Water Quality Field Guide

PASCO®

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Introduction

Introduction

Welcome to the Water Quality Field Guide Reference Manual

Dear Instructor,

Hands-on water quality studies are motivational for students of all ages and levels of science education. From elementary school through college, water quality studies and activities can be conducted such that students learn both basic science concepts—such as measuring, hypothesis testing, data analysis, and documenting and reporting—and concepts specific to environmental studies—such as interrelationships of factors in the biosphere, seasonal and regional variability, and impacts on the environment of human activity and technology.

The water quality sensors and ezSample™ test kits combined with PASCO data collection systems, as well as the ezSample Field Titrator kits, are powerful, sensitive, accurate, and easy-to-use tools for measuring water quality parameters.

The purpose of this field guide is to provide you with all the background, reference information, and procedural instruction needed to use these tools in a meaningful way in teaching and hands-on learning about environmental water quality.

Sincerely,

Korey Champe

Earth and Environmental Sciences Education Manager

PASCO Scientific

Using PASCO's water quality measuring technologies: an overview

The water quality parameters that can be measured using PASCO's technology are listed in Table 1. The sensors and ezSample test kits can be used with all of PASCO's data collection systems and can be interfaced directly to a computer running PASCO's DataStudio® or SPARKvue™ software.

"Data collection system" refers to the data collection, display, and analysis device used to carry out the various PASCO Water Quality Field Guide activities. These include PASCO's DataStudio, the Xplorer GLX™, SPARKvue, and SPARK Science Learning System™. In some cases, a sensor extension cable facilitates the connection between the data collection system and the sensor. Finally, the ezSample field titrator kits do not involve electronic sensing or electronic data collection.

In addition to these technologies, PASCO's GPS (global positioning system) Sensor can be incorporated into water quality studies, enabling students to participate in global information system (GIS) activities.

This field guide has been written for use with all of PASCO's data collection systems. The actual instructions for the technical procedures (referred to as "Tech Tips") are in the appendix that corresponds to the PASCO data collection system being used for the activity. The Tech Tip number at the end of a technical step is used to locate that Tech Tip in the appropriate appendix.

Integrating water quality field studies into curricula across the grade levels

The field of water quality monitoring can be challenging even for college-level students and professionals, since it encompasses a wide range of monitoring activities, technologies, and expertise requirements.

However, specific elements of water quality testing can be easily integrated into curricula in such a way that any student's current scientific development level can be advanced. One way to illustrate this is to examine the relevant National Science Standards at different grade levels correlated with examples of hands-on water quality activities that might address those standards (Tables 2 and 3).

At the advanced high school and college levels, students will design investigations using the appropriate analytical technologies. At this level, accuracy and reliability of data is emphasized, so instrument calibration and the use of external standards will be emphasized. Regional comparisons and cooperative databases may then be incorporated.

Table 1.

Listing of water quality parameters that can be measured using PASCO technology*

Water Quality Parameter	PASCO Measuring Technology
Alkalinity	Total Alkalinity ezSample Field Titrator Kit.
Ammonia	Ammonia ezSample Test Kit*
Biological Oxygen Demand (BOD)	PASPORT Dissolved Oxygen Sensor *or* PASPORT Water Quality MultiSensor
Chlorine	Chlorine ezSample Test Kit, PASPORT Water Quality Colorimeter
Conductivity (TDS estimation)	PASPORT Conductivity Sensor *or* PASPORT Water Quality MultiSensor
Dissolved CO_2 Gas	Dissolved CO_2 Gas ezSample Field Titrator Kit
Dissolved Oxygen (DO)	Dissolved Oxygen Sensor *or* PASPORT Water Quality MultiSensor
Hardness	Total Hardness ezSample Field Titrator Kit
Iron	Iron ezSample Test Kit, PASPORT Water Quality Colorimeter
Nitrate	Nitrate ezSample Test Kit, PASPORT Water Quality Colorimeter
pH	PASPORT pH Sensor *or* PASPORT Water Quality MultiSensor
Phosphate	Phosphate ezSample Test Kit, PASPORT Water Quality Colorimeter
Salinity	PASPORT Salinity Sensor
Stream Flow	Flow Rate/Temperature Sensor
Temperature	PASPORT Temperature Sensor *or* Flow Rate/Temperature Sensor *or* PASPORT Water Quality MultiSensor
Thermocline	PASPORT Thermocline Sensor
Turbidity	PASPORT Turbidity Sensor

* See ordering information.

Table 2.

Correlation of National Science Content Standards and suggested water quality testing activities: grade levels 5 through 8

Content Standard	Example Activities
A: Science as Inquiry Develop abilities to do scientific inquiry—Use appropriate tools for measuring and recording and analyzing data; Develop understandings about scientific inquiry.	Develop hypotheses about variations in temperature in a body of water over time, at different depths, at different locations, etc. Using a temperature sensor, record temperature to test the hypotheses. Display graphs and tables of the data and explain temperature variations, verifying or nullifying the hypotheses.
B: Physical Science Develop understanding of properties and changes of properties in matter.	Using sensors, explore temperature, pH, turbidity (suspended solids), and conductivity (total dissolved solids) in two different bodies of water, in a body of water at different times of the year, or before and after a rain storm.
C: Life Science Develop an understanding of • structure and function of living systems • reproduction and heredity • regulation and behavior; • populations and ecosystems • diversity and adaptation of organisms	Conduct a survey of two distinctly different watershed areas and the waterways. Have students record careful observations of the terrain, terrestrial and aquatic vegetation, insects, birds, and other animals that can be observed. Using sensors, have students record the ambient air temperature and the temperature, pH, conductivity, dissolved oxygen, and turbidity of the water bodies. Identify some correlations (e.g., dissolved oxygen levels are higher in clear cold water, which supports the reproduction and living process of a different variety of organisms compared to warm, turbid water; birds that feed on fish, like herons, seem to be found more frequently around richly various ecosystems; pH may be lower in high mountain streams, which have lower conductivity levels; watersheds that have thick vegetation surrounding the water body are associated with less turbid water, etc.).
D: Earth and Space Science Develop understanding of the structure of the earth system, including • water cycle, • water as a solvent, • the sun's role in water quality and the water cycle	In a watershed study, trace the path of water from its origin to the ocean, and then back to the watershed. Using the conductivity and pH sensors, map the conductivity and pH of the water body. Develop hypotheses regarding the origins of the dissolved solids, acids, and bases. Track water temperature over time and at different locations and consider the role of the sun in variations in temperature that are recorded.
E: Science and Technology Develop abilities to design and assess technology; Develop understandings about science and technology.	Design a water treatment system. Using filtration (sand, gravel), sedimentation, coagulation (egg white), and activated charcoal, design a system that removes suspended solids, sediment, odor, and color from a "polluted" water source (with coffee grinds and pulverized kitty litter). Use the turbidity, pH, and conductivity sensors to analyze the water before and after treatment to determine which properties were most affected by the purification technology and which design was most effective. Take a field trip to the local waste water treatment facility.
F: Science in Personal and Social Perspectives Develop understanding of • personal health • populations, resources, and environments • natural hazards • risks and benefits • science and technology in society	Using the conductivity, pH, and dissolved oxygen sensors, explore a variety of drinking water sources (tap water, bottled water, filtered water, distilled water, well water, etc.). Compare findings with EPA drinking water standards. Conduct taste tests of these water sources. Discuss the variability found in the context of health and safety (i.e., despite considerable variability, especially regarding dissolved solids and taste, all sources are safe to drink and promote personal health). Discuss the risks-benefits of bottled water versus tap water (e.g., the trade-off of spending limited financial resources for bottled water instead of fruits and vegetables). Take a field trip to the local drinking water treatment facility.

Table 2.

Correlation of National Science Content Standards and suggested water quality testing activities: grade levels 5 through 8

Content Standard	Example Activities
G: History and Nature of Science Develop understanding of • science as a human endeavor • nature of science • history of science	Throughout the hands-on exploration of water quality, provide historical context regarding scientific and technological progress and processes regarding water quality, including videos, readings, internet research on topics such as: • the impact of advances in scientific understanding and sanitation technologies on improved human health (e.g., the role of Louis Pasteur in developing the germ theory of disease, the role of epidemiologists, John Snow and William Farr in understanding the cause of the 1849 cholera epidemic in London); • the role of environmentalists in identifying, documenting, and creating social changes regarding acid rain and its harmful effects; • current concerns regarding global warming and its effects on water temperatures and levels and consequences for plants, animals, and humans, set in the context of historical climate changes (e.g., ice ages, dinosaur age, etc.); • the role of satellites and global information systems on monitoring water quality.

