.

35. ☐ Clean the lab station according to the teacher's instructions, especially those for the disposal of the solutions.

Analysis Questions

1. How were the equilibria of the solutions in the experiment disturbed?

The equilibria were upset by changing the concentrations of various products. The concentrations were changed by either adding a product or by using a chemical reaction to remove some of a product.

2. Why was pH data recorded?

pH data was recorded to monitor changes in the amount of H_3O^+ ions as the equilibrium of a system was disturbed. pH is a measure of H_3O^+ ion concentration in a solution and therefore allow us to “see” what is happening at the molecular level.

3. What are two different ways H_3O^+ concentrations were detected in this experiment?

A pH sensor gave numerical results, while the use of an indicator (phenolphthalein) changed color.

Synthesis Questions

Use available resources to help you answer the following questions.

1. Write the chemical equation for the equilibrium established in a saturated sodium chloride (NaCl) solution. Do not include water in the equation, but be sure to include the appropriate state symbols.

$NaCl(s) \rightleftharpoons Na^{+}(aq) + Cl^{-}(aq)$

2. What two ways could you cause NaCl to precipitate out of a saturated solution without adding more NaCl? Explain your answer?

Add more Na^{+} ions or Cl^{-} ions to the solution. This can be done using compounds such as NaOH or HCl that will dissociate. This works because the reaction shifts to the products in order to alleviate the disturbance. The newly formed NaCl precipitates out because the solution is already saturated.

3. What are three variables of a system that can be changed to disrupt its equilibrium?

Changing the temperature, pressure, or concentration of either a reactant or product can disrupt a system in equilibrium.

4. In the following system at equilibrium, explain how cooling the system shifts the equilibrium.

$$C(s) + H_2O(g) + heat \rightleftharpoons CO(g) + H_2(g)$$

Cooling the reaction shifts the equilibrium to the left (favoring the reactants).

Multiple Choice Questions

Select the best answer or completion to each of the questions or incomplete statements below.

1. Le Châtelier's principle states that

- **A.** A chemical reaction always goes to completion.
- **B.** A system is at equilibrium when the concentration of products equals the concentration of the reactants.
- **C.** A system favors the products when a stress is placed on it.
- **D.** **A system responds to a stress in such a manner that it relieves the stress.**

Use the following equilibrium equation to answer Multiple Choice Questions 2 and 3 below.

$$Fe^{+3}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

(colorless) + (colorless) (brown/deep red)

2. What happens to the color of the solution when you add $Fe(NO_3)_3$ to it?

- **A. It turns darker because the reaction shifts to make more products.**
- **B.** It turns darker because the reaction shifts to make more reactants.
- **C.** It turns lighter because the reaction shifts to make more reactants.
- **D.** It turns lighter because the reaction shifts to make more products.

3. What will happen when SCN^{-} ions are removed from the solution?

- **A.** More products will form.
- **B. More reactants will form.**
- **C.** More reactants and products will form.
- **D.** Nothing will happen.

4. How might the equilibrium be shifted towards the products in the following reaction?

$$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)$$

- **A. Add more O_2.**
- **B.** Add more CO_2.
- **C.** Remove CO.
- **D.** Remove O_2.

5. Which change may disrupt a system in equilibrium?

- **A.** A change in the pressure of the system (if gases are present).
- **B.** A change in the temperature of the system.
- **C.** A change in the concentration of the products or reactants.
- **D. All of the above.**

Key Term Challenge

Fill in the blanks from the list of randomly ordered words in the Key Term Challenge Word Bank.

1. A **forward reaction** is one in which reactants are converted to products. A **reverse reaction** is also possible in which the newly created products revert back to the original reactants. Many chemical reactions can proceed in both directions at the same time. Shortly after the forward reaction begins, the reverse reaction also takes place such that both are occurring at the same time. Eventually, a **chemical equilibrium** is established where the rate of the forward reaction equals the rate of the reverse reaction. The equation recognizes the dual process with opposite arrows ($\rightleftharpoons$). Even though both the forward and reverse reactions continue to happen, the relative amounts of reactants and products are **the same**. This establishes a **dynamic** equilibrium where it only appears that nothing is changing. Common types of equilibrium reactions include the **dissociation** of weak acids or weak bases in solution, many types of redox reactions, and the dissociation of slightly soluble salts in solution.

2. A system at equilibrium can be disturbed by changing the **amount** of reactants or products, the **temperature** of the system, or the **pressure** of the system (if gaseous substances are involved). The system will then adjust in order to reestablish a new equilibrium between its reactants and products. **Le Châtelier's principle** states that a system at equilibrium responds to a change in such a manner that it relieves the stress. For example, a weak acid partially dissociates in aqueous solutions until it reaches equilibrium. If additional ions are **added** the weak acid will reform. If ions are **removed**, however, the weak acid will dissociate to produce more ions.