Table 3.

Correlation of National Science Content Standards and suggested water quality testing activities: grade levels 9 through 12

Content Standard	Example Activities
A: Science as Inquiry Develop abilities to do scientific inquiry—Use appropriate tools for measuring and recording and analyzing data; Develop understandings about scientific inquiry.	Develop hypotheses about variations in water quality parameters in a body of water over time, at different depths, at different locations, etc. Design investigations using appropriate sensors and test kits to measure parameters and collect data to test the hypotheses. Display graphs and tables of data and mathematically analyze variations, verifying or nullifying the hypotheses. Write reports that clearly and succinctly include the following: statement of the problems and related historical information, the hypotheses, the experimental design, the methodologies used to test the hypotheses and analyze the data, present the data, discuss the data and the significance of the findings as well as an alternative hypotheses that might explain the data, and list the references used.
B: Physical Science Develop understanding of • structure of atoms • structure and properties of matter • chemical reactions • motions and forces • conservation of energy and increase in disorder • interactions of energy and matter	Using sensors and colorimetric kits, explore water quality parameters in two different settings or in a body of water at different times of the year, before and after a rain storm, or at different locations relative to a point source of pollution. Record all data, including air temperature and relevant weather and ecosystem data (e.g. surrounding terrain, nature of run-off, surrounding vegetation, terrestrial and aquatic plants and animal species observed, etc.) Analyze data, looking for notable differences in parameter levels between the two situations. Develop hypotheses that might explain these differences in terms of thermal energy, stream flow, ionic content of the water or run-off, action of erosion, chemical reactions of rain or groundwater comprising run-off with minerals in the watershed, diluting influence of water, buffering capacity of the water, or other chemical or physical properties of the water and watershed.
C: Life Science Develop understanding of • the cell • molecular basis of heredity • biological evolution • interdependence of organisms • matter, energy, and organization in living systems • behavior of organisms	Using sensors and colorimetric kits, explore water quality parameters in two different settings or in a body of water at different times of the year, before and after a rain storm, or at different locations relative to a point source of pollution. Record all data, including air temperature and relevant weather and ecosystem data (e.g. surrounding terrain, nature of run-off, surrounding vegetation, terrestrial and aquatic plants and animal species observed, etc.) Collect water samples and examine using a microscope. Conduct microbiological analysis for coliform bacterial. Analyze data, looking for notable differences in water quality parameter levels between the two situations. Develop hypotheses regarding the effect these differences in parameter levels might have on the content of the plant and animal species in the ecosystem, as well as on humans.
D: Earth and Space Science Develop understanding of the structure of the earth system, including • energy in the earth system • geochemical cycles • origin and evolution of the earth system • origin and evolution of the universe	In a watershed study, trace the path of water from its origin to the ocean, and then back to the watershed. Using sensors and colorimetric kits, explore water quality parameters. Develop hypotheses regarding the origins of the substances detected in the water. Track water temperature over time and at different locations, and consider the role of the sun in the variations in temperature that are recorded. Develop hypotheses regarding the impact of local climate changes, such as those produced by global warming, on the quality of the water and the composition of the ecosystem in this watershed.
E: Science and Technology Develop abilities to design and assess technology; Develop understandings about science and technology.	Design a water treatment system. Using filtration (sand, gravel), sedimentation, coagulation (egg white), and activated charcoal, and other devices, design a system that removes pollutants from a polluted water source. Using sensors and colorimetric kits, analyze the water before and after treatment to determine which properties were most affected by the purification technology and which design was most effective. Write a report that supports the conclusion with the data collected. Include in the report the results of Internet research on new developments in water treatment technologies.

Table 3.

Correlation of National Science Content Standards and suggested water quality testing activities: grade levels 9 through 12

Content Standard	Example Activities
F: Science in Personal and Social Perspectives Develop understanding of • personal and community health • population growth • natural resources • environmental quality • natural and human-induced hazards **science and technology in local, national, and global challenges**	Using sensors and colorimetric kits, explore water quality parameters in a variety of drinking water sources (e.g., tap water, bottled water, filtered water, distilled water, well water, etc.). Compare findings with EPA drinking water standards. Conduct taste tests of these water sources. Discuss the variability found in the context of health and safety (i.e., despite considerable variability, especially regarding dissolved solids and taste, all sources are safe to drink and promote personal health). Discuss the risks-benefits of bottled water versus tap water (e.g., the trade-off of spending limited financial resources on bottled water instead of fruits and vegetables). Develop hypotheses regarding potential situations in which the risk-benefit ratio might be different.
G: History and Nature of Science Develop understanding of • science as a human endeavor • nature of science • history of science	Throughout the hands-on exploration of water quality, have students write reports on such topics as: • the impact of advances in scientific understanding and sanitation technologies on improved human health (e.g., the role of Louis Pasteur in developing the germ theory of disease, the role of epidemiologists, John Snow and William Farr, in understanding the cause of the 1849 cholera epidemic in London); • the role of environmentalists in identifying, documenting, and creating social changes regarding acid rain and its harmful effects; • concerns regarding global warming and its effects on water temperatures and levels as well as on humans, set in the context of historical climate changes (e.g., ice ages, dinosaur age, etc.); • the role of satellites and global information systems on monitoring water quality.

About integration of water quality studies into state science standards

The suggested ways of integrating water quality studies into the National Science Content Standards shown in the above tables are given as examples for your use. However, a similar matrix for integrating water quality studies into individual state science standards is beyond the scope of this manual. Consider matching the activities you choose to your state standards that are equivalent to the National Science Content Standards.

Basic Water Quality Concepts

Designated Use

Designated Use and PASCO's Measuring Technology

A water quality standard depends on the use or uses to be made of the water. Therefore, standards vary, depending on the designation of use. Some examples of designated use include[3]:

- Public water supplies (for tap water)
- Protection and propagation of fish, shellfish, and wildlife
- Recreation
- Agriculture/industry
- Navigation
- Coral reef preservation
- Marinas
- Groundwater recharge
- Aquifer protection
- Hydroelectric power

This manual is targeted primarily towards exploring water quality as it relates to public water supplies; protection and propagation of fish, shellfish, and wildlife; and recreation. These issues are of most immediate concern to students and teachers, and these relate most closely to science curriculum standards.

Quality standards for drinking water

The most stringent water quality standards are those for drinking water. The U.S. Environmental Protection Agency (EPA) has published regulations for maximum contamination levels, as well as public health goals, for 87 substances,[1] comprising the following:

- Disinfectants
- Disinfection byproducts
- Inorganic chemicals
- Microorganisms
- Organic chemicals
- Radionuclides

Among these, PASCO provides measuring technology for two of these: chlorine and turbidity.

In addition, the EPA has designated 15 substances as nuisance substances that may cause cosmetic effects (such as skin or tooth discoloration) or aesthetic effects (such as taste, odor, or color) in drinking water. For these substances, the EPA recommends secondary standards[2]; however, these recommendations are not enforced. PASCO provides measuring technology for two of these: pH and total dissolved solids.

Quality of natural water bodies quality considerations for natural bodies of water and their ecosystems

When evaluating water quality for natural water bodies, a complex interplay of physical, chemical, and biological factors must be considered.[3] PASCO offers measuring technology for many of the typical chemical factors, as well as for some of the physical factors (see the table below).

Summary of Typical Factors Used in Conducting a Water Body Survey and Assessment*

Physical Factors	Chemical Factors	Biological Factors
instream characteristics	**dissolved oxygen (DO)****	**biological inventory (existing use analysis)**
size (mean width/depth)	**toxicants**	fish
flow/velocity**	**suspended solids (turbidity)****	macroinvertebrates
annual hydrology	**nutrients**	microinvertebrates
total volume	nitrogen (nitrate, nitrite)**	phytoplankton
reaeration rates**	phosphorous (phosphate)**	macrophytes
gradient/pools/riffles	**sediment oxygen demand***	**coliform bacteria**
temperature**	**salinity****	**biological potential analysis**
sedimentation	**hardness****	diversity indices
channel modification	**alkalinity****	habitat suitability index (HSI) models
channel stability	**pH****	tissue analyses
substrate composition and characteristics	**dissolved solids****	recovery index
channel debris		intolerant species analysis
sludge deposits		omnivore-carnivore analysis
riparian characteristics		reference reach comparison
downstream characteristics		

**Water Quality Handbook, 2nd Edition; Chapter 2: Designation of Uses*. 1994, Environmental Protection Agency.

**PASCO offers measuring technology for these parameters.

References

1. *National Drinking Water Standards*. 2002, U.S. Environmental Protection Agency.
2. *National Secondary Drinking Water Regulations. Current Drinking Water Standards* 2006, U.S. Environmental Protection Agency.
3. *Water Quality Handbook, 2nd Edition; Chapter 2—Designation of Uses*. 1994, U.S. Environmental Protection Agency.

Water Quality Index

A water quality index (WQI) incorporates several key indicators of water quality into one score for purposes of monitoring and comparing water quality over time. The first WQI was proposed by researchers associated with the National Sanitation Foundation in 1970.[1,2]

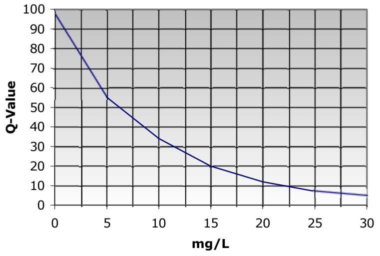
> **Note:** This WQI is essentially a eutrophication index.

For this index, the following 9 parameters are measured:

- Biological oxygen demand (also known as the 5-day biochemical oxygen demand or BOD)
- Dissolved oxygen
- Fecal coliform bacteria
- Nitrate
- pH
- Temperature
- Total dissolved solids (TDS)
- Total phosphates
- Turbidity

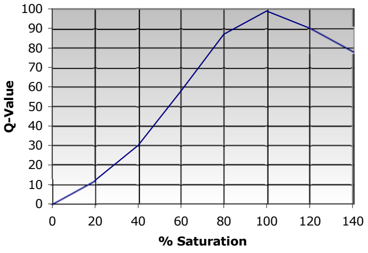
A Q-value is a value statistically derived that allows varying water quality tests to be compared. Each measurement is associated with its Q-value as shown on the graphs.[2]

Chart 1
BOD Test Results



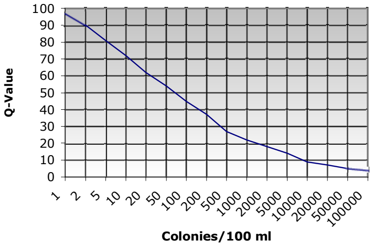
(Note: If BOD > 30, Q=2.0)

Chart 2
Dissolved Oxygen Results



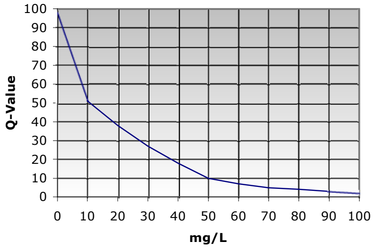
(Note: if % saturation is > 140.0, Q=50.0)

Chart 3
Fecal Coliform Results



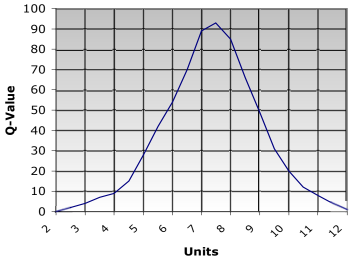
(Note: if Fecal Coliform > 100000, Q=2.0)

Chart 4
Nitrate Results



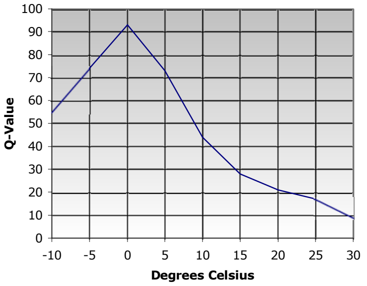
(If Nitrates > 100.0, Q=1.0)

Chart 5
pH Results



(Note: If pH is < 2.0 or > 12.0, Q=0)

Chart 6
Temperature Results



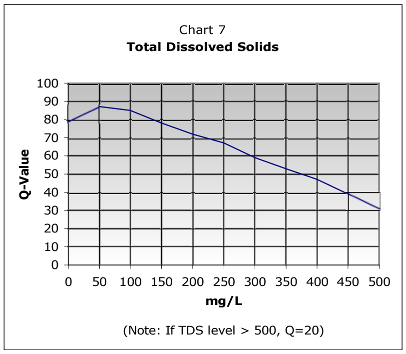


Chart 7
Total Dissolved Solids

(Note: If TDS level > 500, Q=20)

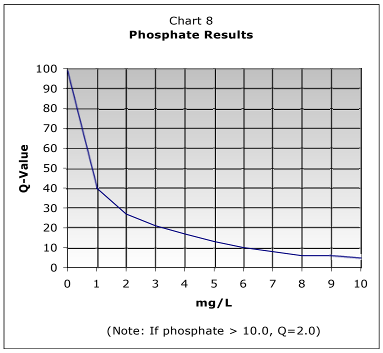


Chart 8
Phosphate Results

(Note: If phosphate > 10.0, Q=2.0)

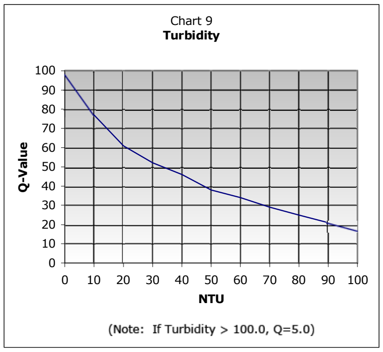


Chart 9
Turbidity

(Note: If Turbidity > 100.0, Q=5.0)

After data is collected, the tables are used to find the associated Q-value. The Q-value is then entered into a table such as the one on the worksheet below. A weighting factor is assigned to each parameter, depending on its effect on water quality, which then converts the Q-value into a number that can be added to other values to determine an overall water quality number.

Water Quality Index Worksheet

Tester's Name ____

Location of Sample ____

Date and Time of Test ____

Test Parameter	Test Results	Q-Value	Weighting Factor	Total
BOD	(mg/L)		0.11	
Dissolved Oxygen	(% saturation)		0.17	
Fecal Coliform	(colonies/100 mL)		0.16	
Nitrates	(mg/L)		0.10	
pH			0.11	
Temperature	°C		0.10	
Total Dissolved Solids	(mg/L)		0.07	
Total Phosphate	(mg/L)		0.10	
Turbidity	(NTU)		0.08	
			Overall Water Quality Index: :	

Note: This worksheet is included for general information purposes. You can download a full-size version from the National Science Foundation Web site. Refer to the Recommended Reading and Resources section for availability information.

Finally, the overall water quality index value as determined from the worksheet is then assigned a qualitative description as shown on the following table:

Qualitative water quality descriptions

Water Quality Index (WQI) Scale	
91–100	Excellent water quality
71–90	Good water quality
51–71	Medium or average water quality
26–50	Fair water quality
0–25	Poor water quality

Additional types of WQIs

Additional WQIs that are modifications of the original version have been adopted worldwide for monitoring water quality regionally. The various modifications have been added to respond to variations in regional concerns. For example, the United States state of Iowa has adopted a modified WQI for monitoring its natural water bodies that includes the following 9 parameters: biological oxygen demand, dissolved oxygen, *E. coli* bacteria, nitrate and nitrite as nitrogen, total detected pesticides, pH, total phosphorus, total dissolved solids, and total suspended solids.

These modifications of the original WQI procedure have resulted in a more sensitive and specific assessment of Iowa waters.[4] In an additional example, Canada uses a WQI that includes 10 parameters as follows: dissolved oxygen, pH, total phosphorus, total nitrogen, fecal coliform bacteria, arsenic, lead, mercury, 2,4-D, and lindane.[5] In a final example, the United States state of Oregon uses a WQI that includes the following 8 parameters: ammonia plus nitrate nitrogen, biochemical oxygen demand, dissolved oxygen, *E. coli* bacteria, pH, temperature, total phosphorus, and total solids.[6]

Although a detailed discussion of the procedure for calculating a WQI is beyond the scope of this field guide, some of the cited references provide detailed procedures, such as the National Science Foundation web site (refer to the Recommended Reading and Resources section).