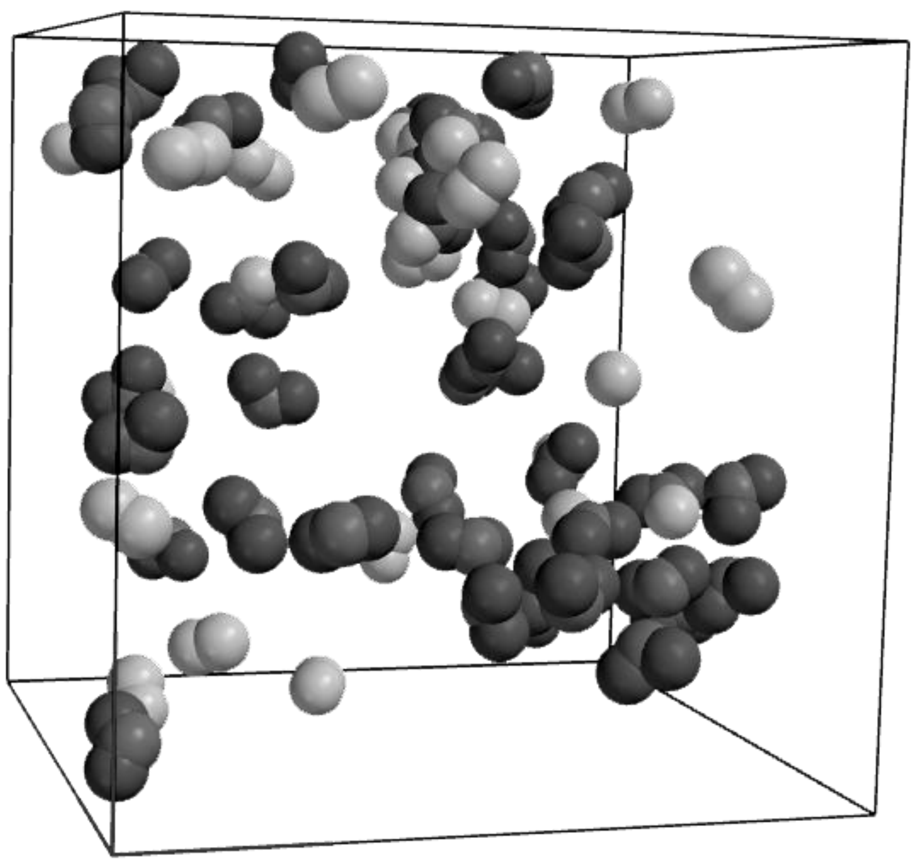
Extended Inquiry Suggestions

Find the pH range of other indicators, such as methyl orange.

Repeat the experiment using a different indicator, such as methyl orange.

Investigate the physical equilibrium of phase changes.

ODYSSEY Molecular Labs



ODYSSEY Molecular Labs

This section contains the student worksheets that accompany the Molecular Lab simulations. The simulations and the answer key can be found on the *ODYSSEY* storage device that came with this manual.

Introduction

Using ODYSSEY Molecular Labs

Wavefunction's *ODYSSEY* is a unique software program for use in chemistry classes. Maker of *Spartan* molecular modeling software, Wavefunction created *ODYSSEY* so students could use scientifically-based simulations to experiment with core chemistry topics from a molecular perspective. The additional perspective provided by the *ODYSSEY* software enhances and complements the hands-on, experiential PASCO activities in this manual.

ODYSSEY includes a collection of ready-to-use chemistry experiments (called “Molecular Labs”) and student worksheets. A number of the Molecular Labs applicable to the PASCO activities are identified under selected topic areas in the table of contents. The student worksheets for these labs are provided in the *ODYSSEY* Molecular Labs section of this manual and the answer key can be found on the accompanying storage device. A fully functional 60-day licensed version of the *ODYSSEY* Instructor Edition, containing the complete set of *ODYSSEY’s* Molecular Labs, is included with this manual.

In addition to the Molecular Labs, *ODYSSEY* provides:

- Prelabs (Tutorials) – ideal for learning about how to use the program
- Applied Chemistry – a collection of chemistry samples commonly encountered in modern society
- Molecular Stockroom – the electronic equivalent of your chemistry stockroom with more than a thousand pre-constructed samples spanning the periodic table

To successfully get started with *ODYSSEY,* check the system requirements and install the software that is on the accompanying *ODYSSEY* storage device; use the activation code provided to access the software for 60 days. Contact PASCO (www.pasco.com) for information on instructor and student licensing.

Classifying Chemical and Physical Properties

Objectives

- To distinguish between the chemical and physical properties of ammonia.
- To classify select properties of ammonia as either extensive or intensive.
- To consider the conversion between the extensive and intensive properties.

Concepts

Chemical properties relate to change on the molecular level that leads to *new* substances. Such change can occur via interaction with other substances:

- Flammability (an interaction with oxygen)
- Susceptibility to oxidation ("rusting")
- Reaction (or inertness) with strong acids

Or there can be chemical change of the substance itself:

- Decomposition of the substance
- Rearrangement of the molecules in a process called isomerization

The important thing to remember is that chemical properties can only be observed when a substance changes into something new. In other words, you can't just look at a substance and know its chemical properties.

All properties of a substance that do not involve chemical reactions constitute its ***physical*** properties. Examples include:

- Boiling point
- Mass
- Index of refraction

The physical properties again can be classified as either extensive or intensive:

- *Extensive* = Property scales with the system size
- *Intensive* = Property does not scale with system size

Procedure and Analysis

1. Label the following properties of dense gaseous ammonia as chemical or physical properties:
   a. Ammonia has a density of 0.028 g/cm^3 at ambient temperature and pressure of ~40 atm..

   b. Ammonia forms a compound with hydrogen chloride that is a white crystalline solid and is 26.2% nitrogen, 7.5% hydrogen and 66.3% chlorine by mass.

   c. At ambient pressure ammonia condenses into a liquid at a temperature of -78 °C.

   d. In the liquid state, ammonia has a specific heat capacity of 80.8 J/(mol·K).

   e. Ammonia does not form any stable chemical compounds with xenon gas.