PASCO's measuring tools can be used in the particular WQI determination your group has chosen to use. Remember that careful, standardized calibration procedures are critical to the usefulness and reliability of a WQI.

Further information is available from the references which can be found in the Recommended Reading and Resources section.

References

1. Brown R, McClelland N, Deininger R, Tozer R. *A Water Quality Index—Do We Dare?* Water and Sewage Works. 1970:339–43.
2. National Science Foundation. *Water Quality Index*. 2004.
3. Delzer GC, McKenzie SW. *Five-Day Biochemical Oxygen Demand*. National Field Manual. U. S. Geological Survey; 1999.
4. *Why a Water Quality Index? Water Fact Sheet 2006–8*. Iowa City, IA Iowa Department of Natural Resources, Geological Survey; 2006.
5. Saffran K, Cash K, Hallard K, Neary B, Wright R. *Water Quality Index 1.0 User's Manual*. Canadian Council of Ministers of the Environment (CCME)
6. *Oregon Water Quality Index Methodology*. Oregon Department of Environmental Quality; 2006.

Parameter and Measurement Information

Alkalinity

What is alkalinity?

Alkalinity is the water's capacity to resist changes in pH that would make the water more acidic. This capacity is commonly known as its *buffering capacity*, the capability of water to neutralize acid.

A buffer is a solution to which an acid can be added without changing the concentration of available H^+ ions (without changing the pH) appreciably. The components of the buffer absorb the excess H^+ ions and protect the water body from fluctuations in pH (refer to the section on pH for additional background). For example, if you add the same weak acid solution to two vials of water—both with a pH of 7, but one with no buffering capacity and the other with buffering capacity—the pH of the water with no buffering capacity will immediately drop while the pH of the buffered water will change very little or not at all. The pH of the buffered solution would change when the buffering capacity of the solution is overloaded.[1]

Technically, the term *alkalinity* refers to filtered water and the term *acid-neutralizing capacity* (ANC) refers to unfiltered water, but both are measured using the same procedure involving acidimetric titration with a strong acid to a pH of 4.5.[4] Alkalinity is measured in units of ppm or mg/L as $CaCO_3$.

Alkalinity is often related to hardness because one source of alkalinity is carbonate rocks (limestone), which are mostly $CaCO_3$. If $CaCO_3$ actually accounts for most of the alkalinity, hardness in terms of $CaCO_3$ is equal to alkalinity. Conversely, soft water usually has low alkalinity and little buffering capacity unless the carbonate ions are associated with sodium or potassium ions, which do not contribute to hardness. So generally, soft water is much more susceptible to fluctuations in pH from acid rains or acid contamination.[1] (For more information, refer to the section on Hardness.)

Why measure alkalinity?

Alkalinity is important for fish and other freshwater aquatic life because it protects or buffers against rapid pH changes that occur naturally as a result of photosynthetic activity or against acid shock that may occur in spring when acidic snow melts or following acidic rain or other acidic discharges that may enter the stream.

If increasing amounts of acids are added to a body of water, the water's buffering capacity is consumed. If additional buffering material can be obtained from surrounding soils and rocks, the alkalinity level may eventually be restored. However, a temporary loss of buffering capacity can permit pH levels to drop to those harmful to life in the water.

The ions that contribute to alkalinity—primarily carbonate (CO_3^{2-}), bicarbonate (HCO_3^-), phosphate (PO_4^{2-}), and hydroxide (OH^-)—also combine with toxic heavy metals and markedly reduce their toxicity.[3]

Excessive alkalinity can cause problems for swimmers by altering the pH of the lacrimal fluid around the eye, causing irritation.

In water used for irrigation, high alkalinity may indirectly increase the relative proportion of sodium in soil water. This may occur when bicarbonate concentrations are high; as water evaporates, the calcium and magnesium ions that are in solution precipitate as carbonates in the soil, causing a relative increase in sodium concentration, which can result in soil and plant damage. High levels of alkalinity may also lead to chlorosis (yellowing of leaves) in plants because it causes the iron to precipitate as a hydroxide, making the iron unavailable to plants.

For industrial water supplies, high alkalinity can be damaging, particularly for food production in which acidity accounts for flavor and stability, such as for carbonated beverages. In other instances, alkalinity is desirable because water with higher alkalinity is much less corrosive; however, water that is very high in alkalinity can be corrosive.[3]

What factors affect alkalinity?

Both alkalinity and the acid-neutralizing capacity (ANC) of natural water are determined by the soil and bedrock through which it passes. The main sources for natural alkalinity are rocks that contain carbonate, bicarbonate, and hydroxide compounds. Borates, silicates, and phosphates also may contribute to alkalinity.[3] The presence of calcium carbonate or other compounds such as magnesium carbonate contribute carbonate ions (CO_3^{2-}), bicarbonate ions (HCO_3^-), or carbonic acid (H_2CO_3), depending on the water's pH, to the buffering system. At a pH > 10.33, carbonate is predominant; at pH 6.4 to 10.33, bicarbonate is predominant; at pH < 6.4, carbonic acid is predominant. The negative carbonate and bicarbonate ions combine with the positive hydrogen ions (H^+), removing them from the solution and thereby increasing the pH.[2]

Limestone is rich in carbonates, so water flowing through limestone regions or bedrock containing carbonates generally has high alkalinity and therefore a high buffering capacity. Conversely, water flowing though areas rich in granites and some conglomerates and sandstones may have low alkalinity and therefore poor buffering capacity.

Human activity, such as discharging waste water containing cleaning agents made from carbonate and bicarbonate as well as residues from some food substances, can increase alkalinity. On the other hand, other human activities, such as some mining operations and operating combustion engines, add large amounts of acid to natural waters, which "use up" the chemicals that produce alkalinity, resulting in lower alkalinity. In situations where this results in very low alkalinity levels, carbonate in the form of limestone ($CaCO_3$), is sometimes added to the water to increase alkalinity to protective levels.[2]

What are ideal alkalinity values?

Insufficient alkalinity results in harmful decreases in pH when acid is added to water, while too much alkalinity can be harmful to humans, aquatic organisms, plants, and industrial processes because of its corrosive effects. High alkalinity (up to 400 mg/L $CaCO_3$) is not considered detrimental to humans but is associated with high pH values, hardness, and high levels of dissolved solids. High alkalinity waters may also have a flat, unpleasant taste.

The table below lists some examples of minimum and maximum levels of alkalinity by designated use. A general rule of thumb for natural waters is that total alkalinity should not allow the pH to exceed 9 or be less than 5.[2]

Minimum and maximum alkalinity levels for various water usages (mg/L as $CaCO_3$)[2,3]

Designated Use	Minimum	Maximum
Tap water		400
Fresh water organisms	20	
Marine organisms	25	
Irrigation		600
Industrial: Textile mill products		50-200
Paper products		75-150
Chemical products		500
Petroleum refining		500
Metals industry		200
Food canning		300
Bottled and canned soft drinks		85

Measuring Procedure

The measuring procedure for total hardness involves the Alkalinity ezSample Field Titrator kit as detailed in the instruction card that is included with the kit. The ezSample Field Titrator is shown below.

Note: If the water is turbid, it should be filtered before testing it. However, for most applications, the sediment can be allowed to settle for several hours instead being removed by filtration.

Note: The test range of the Alkalinity ezSample Field Titrator is 10–100 ppm (mg/ L) $CaCO_3$. Because the ampoules in the Alkalinity ezSample Field Titrator kit have nonlinear scales, the accuracy of these kits varies with the analyte concentration. At the low end of the test range, the accuracy is ± 5%. At the high end of the range, the accuracy falls to ± 20%.

Note: If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

References

Further information can be found in the referenced material listed below (refer to the Recommended Reading and Resources section for more availability information).

1. *Alkalinity and Stream Water Quality*. 2006, Wilkes University Center for Environmental Quality, Environmental Engineering, and Earth Sciences.
2. *Healthy Water, Healthy People Testing Kit Manual*. 2002, Bozeman, Montana: The Watercourse, International Project WET.
3. *Quality Criteria for Water (Red Book)*. 1976, Washington, D.C.: U.S. Environmental Protection Agency.
4. *National field manual for the collection of water-quality data: U.S. Geological Survey Techniques of Water-Resources Investigations*. Vol. 9. variously dated. Chaps. A1–A9, U.S. Geological Survey.

Ammonia

What is ammonia?

Ammonia is a pungent, colorless, gaseous, alkaline compound of nitrogen and hydrogen that is highly soluble in water. When ammonia dissolves in water, some of the ammonia reacts with the water to form ammonium ions. A chemical equilibrium is established that contains un-ionized ammonia (NH_3), ionized ammonia (NH_4^+), and hydroxide ions (OH^-), as shown in the following equation:

$$NH_3 + H_2O \rightleftharpoons NH_3 \cdot H_2O \rightleftharpoons NH_4^+ + OH^-$$

In this equation, NH_3 represents ammonia gas, $NH_3 \cdot H_2O$ represents the un-ionized ammonia molecule that is loosely attached to water molecules through hydrogen bonding, and the ionized form of ammonia is represented as NH_4^+. Because ionization of NH_3 has a pK value of about 9.3, NH_4^+ is the dominant chemical form in natural waters. For example, at a pH of 8.5 and temperature of 25 °C, only about 15% is present as NH_3.

Importantly, the toxicity of aqueous solutions of ammonia is attributed to the NH_3 form; NH_4^+ is thought not to be toxic or is perhaps only very slightly toxic to aqueous organisms.[3]

The term total ammonia nitrogen refers to the sum of NH_3 and NH_4^+. Aqueous ammonia is commonly measured in terms of total nitrogen due to ammonia (mg/L NH^{3-} N or mg N/L).

Why measure ammonia?

Un-ionized ammonia (NH_3) is toxic to fish and invertebrates. Concentrations as low as 0.02 mg/L of NH_3 are lethal to some fish, such as trout, and to some invertebrates, such as clams and mussels. At even lower concentrations, fish and invertebrates can sustain damage to gills, skin, various organs, and reproductive capability. Other fish, such as carp, are much less sensitive.[3] When setting up an aquarium or fish farm tank, ammonia is a major concern, since tap water lacks the presence of nitrifying bacteria that can break down the ammonia that is naturally excreted by the fish. Measuring ammonia enables the identification of point sources and nonpoint sources of ammonia-containing pollution.

What factors affect ammonia levels?

Ammonia is present in most waters as a natural biological degradation product of nitrogenous organic matter. Organisms such as fish excrete ammonia as a waste product. Ammonia may also reach ground and surface waters through discharge of fertilizers, raw sewage, or industrial wastes containing ammonia as a byproduct or wastes from industrial processes using "ammonia water."

In addition to the contribution of total ammonia levels to the toxic NH_3 levels, the concentration of NH_3 is dependent upon pH, with increasing concentrations at increased pH levels because of the equilibrium relationship among NH_3, NH_4^+, and OH^-. Temperature also affects the concentration of NH_3, which increases with increasing temperature. Ionic strength is also a factor that influences the concentration of NH_3, with slightly decreased concentration associated with increased salinity for dilute saline concentrations.[3]

Nitrifying bacteria, (e.g., *Nitrosomonas sp*., and *Nitrobacter sp*.) convert ammonia to less toxic nitrite (NO^{2-}) and then to nitrate (NO_3^-), which is considered nontoxic at levels less than 90 mg/L for warm-water fish, a level that rarely occurs in nature.[2] Thus, these bacteria are important factors in reducing ammonia levels, reducing its toxicity in natural waters as well as in aquaria and cultivated fish tanks.

What are ideal ammonia values?

It is generally agreed that in natural aqueous environments, the less ammonia the better.

The EPA has established regulations regarding maximum ammonia concentrations in terms of total ammonia nitrogen per liter (mg N/L) for fish in natural water bodies. Values for both acute and chronic exposure are stated. These regulations are divided according to species—with and without salmonids—and according to whether early life stages (eggs, embryos, and forms up to about 34 days after hatching) are present. The values are stratified according to pH and temperature.[1]

Freshwater plants are appreciably more tolerant to NH_3 than are invertebrates or fish, so ammonia criteria that are appropriate for the protection of aquatic animals will also be protective of plant life. However, many saltwater fish and invertebrates are more sensitive to NH_3 than are those living in fresh water.[2] In saltwater aquaria, removal and replacement of tank water on a regular basis is the best way to control ammonia levels.

The EPA does not have regulations or recommendations regarding ammonia in drinking water.

Further information can be found in the referenced material listed below (refer to the Recommended Reading and Resources section for availability information).

U.S. EPA chronic criteria for ammonia ($NH_3 \cdot H_2O + NH_4$, maximum allowed) for fish when early life stages are present (mg N/L)[1]

pH	Temperature °C					
	0	14	18	22	26	30
6.5	6.67	6.67	5.33	4.12	3.18	2.46
7.0	5.91	5.91	4.72	3.78	2.92	2.25
7.5	4.36	4.36	3.49	2.69	2.26	1.61
8.0	2.43	2.43	1.94	1.50	1.16	0.773
8.5	1.09	1.09	0.870	0.672	0.520	0.401
9.0	0.486	0.486	0.389	0.300	0.232	0.179

U.S. EPA chronic criteria for ammonia ($NH_3 \cdot H_2O + NH_4$, maximum allowed) for fish when early life stages are absent (mg N/L)[1]*

pH	Temperature °C		
	0–7	12	16
6.5	10.8	7.84	6.06
7.0	9.6	6.95	5.37
7.5	7.09	5.13	3.97
8.0	3.95	2.86	2.21
8.5	1.77	1.28	0.990
9.0	0.790	5.72	0.442

*** Chronic criteria for temperatures greater than 16°C are the same as for when early life stages are present.[1]**

U.S. EPA acute criteria for ammonia ($NH_3 \cdot H_2O + NH_4$, maximum allowed) for fish (mg N/L)[1]

pH	Salmonids Present	Salmonids Absent
6.5	32.6	48.8
7.0	24.1	36.1
7.5	13.3	19.9
8.0	5.62	8.4
8.5	2.14	3.2
9.0	0.885	1.32

Measuring Procedure

The measuring procedure for ammonia uses the PASPORT Water Quality Colorimeter and the Ammonia ezSample Test Kit.

Setting up the Water Quality Colorimeter

1. **Plug the Water Quality Colorimeter into the data collection system.** �(2.1)

   Task Result: By default, two ammonia readings will become available, one for low-range ammonia readings (Ammonia [L]) and one for high-range ammonia readings (Ammonia [H]).

2. **Calibrate the Water Quality Colorimeter.** �(3.1)

Taking an Ammonia Measurement

SAFETY PRECAUTIONS

❖ Wear safety glasses and protective gloves.

❖ Review the MSDS for each of the Ammonia ezSample reagents, and have them available for reference.

1. **Follow the included Test Procedure instructions for the Ammonia ezSample Test Kit for sample preparation (Steps 1 through 6).**

   Note: If ammonia is present in the sample, a blue or cyan color will develop after the 15-minute incubation period. If the color after 15 minutes is yellow or gold, there is no appreciable ammonia present in the test sample, and results should be reported as 0 mg/L.

2. **If required for your data collection system, begin by building an experiment. Otherwise, set the display to show both Ammonia measurements.** �(2.2)

3. **Set your data collection system to Manual Sampling.** �(2.3)

4. **Place the prepared test ampoule in the Water Quality Colorimeter and cover it with the black cap.**

   Note: Handle the ampoule by the tip, and wipe the outside glass lens clean with a non-abrasive cleaning tissue.

   Task Result: The ammonia reading will automatically appear in both displays (Ammonia [H] and Ammonia [L]). There may be slight fluctuations, so wait until these settle around a point.

5. **Record the ammonia reading.** �(2.4) **If the value is greater than 1.0, the high-range reading (Ammonia [H]) should be used; otherwise, use the low-range reading (Ammonia [L]).**

Note: The units *mg/L* and *ppm* (parts per million) are equivalent.