2. Three samples of dense gaseous ammonia are provided that are chemically identical, but differ in the system size:

Ammonia: ○ **Size I** ○ **Size II** ○ **Size III**

Simulations are required to answer some of the subsequent questions, thus click on the Start button below the models:

❑ *When Running, Autostart After Sample Change*

- From the **View** menu, select **Properties**.
- From the **Add Property** menu (lower left corner), select **System** → **Total Number of Molecules**.
- One after the other, select samples I, II, and III above—the value of the property immediately gets updated in the table.

Do the total number of molecules scale with the system size, i.e., is the property extensive or intensive?

3. One by one, add the following to the list of "Properties" and for each determine if its value changes (within the margin of error for the simulations) when you switch between samples I, II, and III:

   a. **System → Density**

   b. **Thermodynamics → Volume**

   c. **Energy → Total Energy (Time Averaged)**

   d. **Thermodynamics → Temperature**

   e. **Energy → Kinetic Energy (Time Averaged)**

   f. **Thermodynamics → Enthalpy (Molar)**

   g. **Energy → Potential Energy (Time Averaged)**

   h. **Thermodynamics → Entropy (Molar)**

   i. **Thermodynamics → Free Energy (Molar)**

   j. **Dynamics → Speed**

   k. **Thermodynamics → Pressure**

   l. **Other → PV/nRT** ("Compression Factor")

   Which of the properties do you find to be extensive?

4. Which of the properties do you find to be intensive?

5. Generally, you cannot just tell from the name whether the property is extensive or intensive. However, there is one exception in that a certain label will immediately tell you that a property is extensive. What is this label?

6. Extensive properties can always be converted to intensive properties by dividing by the system size. For example "food intake" is an extensive property, and "food intake per pound of bodyweight" is its intensive equivalent. Give an example from the extensive properties that you identified above together with its intensive counterpart.

7. Intensive properties can often be converted to extensive properties by multiplying with the system size. For two of the intensive properties that you identified above, give their extensive counterparts.

8. While it works for molar properties, the conversion from an intensive property to an extensive property by multiplying with the system size does *not* always make sense. For example, if you take the intensive property "pulse rate", then "pulse rate times pounds of bodyweight" is not a very meaningful concept. Which of the properties that you identified as intensive do not have an extensive counterpart?

9. Suppose you need to identify some unknown substance by making a *single* measurement using a small amount of the substance. Would the measured property have to be extensive or intensive? Or would the answer depend on the type of substance?

10. Which of the following would be an example of a chemical property?
    a. Water boils at 100 °C
    b. Silver is malleable
    c. Methane will combust
    d. The density of copper is 8.96 g/cm^3

11. Which of the following is not a physical property?
    a. Ice melting
    b. A car rusting
    c. Rubbing alcohol evaporating
    d. Liquid bromine is a red-brown color

Identifying a substance by Its Density

Objectives

- To describe the relationship between mass and volume for samples of a given substance that vary in the size and shape.
- To identify a substance via its density value

Models

The following models represent molecular-sized "chunks" of an unknown metal:

○Sample 1 ○Sample 3 ○ Sample 5
○Sample 2 ○Sample 4 ○Sample 6

A mass-volume plot for the six samples can be generated in the following way:

- Click on **Plots** in the toolbar (or select the menu item **View** → **Plots**). From the **Add Plot** menu (upper left corner), select **XY Plot....**
- For the "X Axis," select **Volume**. For the "Y Axis," select **Total Mass**.
- Click **Advanced** and select **Different samples are datapoints in a single curve**. Click **OK**.
- Click **Next** and select **Linear Fit**. Click **Finish**.
- Note that you can "drag" the caption with the fit equation (so that it won't interfere with the graph).

Questions

1. What is the relationship between mass and volume?

2. What parameter of the fit line gives you the density value?

3. Directly from the fit equation the density is obtained in units of g/nm^3. What is the conversion factor to the more common density unit of g/cm^3?

4. What is the density value (in g/cm^3) that you obtain?.

5. What substance so you conclude that the molecular models represent?

6. The density can also be obtained from measurements of mass and volume of individual samples:

   - Click on **Properties** in the toolbar (or select the menu **View → Properties**).
   - From the **Add Properties** menu (lower left corner) select **System → Total Mass** and (if it is not yet displayed) also **Thermodynamics → Volume.**

   Obtain an individual density value for the *most* elongated of the six samples by taking the mass to volume ratio. Repeat for the least elongated sample. What is the apparent relationship between sample *shape* and density?

7. What is the apparent relationship between sample *size* and density?

8. The fact that the density can be obtained from the mass to volume ratio of just a single sample is crucially dependent on the fact that the fit line in the mass-volume plot goes through a very special point::

   - Bring up the **Plot Edit** dialog (above the plot, click on the **Pencil** icon).
   - In the”X-Axis” section, set the checkmark for **User Defined Range** and change the “From” filed to **0.**
   - Do the same for the “Y Axis” section, then click **OK.**

   What is the special point that the graph goes through?

9. Your calculated fit equation presumably shows that the graph has a intercept that is *not* exactly zero. How can the intercept be non-zero and the graph still do through the special point that you identified in the previous question?

10. What could be some reasons for finding a non-zero intercept?

Comparing s- and p-Orbitals

Objectives

In this experiment, you will consider the atoms of the three lightest noble gases:

⦿ **Helium / Neon / Argon** He / Ne / Ar

Specifically, you will study the *internal* electronic structure of these atoms. You will

- characterize the orbitals for different quantum numbers,
- examine and compare a variety of occupied orbitals,
- elucidate the effect of nuclear charge on orbital size.