Disposal of the test vial

1. **The ampoules may contain very low levels of potentially toxic substances (see the MSDS), so check with your local regulatory agency regarding disposal procedures.**
2. **Carefully pour the test water down the drain. Shake the glass tip remaining in the sample cup into the solid waste container.**
3. **Thoroughly rinse the sample cup with deionized or distilled water.**

Accuracy, resolution, and practical detection limit (PDL)

The accuracy of the Ammonia ezSample Test Kit is ±10% at 75% of full scale range. Accuracy may be compromised if test results are outside the stated test ranges. The lower limit of the stated test range is the "practical detection limit (PDL)," defined as the lowest concentration at which less than ±30% error is routinely obtained.

For the Ammonia ezSample Test Kit, the range, accuracy, PDL, and resolution are shown on the table below.

> **Note**: If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

Ammonia ezSample Test Kit: accuracy, resolution, and PDL (mg/L)

Range	Accuracy	PDL	Resolution
0–1 (low)	0.05	0.05	0.01
1–3 (high)	0.1	1	0.1

References

1. U.S. Environmental Protection Agency, *1999 Update of Ambient Water Quality Criteria for Ammonia.* 1999.
2. U.S. Environmental Protection Agency, *Quality Criteria for Water 1986 (Gold Book).* 1986. U.S. Environmental Protection Agency.
3. U.S. Environmental Protection Agency, *Quality Criteria for Water (Red Book).* 1976, Washington, D.C.: U.S. Environmental Protection Agency.

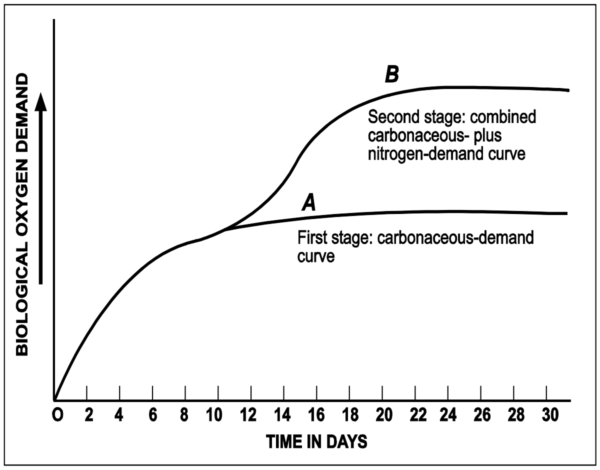
Biological Oxygen Demand

What is biological oxygen demand (BOD)?

Biological oxygen demand (BOD), also known as *biochemical oxygen demand*, is a bioassay procedure that measures the dissolved oxygen (DO) consumed by bacteria from the decomposition of organic matter. The BOD analysis is an attempt to simulate by a laboratory test the effect that organic material in a water body will have on the DO in that water body.

Biochemical oxygen demand values are a measure of food for naturally occurring microorganisms or, in other words, a measure of the concentration of biodegradable organic material. When nutrients are introduced, naturally occurring microorganisms begin to multiply at an exponential rate, resulting in the reduction of DO in the water. The test does not determine the total amount of oxygen demand present, since many compounds are not oxidized by microorganisms under conditions of the test.

There are two stages of decomposition involved in biological oxygen demand (BOD): a carbonaceous stage and a nitrogenous stage (refer to the figure below). The typical carbonaceous-demand curve (A) shows the oxidation of organic matter. The typical carbonaceous-plus-nitrogenous-demand curve (B) shows the additional oxidation of ammonia and nitrite.[1]



The carbonaceous stage, or first stage, represents that portion of oxygen demand involved in the bacterial conversion of organic carbon to carbon dioxide. The nitrogenous stage, or second stage, represents a combined carbonaceous plus nitrogenous demand, when organic nitrogen, ammonia, and nitrite are converted to nitrate by bacteria, a process that also consumes DO.

The most common method for measuring BOD is the 5-day BOD method (BOD_5), which is a method approved by the U. S. Environmental Protection Agency (EPA) for use during regulatory monitoring. In the BOD_5 test, which uses naturally occurring microorganisms to oxidize the carbonaceous organic matter, the change in DO concentration is measured before and after a 5-day period in water samples that are incubated at a specified temperature (20 °C ± 1 °C) in darkness. The BOD test results are reported as mg/L DO.[1]

The BOD_5 method measures most of the carbonaceous stage of the BOD (typically about 60% to 70%; refer to the figure showing Biological Oxygen Demand). This method has been widely adopted as a standard, based on historical use and convenience, since it is usually impractical to wait 20 or more days for the outcome of the test.

Why measure BOD?

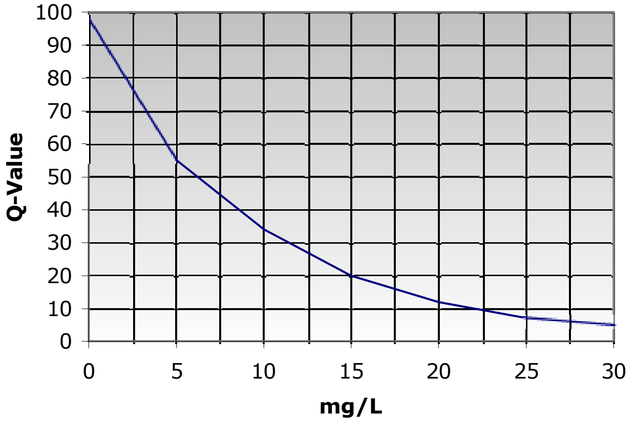
Biological oxygen demand is an important water quality parameter because it greatly influences the concentration of DO that will be in the water. The Q-value curve for BOD used in calculating a water quality index[2] illustrates the relationship between BOD and water quality—the higher the BOD, the lower the water quality, with 0 to 2 mg/L being associated with high water quality and values greater than 10 mg/L being associated with low water quality (refer to the BOD Test Results chart).

The BOD_5 test is used to measure the potential of wastewater and other waters to deplete the oxygen level of receiving waters. The test is also used to examine influents and effluents from wastewater processing facilities to compute the efficiency of operation of the treatment units.

The Q-curve for BOD used in calculating a water quality index[2] illustrates the relationship between BOD and water quality.

Chart 1

BOD Test Results



(Note: If BOD > 30, Q=2.0)

What factors affect BOD levels?

Human and animal waste in sewage is a significant contributor to elevated BOD levels. Runoff containing fertilizer from farms and other sources contributes to accelerated eutrophication, in which water bodies become choked with excessive plant growth, which contributes to the carbonaceous (organic) content of water bodies as these plants die and decompose.

Levels of BOD can be reduced by the introduction of low-BOD water from rain or snow melt. Additionally, when animals and plants die and settle on the bottom of the water body under conditions that permanently remove them from the water column, BOD levels can drop. Ultimately, river waters with high BOD discharge into the oceans where nutrients become highly diluted. The high BOD levels of untreated sewage are greatly reduced by wastewater treatment procedures.

What are ideal BOD values?

As shown on the BOD Q-curve, the less the BOD value, the better. High BOD levels will result in anoxic conditions, with the resulting growth of anaerobic microorganisms that produce noxious gases and cause the death of aerobic aquatic organisms. In the case of wastewater, the larger the drop in BOD following treatment procedures, the better. The regulated level of wastewater effluents depends on local conditions.

Overview of the two procedures for the 5-day BOD (BOD_5) test

About the official procedure

The procedure specified by the United States Geological Survey for the BOD_5 test is done in the laboratory and is quite complex to perform and difficult to control for reliable results. Among the considerations for a reliable test are the following:[1]

- The water sample must be obtained in a fashion such that it is representative of the larger water body.
- The test must be initiated within 2 hours of collecting the water sample, or the sample must be refrigerated (4 °C) for up to 48 hours before initiating the BOD_5 test.
- Special BOD bottles that minimize the accidental introduction of oxygen or nutrients must be used.
- Glassware and pipets must be scrupulously cleaned with a non-phosphate detergent and with multiple rinses with distilled or deionized water to avoid introducing nutrients or toxins.