Procedure and Analysis

11. Closely related to the notation for electron configurations, atomic orbitals have unique labels, such as "$3p_x$". The first item of the label, the so-called "principal quantum number", relates to the orbital's size. Three orbitals only differing in that quantum number are shown:

    Argon: ○ **1s** ○ **2s** ○ **3s**

    What is the correlation between principal quantum number and orbital size?

12. The second item of an orbital label relates to the orbital's shape. Two orbitals differing in the shape are shown:

    Neon: ○ **2s Orbital** ○ **2p Orbital**

    What is the characteristic shape of an s-orbital?

13. What is the characteristic shape of a p-orbital?

14. The third (subscripted) item of the orbital label relates to the orbital's orientation. An example is shown:

Argon: ○ **$3p_y$** ○ **$3p_z$**

Why is there no subscript for s-orbitals?

15. How many electrons (total) do we expect an atom of argon to have? Bring up the periodic table (**Tools → Periodic Table**). What item in the periodic table provides the needed information?

16. How many *orbitals* total do we expect to be occupied in an atom of argon?

17. The *highest* principal quantum number for the occupied orbitals of an argon atom is "3". How does that number relate to the element's position in the periodic table? (See **Tools → Periodic Table**.)

18. What is the relationship between the nuclear charge and the number of electrons of an argon atom?

19. Compare the size of the 1s orbital for the first three noble gases:

1s: ○ **Helium** ○ **Neon** ○ **Argon**

Even though an atom neon is clearly heavier and bigger than an atom of helium, its 1s orbital is much smaller! What physical reason can explain this at first glance puzzling behavior?

20. Again examine the last set of models: The size difference between helium and neon is bigger than the size difference between neon and argon. Try to rationalize this observation (think of the *ratio* of the nuclear charges for the two cases).

21. Next, compare the 2s orbitals of neon and argon (why is helium not included in the comparison?)

    2s: ○ **Neon** ○ **Argon**

    Is the variation qualitatively the same as for the 1s orbital?

22. Models for the three degenerate (equal-energy) 2p orbitals of neon are available:

    Neon: ○ **$2p_x$** ○ **$2p_y$** ○ **$2p_z$**

    The same set of orbitals also exists for argon, only the orbitals are smaller:

    Argon: ○ **$2p_x$** ○ **$2p_y$** ○ **$2p_z$**

    What is true about the mutual orientation of the p-orbitals in either case?

23. Some of the preceding questions elaborated on the fact that a given orbital gets smaller as you proceed from a lighter atom to a heavier atom. If that is true, why are the atoms of the heavier noble gases *still* bigger than the atoms of the lighter ones?

24. Models **I** to **IV** show, in random order, the complete set of *valence* orbitals of argon:

    Valence Orbitals: ○ **I** ○ **II** ○ **III** ○ **IV**

    Assign the correct label to each of the four orbitals.

Electronegativity and the Formation of Bonds

Objectives

The degree of ionic and covalent bonding can often be estimated by simply looking at the difference in electronegativity (EN) between the bonded atoms. In this experiment, you will apply this technique to 18 dimers that show a systematic variation in the electronegativity difference between the atoms.

Procedure and Analysis

The dimers in this experiment are all shown with a connecting "bond"—not only if the bonding is covalent in character (in which case it is "normal" to draw an explicit bond), but also if it is ionic in character (for which we usually do *not* draw explicit bonds).

1. From the **View** menu, select **Properties**. From the **Add Properties** menu (lower left corner), select **Atom → Electronegativity.** Find the electronegativity for both atoms and enter the *difference* in the following table:

	ΔEN	Bonding Type
⊙ **Cl_2**		
○ **BrCl**		
○ **IBr**		
○ **HI**		
○ **ICl**		
○ **HBr**		
○ **HCl**		
○ **BrF**		
○ **LiI**		
○ **NaI**		
○ **KI**		
○ **LiBr**		
○ **NaBr**		
○ **KBr**		
○ **NaCl**		
○ **KCl**		
○ **NaF**		
○ **KF**		

2. Of the elements appearing in these diatomics, which one is the most electronegative and which one is the most electropositive?

3. To which corner of the periodic table would you have to go to find *the* most electropositive element?

4. Display the electrostatic potential maps of the diatomic molecules:

   ○ **Hide** ○ **Mesh** ○ **Transparent** ○ **Solid**

   Blue indicates that the area interacts as if there was *effectively* positive charge and red indicates that the area interacts as if there was *effectively* negative charge; green is the neutral color.

   Using your subjective judgment, classify the bonding type as either "covalent," "polar covalent," or "ionic" in each case (fill in the table above).

5. At what electronegativity difference would you draw the line between ionic and covalent bonding?

6. Do you think that your answers to the last two questions are unique? If yes, explain why. If no, also explain why.

7. Hydrogen fluoride deviates somewhat from "average" behavior with regard to the degree of covalent bonding. Given that hydrogen fluoride is often classified as "polar covalent," is the molecule more ionic or more covalent than one might have expected?

Naming Molecular Compounds

Objectives

Some chemical compounds are practically always referred to by their common name (or "trivial name"), for example ammonia. Common names must be memorized—there is no way around it. In the vast majority of cases, however, knowing a common name is not really required to be able to uniquely identify a compound while still being easily understood. In the following, you will

- consider the rules for giving systematic names to binary molecular compounds,
- apply the rules to the molecules of a number of unknowns.

Concepts

Compounds of covalently bonded molecules with just two types of atoms are not difficult to name. In the most widely used system, the rules are as follows:

- Start the compound name with the element that appears furthest left in the periodic table. If both elements are in the same group, then start with the one from the lower period.
- Make the name of the second element end in “-ide".
- Use numerical *prefixes* to indicate the number of atoms in the molecule:

# of Atoms	Prefix	# of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

Exception—don’t use “mono” for the first element in the compound name.

- Drop the "o" or "a" at the end of the prefixes if the element name begins with a vowel.

Questions

1. Name the molecular compounds listed below using the rules that you have learned:

⊙ **Molecule #1**

○ **Molecule #2**

○ **Molecule #3**

- **Molecule #4**

- **Molecule #5**

- **Molecule #6**

- **Molecule #7**

- **Molecule #8**

- **Molecule #9**

- **Molecule #10**

- **Molecule #11**

- **Molecule #12**

- **Molecule #13**

- **Molecule #14**

2. The first rule above could also have been phrased in terms of electronegativity. Are electropositive or electronegative elements more likely to be listed first in the systematic name?

3. Conversely, are electropositive or electronegative elements more likely to be listed *last* in the systematic name?

4. Look through the elements in all the models that you have considered. Is the first element in binary molecular compounds typically a metal, nonmetal, or semimetal?

5. In continuation of the previous question, is the second element typically a metal, nonmetal, or semimetal?

6. One of the questions above related the required sequence of element names to the electronegativity concept. Can you find an example among the 14 molecules listed where this type of “mapping” does not work? If yes, name the molecule.

Exploring Ionic Attraction

Objectives

- To investigate how charged atoms ("ions") attract and repel each other.
- To correlate the strength of the forces with the magnitude of the charges.

Introduction

An *ion* is an atom that has become charged:

- A positive ion forms if an atom loses one (or multiple electrons).
- A negative ion forms if an atom gains one or sometimes multiple electrons.

How will ions interact once they have formed? In this experiment, you will use a simple model setup of molecule-sized moving balls to examine the interactions between particles that are either uncharged (atoms) or charged (ions). By comparing different simulations, you will get a feel for the driving force of the formation of ionic compounds.

Procedure

The model on the left shows a system of 10 red atoms ("A") and 10 blue atoms ("B") confined to a container:

⊙ **Model System for Studying Ionic Interactions**

All particles are initially uncharged.

- When you start the simulation (below the model, click on the **Start** Button), you can follow the dynamics of the system. Atoms collide randomly with each other as well as with the walls of the container.

You can associate electric charges with the atoms, either for just one or for both of the species. The needed control is in the **Build** menu and is called **Charges**....

- Change the charge of both A atoms and B atoms to **+**1. Click OK and, after restarting, observe the simulation.

With the charge of one of the two types of ion still at +1, change the charge of the other type of ion to –1. Again observe the simulation.

Change the charges of A and B ions to either both **+5** or both **–5**. Observe the simulation.

- Try a permutation where the charge of one type of particle is +1 (or –1) and the other type of particle is uncharged.

Note that you can display charge labels: From the **Style** menu, select first **Ball and Spoke** and subsequently **Charge Labels → Net (Ionic)**.

Questions

1. Describe your observations if the charge of both species is **+1**.

2. Compare your observations for the case where both charges are **+1** with the case where both charges are **-1**. Is there a difference?

3. Describe your observations if the charges of two species are **+1** and **-1**, respectively, i.e., if they are opposite.

4. What changes if both charges are **+5** (or **–5**)?

5. Describe your observations for the case that only one of the two atom types is charged.

6. How do ions of like charge act on each other?

7. How do ions of unlike charges act on each other?

8. In general terms, how does the strength of the electrostatic interactions very with the charge?

9. Examining the following molecular model of sodium chloride:

⊙ **Sodium Chloride** NaCl(s)

The model is closely related to the abstract model studied above. What are the ions in sodium chloride?

10. In what sense does the model of sodium chloride differ from the abstract model?

Comparing the Density of Liquids and Gases

Objectives

Compared to the density of liquids and solids, the density of gases is exceedingly low. In this experiment, you will quantify exactly how low.

Procedure and Analysis

The following molecular samples represent a typical gas and a typical liquid, both at room temperature (*T* = 25 °C) and ambient pressure (*P* ~1 atm):

⊙ **Carbon Monoxide** CO (*g*)
○ **Methanol** CH_3OH (*l*)

1. Start with the sample of carbon monoxide. In the "Space Filling" model style, all atoms are shown in a size that is physically “realistic”:

   Model Style: ⊙ **Space Filling** ○ **Ball and Spoke**

   Clearly the molecules take up only a fraction of the available space. Start to make this statement more precise by selecting **View → Properties**. From the **Add Property** menu (lower left corner), select **Molecule → Molecular Volume**. Clicking on any one of the molecules will provide you with a rough estimate for the molecular volume of a single molecule.

2. Next, select **System → Total Number of Molecules**. This gives you the total number of molecules in the sample.

3. Finally, find the total system volume (= "box size") by selecting **Thermodynamics → Volume**.

4. What fraction of space (in %) is taken up by the molecules?

5. Repeat all three measurements for the molecular sample of methanol, i.e., for a substance that is a *liquid* at room temperature and ambient pressure.

6. What fraction of space (in %) is taken up by the molecules in the liquid?

7. By what factor do the two fractions differ, i.e., approximately how much "denser" is methanol than carbon monoxide?

Bonding in Crystalline Solids

Objectives

An analysis of the binding forces in crystals suggests to distinguish between four major types of solids: ionic, network-covalent, molecular, and metallic. In this experiment, you will

- examine representative models for the four major types,
- correlate structural features with macroscopic properties.