- An aquarium pump, tubing, and air diffusion stone are needed, and they must be scrupulously cleaned as above.
- The pH must be maintained between 6.5 and 7.5 using 1 M NaOH and H_2SO_4, if necessary.
- A nutrient dilution solution containing $CaCl_2$, $FeCL_3$, $MgSO_4$, and phosphate buffer must be prepared using very high quality water that contains no nutrients or toxins. This solution supports the growth of the bacteria used in the assay. It should be discarded in the event that bacterial growth occurs in it.
- The DO sensor must be calibrated at the current atmospheric pressure and the temperature of the water sample before each day's measurements.
- The test samples must be incubated in the dark at 20 °C ± 1 °C.
- Any residual chlorine, if present, must be removed before the test.
- A sample that contains any toxic metals, arsenic, or cyanide must be specially treated before the test.
- A sample that has been treated by disinfectants may need to be "seeded" with bacteria.
- A negative (blank) and positive (glucose/glutamic acid standard solution) control must be run with the test, and the results of these must meet standards.

About the modified 5-day BOD (BOD_5) test for use in the classroom

The BOD_5 procedure presented in this field guide is a modification of the standard procedure that will produce results that are acceptable in a classroom environment and that support learning of the basic concepts of BOD_5 testing. It should not be used in a regulatory environment or when meaningful comparisons to test values obtained using the EPA-approved BOD_5 procedures[1] are needed. It is intended for use with natural water bodies, not treated sewage effluent. The following assumptions, which are reasonable for most natural water bodies and typical classroom lab conditions, apply to this procedure:

- there is an adequate population of aerobic bacteria in the test water;
- there are no toxins in the test water, test equipment, or dilution water that would prevent these bacteria from growing and multiplying;
- the pH of the test water is between 6.5 and 7.5;
- there is a negligible amount of nitrogenous material dissolved in the test water; and
- the ambient room temperature is about 20 °C and does not vary appreciably.

Modified BOD_5 procedure for the classroom

Materials List

Item	Quantity
Dissolved Oxygen Sensor*	1
Temperature Sensor* or Thermometer	1
Plastic lab bottle with screw top, 2-L	2
Plastic lab bottle with screw top, 1-L	1
Dark plastic or glass lab ("BOD") bottle with air-tight top, 300-mL	3
Graduated cylinder, 100-mL	2
Graduated pipet*	1
Wash bottle containing deionized water	1
Waste container	1
*The Water Quality MultiSensor may also be used. **Note:** If you are making your own nutrient dilution solution from scratch, you will also need the chemicals listed in the table below, a mass scale, and weighing supplies.	

Preparation for the BOD_5 test

1. **Prior to the procedure, ensure that all equipment has been thoroughly cleaned and rinsed, including final rinses with deionized water, and is dry.**
2. **The day before beginning the BOD_5 test, aerate 1 L of laboratory-quality distilled water and bring it to 20 °C.**

   Suggestion: To aerate, pour the deionized water into the 2-L bottle, cap, and shake vigorously. Then loosen the cap to allow contact with the atmospheric air.

   Note: It is very important that the distilled water used for the dilution water be of high grade and free from contaminants (such as copper and chlorine) that could inhibit the growth of bacteria. For this reason, it is recommended that ordinary commercial distilled water (for example, for use in car batteries) not be used.
3. **Collect at least 1 L of a representative sample of the natural body of water.**

   Note: If the BOD_5 test will be delayed by more than 2 hours, store the sample at 4 °C (not frozen). The BOD_5 test must be started within 48 hours.
4. **Approximately 1 hour before the BOD bottles are to be set up, prepare about 1 L of the BOD_5 dilution medium, using the distilled water prepared in Step 2 and prepackaged reagents (from retailers that sell water quality testing supplies; see the Recommended Reading and Resources section for suggestions) according to the manufacturers instructions, or using the procedure shown in the table below.**
5. **Store at room temperature in a container that does not have an air-tight lid (to allow the air dissolved in the medium to equilibrate with the atmospheric air).**

Preparing the nutrient buffer dilution solution

Solution	Procedure
Calcium chloride ($CaCl_2$) solution	Dissolve 27.5 g of $CaCl_2$ in deionized water and dilute to 1 L.
Ferric chloride ($FeCl_3$) solution	Dissolve 0.25 g of $FeCl_3 \cdot 6H_2O$ in deionized water and dilute to 1 L.
Magnesium sulfate ($MgSO_4$) solution	Dissolve 22.5 g of $MgSO_4 \cdot 7H_2O$ in deionized water and dilute to 1 L.
Phosphate buffer solution	Dissolve 8.5 g of KH_2PO_4, 21.8 g of $KHPO_4$, 33.4 g of $Na_2HPO_4 \cdot 7H_2O$, and 1.7 g of NH_4Cl in about 500 mL of deionized water. Dilute to 1 L.
Nutrient buffer dilution solution	To 1 L of the aerated deionized water prepared in Step 2, add 1 mL of each of the 4 solutions above and mix well.

*maximum contamination level (MCL) of either chloramines or free chlorine

Conducting the BOD_5 test

1. **Saturate the test water with air by pouring 1 L of it into a 2-L bottle, capping, and shaking vigorously.**
2. **Allow this water to sit undisturbed with the cap off until the air bubbles dissipate and the water temperature is the same as the room temperature (about 20 °C). If necessary, tap the container gently to dislodge air bubbles from the sides and bottom.**
3. **Using a Dissolved Oxygen Sensor that has just been calibrated �$^{(2.6)}$ and a data collection system, measure the DO in the test sample and record this value.**

   Note: This value should be approximately that of 100% DO saturation.

4. **Pour about 50 mL of the nutrient buffer dilution solution into each of the BOD bottles that will contain diluted test samples.**

   Note: Use a pouring method that minimizes the introduction of air bubbles, such as pouring slowly down the side of the bottle.

5. **From the table below, select a range of 3 dilutions that bracket the anticipated BOD of the test sample.**

   Suggestion: If you do not have an idea of this range, use the first 3 dilutions on the table. Using a graduated cylinder, measure each specified volume of test water and pour it into the BOD bottle as noted in Step 4.

6. **Fill the BOD bottles to the brim with the dilution water, and cap the bottle so it is air-tight, being careful to avoid trapping an air bubble.**
7. **Incubate the BOD bottles at room temperature for 5 days.**

   Note: If the BOD bottles are not completely opaque, incubate in a dark room or box.

8. **After 5 days, measure the DO level in each BOD bottle as in Step 3, and record the result.**

Example dilutions for the 5-day BOD test

Anticipated Range of the BOD_5 Value (mg/L)	Milliliters of Sample	Milliliters of Dilution Water
0–7	300	0
4–12	150	150
6–21	100	200
12–42	50	250
30–105	20	280
60–210	10	290
120–420	5	295
300–1,050	2	298
600–2,100*	1	299

* It is necessary to dilute the original test sample if it has a BOD that is greater than this range.

Calculating the 5-day BOD

1. **Select the dilution for which the 5-day DO reading dropped at least 2 mg/L from the first reading but remained above 1 mg/L.**

   Note: If more than 1 dilution fell into this category, calculate the 5-day BOD for each one, and then average the results for the final BOD_5 value.

2. **Use the following formula to calculate the final BOD_5 value:**

   BOD_5 (mg/L) = $(D_1 - D_2)/P$

   where

   D_1 is the initial DO of the sample

   D_2 is the final DO of the sample after 5 days, and

   P is the decimal volumetric fraction of sample used.

   For example, if a sample of 100 mL is diluted to 300 mL, then $P = 100/300 = 0.33$

   Note: If no dilution was necessary, then $P = 1.0$ and the BOD_5 is determined by $D_1 - D_2$.

More background information

More information about biological oxygen demand and its role in water quality may be found in the references listed below, with availability information listed in the Recommended Reading and Resources section.

References

1. G.C. Delzer, McKenzie SW. *Five-Day Biochemical Oxygen Demand*. In: DN Myers; FD Wilde, editors, translator and editor *National Field Manual for the Collection of Water-Quality Data*. Vol. 7.2, Techniques of Water-Resources Investigations Reports: U.S. Geological Survey; 1999.
2. National Science Foundation. *Water Quality Index*. 2004.

Chlorine

What is chlorine?

Elemental chlorine is a greenish-yellow gas (Cl_2) under standard conditions that is highly soluble in water and is poisonous. It is a powerful oxidant used in bleaching and disinfectants that reacts readily with many inorganic substances and all animal and plant tissues.