Concepts

The crystals of *ionic* solids are hard and brittle and melt at high temperatures:

- ⊙ **Cesium Iodide** CsI (*s*)
- ○ **Calcium Fluoride (Fluorite)** CaF_2 (*s*)
- ○ **Magnesium Oxide** MgO (*s*)
- ○ **Ammonium Chloride** NH_4Cl (*s*)

Network-covalent solids are even harder than ionic solids and tend to melt at exceedingly high temperatures:

- ○ **Diamond** C (*s*)
- ○ **Silicon Dioxide** (Quartz) SiO_2 (*s*)
- ○ **Boron** B (*s*)

Molecular solids are the other extreme—they tend to be fairly soft and have moderate melting points:

- ○ **Iodine** I_2 (*s*)
- ○ **Solid Water** (Ice) H_2O (*s*)
- ○ **Glucose** $C_6H_{12}O_6$ (*s*)

Metallic solids are not very uniform in that there is considerable variation with regard to hardness and melting points:

- ○ **Magnesium** Mg (*s*)
- ○ **Gold** Au (*s*)
- ○ **Tungsten** W (*s*)

What sets metallic solids apart from each of the other types of solids is that they exhibit excellent electrical as well as thermal conductivity.

Analysis

1. What group of solids is held together by electrostatic attractions between positive and negative ions?

   Charge Labels: ⦿ **Hide** ○ **Show**

2. What type of solid structure amounts to the presence of "giant molecules"?

3. Which of the four types of solids include examples where there is some covalent bonding?

4. Which of the four types of solids cannot possibly occur as the structure type of an element?

5. The melting points of cesium iodide, fluorite and magnesium oxide are 621 °C, 1,402 °C, and 2,852 °C, respectively:

   ⦿ **CsI** (*s*) ○ **CaF_2** (*s*) ○ **MgO** (*s*)

   What could be a simple physical explanation for the variation of the melting points?

6. The binding forces between the particles of molecular solids are "Van der Waals forces":

   ○ **I_2** (*s*) ○ **H_2O** (*s*) ○ **$C_6H_{12}O_6$** (*s*)

   Two of the molecular solids shown include a particularly *strong* form of Van der Waals force. What force is that and what are the models affected by it?

7. The metallic solids are "densely packed" in that each particle has a large number of immediate neighbors:

   ○ **Mg** (*s*) ○ **Au** (*s*) ○ **W** (*s*)

   - Change the model style of the selected structure to **Model → Ball and Spoke**.
   - *Right-click* on an atom. Select **Set Clipping Center**.
   - Use the scrollwheel to display the clipping sphere screen-filling.
   - Now *select* the clipping sphere by clicking on it → its color changes to golden. Use the scrollwheel to adjust the size of the clipping sphere so that you can see the shell of *nearest* neighbors of the chosen atom. Click on the background to deselect the clipping sphere.

   What is the number of direct (nearest) neighbors for the three metallic structures shown here?

8. Compare the "cohesive energies" of the three metallic substances (those are the energies *relative* to the separated atoms in the gas phase):

   ○ **Mg** (*s*) ○ **Au** (*s*) ○ **W** (*s*)

   - From the **View** menu, select **Properties**.
   - From the **Properties** menu (lower left corner), select **Energy → Potential Energy**.

   What values do you find and what do you *predict* for the relative melting temperatures of the three metals?

9. The crystal structures of noble gases look very much like that of metals:

   Examples: ○ **Argon** Ar (*s*) ○ **Xenon** Xe (*s*)

   However, the bonding is not metallic, but rather that of weak dispersion forces. Accordingly, noble gas solids melt at very low temperatures (at –189 °C and –112 °C for argon and xenon, respectively). Given these properties, what is the best classification for this type of atomic solid?

10. The solids shown in this experiment already include several substances that we are familiar with from everyday life. For each of the four major categories, give one more example of a "common" substance that represents that category.

Measuring Gas Pressure

Objectives

The model on the left shows a single helium atom confined to a simulation box:

⦿ **Confined Helium Atom**

When you start the simulation (below the model, click on the **Start** button), you can see how the atom bounces off the container walls—note how the computer draws a small trail to highlight the motion as well as a little "halo" during the collision events themselves.

Each collision imparts a tiny *force* on of the walls of the confining box. If there are many atoms and many such collisions, the total force gives rise to the pressure of the system. In this experiment, you will

- interpret the pressure of a sample of helium in terms of molecular collision events,
- express the pressure value in various units,
- examine how the pressure correlates with the number of particles in the sample

Procedure and Analysis

1. The macroscopic definition of pressure is:

   Pressure = Force / Area

   What are the SI units of "force" and "area"? What is the SI unit of pressure that is obtained by taking the ratio of the SI units for force and area?

2. For buoyancy, balloons are commonly filled with a very light gas:

   ○ **Helium**

   Start the simulation, then after a few seconds stop it again. While stopped, count the number of "halos" indicating collisions with the container walls—the *forces* associated with the collisions are the microscopic origin of the gas pressure. Repeat this four times by restarting, stopping, and counting the number of halos. Give the average number of "simultaneous collisions" (collisions at a given moment) that you determine in this manner.

3. Find the dimensions of the simulation cell that are indicated above the sample: From the **View** menu, select **Properties**, and from the Properties menu (lower left corner), select **System → Cell Dimensions.** From the dimensions, calculate the simulation cell's surface area (total of the six surrounding walls).

4. Consider a typical party balloon with a surface area of ~3000 cm^2. What keeps the balloon inflated is the continuous "bumping" of the gas particles into its walls. If the density of helium in the balloon is the same as in the simulation sample, what do you extrapolate from the simulation result for the number of "simultaneous collisions" with the walls of the balloon?

5. Next, go back to the **Properties** menu in the lower left corner and select **Thermodynamics → Pressure**. By adjusting the unit preference in **Tools → Preferences**, determine the pressure in each of the three units "atmospheres", "kilopascal", and "millimeters of mercury". List the three numbers in order of increasing numerical value.

6. To a precision of two significant figures, what is the conversion factor from atmospheres to kilopascal that the computer must be using?

7. To a precision of two significant figures, what is the conversion factor from kilopascal to millimeters of mercury that the computer must be using?

8. Extend the list of "Properties" with **System → Total Number of Atoms** and thereby find the number of atoms in the system.

9. Edit the value field for the number of atoms and fill in a new value of **300**. Restart the simulation. Does the number of molecule-wall collisions (as indicated by the halos) decrease or increase?

10. Does the change of the pressure parallel the change in the number of collisions?

Measuring the Specific Heat

Objectives

- To determine the specific heat of three different substances.
- To consider the uniqueness of liquid water.
- To calculate the expected temperature changes when heat is added to various systems.

Procedure

In the following lab, you will carry out molecular simulations of several substances with variation of the temperature *T* (the computer will automatically add or remove heat so as to obtain the desired temperature). Monitoring the energy *E* as a function of the temperature and working with the definition:

$$C = \Delta E / \Delta T$$

will lead to the value for the molar heat capacity *C*.

Models of three common substances are available:

⊙ **Water** H_2O (*l*)

○ **Iron** Fe (*s*)

○ **Ethanol** C_2H_5OH (*l*)

Start the simulation of the first substance (below the model, click on the **Start** button). To find the heat capacity, you need to set up a temperature-energy plot:

- From the **View** menu, select **Plots**. From the **Add Plots** menu (upper left corner), select **XY Plot...**.
- For the "X Axis," select **Temperature**. For the "Y Axis," select **Total Energy (Molar)** and set the checkmark for **Use Average**. Click **Next** and then select **Linear Fit**. Click **Finish**.
- After the simulation has run for a little while, click on **Record** above the plot.
- Next, move the shown **Slider** to a new temperature value.
- Run the simulation for *at least* $\sim 10\times10^{-12}$ s, then record the datapoint for the new temperature.
- Collect a total of at least 6 data points. *Note*: The longer you let the simulation run for each new set of conditions (especially after a large jump in temperature), the "smoother" the curve will look.
- Switch to the second substance and start its simulation. *Add a second plot frame* by selecting **Add Plots → XY Plot...**. Follow the instructions as before to obtain the datapoints for this new substance. (Be sure that the *second* plot frame is selected—when selected, it has a bluish border around it.)
- Finish by going through the same steps also for the third substance, i.e., create a *third* plot frame and collect several datapoints. (This time be sure that the third plot frame is selected—when selected, it has a bluish border around it.)

- The molar heat capacity C (short "molar heat") is the slope of the E = f (T) curve. For each substance, obtain its value by selecting the corresponding plot (if selected it is surrounded by a bluish boundary) and then **Plot Edit** (above the plot, click on the pencil icon). Set the checkmark for **Display Equation** and click **OK**. The values found are *molar* because the energy is given per mole.

Analysis

11. What do you find for the molar heat C of the three substances?

12. Convert your results to values for the *specific* heat capacity c (short "specific heat") of the three substances, i.e., the heat capacity per *gram* rather than per mole. The conversion between molar and specific heat involves the molar mass M:

$$c = C / M$$

13. It is often pointed out that the heat capacity of water is "special" and that this has many ramifications for the Earth's climate, for physiology, ecology, and more. Among the three substances examined here, is the specific heat of liquid water "special" (either very large or very small)?

14. Given the few datapoints that you have, does water seem to be "special" with regard to the specific heat, the molar heat, or both?

15. The specific heat capacity of ice (solid water) is ~2.0 J K^{-1} g^{-1}. Steam (gaseous water) in a rigid container has a specific heat capacity of ~1.4 J K^{-1} g^{-1}. Is it water in general or is it just *liquid* water that is "special"?

16. Suppose that you eat a candy bar that according to the label contains "220 Calories." Assuming your body was a perfect calorimeter filled with water, by how much would your body temperature rise? Estimate the amount of water in your body as 60% of your personal weight and use the value for the specific heat that you determined above. Assume that the energy released by the digestion of the candy bar only serves to heat up the water component of your body (i.e., neglect the heat capacity of the non-water component). Don't forget to convert the amount of energy in the candy bar from dietary (nutritional) calories to "scientific" calories (or rather J or kJ).

17. If the fluid in your body was ethanol rather than water, by how much would you heat up upon consumption of a 220 Calories candy bar? (Use your measured value for the specific heat and make the same assumptions as before.)

18. The two simulation cells of liquid water and solid iron have the same dimensions, i.e., they are molecular-level realizations of *equally-sized* samples of water and iron. If both systems are at room temperature and each absorbs 10^{-19} J of heat, what is the final temperature (in °C) of the two molecular samples going to be?

    (You may find it useful to work with the *molar* heat values and to query the **System → Total Number of Molecules** property.)

19. You may have found in your answer to the last question that there is not much difference (⇒ less than a factor of 2) between the effect of a given amount of energy on the temperature change of samples of water and iron that are of equal size. In everyday life, however, we notice that an empty iron bucket exposed to the sunlight heats up much faster than the water in a filled bucket. How can this be reconciled with your answer to Question 8? (*Hint:* Think about the amount of iron involved.)

Chemical Equilibrium and Pressure

Objectives

To examine the applicability of Le Chatelier's principle to the pressure dependence of a simple recombination reaction.

Procedure

The gas phase dissociation and recombination of iodine provides a very simple example of a chemical equilibrium:

$$I_2\ (g) \rightleftharpoons 2\ I\ (g)$$

A corresponding simulation setup is shown:

⊙ **Iodine** (Gas Phase)

Start the simulation (below the model, click on the **Start** button). The system's pressure is:

P = **…**

Since the number of molecules (or moles) is *not* the same on the two sides of the reaction equation, this is an equilibrium that is potentially pressure dependent. An easy way to get a reasonable estimate for the "equilibrium concentrations" of the two kinds of particles is to create two plots for the number of I_2 molecules and I atoms as a function of time:

- From the **View** menu, select **Plots**. From the **Add Plots** menu (upper left corner), select **XY Plot...**.
- For the "X Axis," select **Time**. For the "Y Axis," select **Number of Molecules**. Also set the checkmark for **Use Average**. Click **Finish**.
- In the Y axis label of the plot, click on **<select group>** and then click on one of the I_2 molecules.
- Bring up the **Plot Edit** dialog (above the plot, click on the pencil icon). For the "X Axis", change the "Window" to **40**. Click **OK**.
- Generate a second plot identical to the first one, except that the number of I atoms is shown on the Y axis. Remember to set the checkmark for **Use Average** and to change the "Window" size to **40**.

Analysis

1. Run five simulations in which the volume of the simulation cell varies between about 500 nm^3 and about 2,500 nm^3. To change the simulation volume:
   - From the **View** menu, select **Properties**. From the **Add Properties** menu (lower left corner), select **Thermodynamics** → **Volume**.
   - Edit the number field for the volume to change its value.

   Give the average number of I_2 molecules and I atoms as well as the average pressure for each of the 5 simulation conditions (that is, besides the volumes give a total of 15 numbers).

2. The ratio

$$N_{\text{atomic iodine}}^2 / N_{\text{molecular iodine}}$$

   represents the relative share of the two species in the reaction mixture and thus provides a measure for the *position* of the equilibrium under a given set of conditions. Give the value of this particular ratio for each of the conditions.

3. Using graph paper or a spreadsheet or plotting program, plot the ratio that you calculated in the previous question as a function of the simulation cell volume.

4. Interpret the obtained graph in terms of Le Chatelier's principle.

5. Predict the effect of a *decrease* in pressure on the following chemical equilibria:

   i. $2\ NOCl\ (g) \leftrightharpoons 2\ NO\ (g) + Cl_2\ (g)$

   ii. $2\ SO_2\ (g) + O_2\ (g) \leftrightharpoons 2\ SO_3\ (g)$

   iii. $H_2\ (g) + Cl_2\ (g) \leftrightharpoons 2\ HCl\ (g)$

   iv. $CH_4\ (g) + H_2O\ (g) \leftrightharpoons CO\ (g) + 3\ H_2\ (g)$

   v. $NO\ (g) + CO_2\ (g) \leftrightharpoons NO_2\ (g) + CO_2\ (g)$

Identifying Functional Groups

Objectives

Being able to recognize not only some, but *all* of the functional groups of an organic compound is essential for predicting its chemical and biological reactivity. Your task in this experiment is

- to identify the functional groups in a homologous series of five-carbon compounds,
- to identify the functional groups in a number of biologically active substances.

Following a practice common among chemists, you are asked to account not only for groups that contain elements other than carbon and hydrogen, but also for any double and triple bonds between carbon atoms (but do not include benzene rings).

Procedure and Analysis

Each of the following five-carbon compounds includes exactly one functional group:

⦿ **Compound #1**	○ **Compound #6**
○ **Compound #2**	○ **Compound #7**
○ **Compound #3**	○ **Compound #8**
○ **Compound #4**	○ **Compound #9**
○ **Compound #5**	

1. Match the compound numbers with the functional group names:

a. alcohol	f. carboxylic acid
b. ether	g. ester
c. amine	h. alkene
d. aldehyde	i. alkyne
e. ketone	

2. Although you may not have considered it before, the concept of "structural isomers" can be applied not only to pure hydrocarbons, but also to molecules with functional groups. From the list of compounds above, can you identify any pairs that are structural isomers? For all pairs, provide the compound numbers and the names of the two classes of compounds.

3. Give *all* of the functional groups that pertain to the following amino acids:

   - **Serine**
   - **Lysine**
   - **Aspartic Acid**
   - **Ornithine** (Non-Standard Amino Acid)

4. Examine the following drugs, nutrients, and food additives:

   - **Aspirin** (Analgesic)
   - **Calciferol** (Vitamin D2)
   - **Cholesterol** (Steroid)
   - **Cortisone** (Steroid Hormone)
   - **Ibuprofen** (Analgesic)
   - **Naproxen** (Analgesic)
   - **Simvastatin** (Cholesterol-Lowering Drug)
   - **Tocopherol** (Vitamin E)

   For each of these molecules, identify *all* of the functional groups that are present (do not include any benzene rings).

5. Which of the functional groups that have appeared in this experiment can never be associated with the terminal carbon atom of a chain of carbons?