When chlorine reacts with the amines in proteins of plants and animals, it denatures them, altering their structures and causing enzymes to become inactive. The denaturing effect of chlorine on animal and plant tissues is the basis for its use as an effective water or wastewater disinfectant.

When chlorine dissolves in water, it hydrolyzes according to the following reaction:

$$Cl_2 + H_2O \rightarrow HOCl + H^+ + Cl^-$$

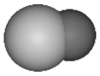
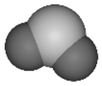
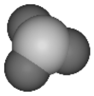
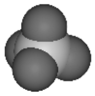
Unless the concentration of the chlorine solution is above 1000 mg/L, all chlorine will be in the form of hypochlorous acid (HOCl) or its dissociated ions, H^+ and OCl^- (hypochlorite ion), which are called free chlorine or free available chlorine.

Chlorine in the free, available form reacts readily with nitrogenous organic materials to form chloramines (mono- and dichloramines). When chlorine or hypochlorites are added to water containing nitrogenous materials, they rapidly form these chloramines, so chlorine toxicity in most natural waters is related to the chloramine concentration. The chloramines are called combined available chlorine.

The sum of free and combined available chlorines is called total residual chlorine (TRC) or total chlorine.[1]

Chlorine can assume oxidation states of –1, +1, +3, +5, or +7, corresponding to the anions Cl^-, ClO^-, ClO_2^-, ClO_3^-, or ClO_4^-, respectively known as chloride, hypochlorite, chlorite, chlorate, and perchlorate, as shown in the table below.

Oxidation states of chlorine

Oxidation State	–1	+1	+3	+5	+7
Anion Name	chloride	hypochlorite	chlorite	chlorate	perchlorate
Formula	Cl	ClO	ClO_2	ClO_3	ClO_4
Structure					

Hypochlorites are commonly used as bleaching agents including sodium hypochlorite (chlorine bleach) and calcium hypochlorite (bleaching powder). Some examples of common bleaching substances containing chlorite include sodium chlorite, $NaClO_2$, and magnesium chlorite, $Mg(ClO_2)^2$. Hypochlorous acid (HClO or HOCl) is swimming pool acid and is used as a bleach and as a disinfectant.

Why measure chlorine?

Free chlorine and chloramines are toxic to fish. Although free chlorine is slightly more toxic than chloramines, the difference in toxicity is not sufficient to merit differentiating between them. Thus, the toxicity to aquatic life of chlorine depends upon the concentration of total residual chlorine (TRC). Field studies have shown that total numbers of fish are drastically reduced and the stream bottom becomes clear of aquatic organisms downstream from facilities discharging chlorinated sewage effluents.

Chlorine in its various forms is used to disinfect water to be used as tap water or in industrial processes, and it is also used to disinfect sewage. However, some of these chlorine-based disinfectants have been shown to cause cancer and reproductive effects in lab animals, suggesting possible similar effects in humans.[2] The balance between adequate disinfection by chlorine compounds and adequate protection from the toxicity of chlorine compounds is challenging to achieve and requires close monitoring of levels of these disinfecting substances. The EPA has established regulations for the maximum residual disinfectant levels in drinking water for chlorine and several of its related compounds, i.e., chlorite, chloramines, and chloride dioxide (see ideal levels section).

What factors affect chlorine levels?

Chlorine is not a natural constituent of water. Free available chlorine and combined available chlorine appear in surface or ground water as a result of use of chlorine gas (Cl_2) or one of the chlorine compounds such as those listed above for disinfection of domestic sewage, to control organisms that grow in cooling water systems, or in industrial processes.

pH affects the ratio between HOCl (hypochlorous acid, the bactericidal agent in the use of chlorine for disinfection) and OCl (hydrochloric acid), with 96% HOCl remaining at pH 6, 72% at pH 7, 22% at pH 8, and 3% at pH 9. As HOCl dissociates into OCl^- and hydrogen ions, pH is lowered and alkalinity decreases.

Chlorine compounds can be removed from water using technologies such as filtration, enhanced coagulation, or enhanced water softening,[2] or through chemical reactions such as with sodium thiosulfate (NaS_2O_3), which reacts with the chlorine (or the chlorine portion of the chloramine) to form harmless chloride ions, a technique used when adding tap water to aquaria. Dissolved chlorine gas will slowly outgas from a warm body of water, particularly if the water is aerated.

What are ideal chlorine values?

In natural waters, the less total chlorine, the better. Species vary as to their sensitivity to chlorine, but all are sensitive to low levels; chlorine levels typically found in tap water are lethal to fish and many other aquatic organisms. Regarding freshwater fish, salmonids (salmon and trout) are the most sensitive, and sticklebacks and bluegill are more resistant to the toxicity of chlorine. Among freshwater invertebrates, snails, daphnia, and rotifers are quite sensitive and crayfish are more resistant.

The most sensitive forms exhibit signs of toxicity at very low concentrations. For example, the LC50 (concentration that was lethal for 50% of the population) for a rotifer (tiny freshwater invertebrate) was found to be 13 µg/L (13 parts per billion) of total residual chlorine (TRC), and juvenile Coho salmon exhibited a reduced growth rate when exposed to 11 µg/L (11 parts per billion) of TRC. Even more resistant species have been reported to be adversely affected at relatively low concentrations. For example, crayfish, which are considered one of the more resistant organisms, were found to have an LC50 of 31 µg/L when exposed for 365 days, because this length of exposure included molting stages at which the crayfish are most sensitive to chlorine.[3] Note that these lethal concentrations of TRC are more than 100 to 2000 times more dilute than the regulated level for tap water as discussed below.

The US EPA has established criteria for levels of chlorine that are protective of salmonids and other aquatic life, and chlorine is regulated in the EPA drinking water standards.[4] The basic criteria and standards are listed in the table below; however, refinements and modifications that apply to local variations of conditions have more recently been adopted.[2,3]

Maximum safe levels for total residual chlorine for various designated uses. [1]

Designated Use	Maximum Safe Level of Total Residual Chlorine
Freshwater aquatic life (salmonids)	2 μg/L[1]
Freshwater aquatic life (non-salmonids)	10 μg/L[1]
Marine aquatic life	10 μg/L[1]
Tap water	4 μg/L[4]*

*maximum contamination level (MCL) of either chloramines or free chlorine

Other chlorine substances regulated by the EPA in drinking water include chlorite (maximum contamination level, MCL, of 1.0) and chlorine dioxide (MCL of 0.8 mg/ L).[2,4]

Further information can be found in the referenced material listed below (refer to the Recommended Reading and Resources section for more information).

Measuring Procedure

The measuring procedure for chlorine uses the Water Quality Colorimeter and the Chlorine ezSample Test Kit as follows:

1. **Plug the Water Quality Colorimeter into the data collection system,** �(2.1) **and turn on the system.**

   Task Result: On some data collection systems, a digits display will appear, by default, for ammonia.

2. **Calibrate the Water Quality Colorimeter.** �(3.1)

Taking a Chlorine Measurement

SAFETY PRECAUTIONS

- Wear safety glasses and protective gloves.
- Review the MSDS for each of the Chlorine ezSample reagents, and have them available for reference.

1. **Follow the included Test Procedure instructions for the Chlorine ezSample Test Kit for sample preparation (Steps 1 through 6).**
2. **If required for your data collection system, begin by building an experiment. Otherwise, set the display to show Chlorine measurements.** �(2.2)
3. **Set your data collection system to Manual Sampling.** �(2.3)
4. **Place the prepared test ampoule in the Water Quality Colorimeter and cover it with the black cap.**

   Note: Handle the ampoule by the tip, and wipe the outside glass lens clean with a non-abrasive cleaning tissue.

   Task Result: The chlorine reading will automatically appear in the display. There may be slight fluctuations, so wait until the reading settles around a point.

5. **Record the chlorine reading.** �(2.4)

   Note: The units *mg/L* and *ppm* (parts per million) are equivalent.

Disposal of the test vial

1. **The ampoules may contain very low levels of potentially toxic substances (see the MSDS sheet), so check with your local regulatory agency regarding disposal procedures.**
2. **Carefully pour the test water down the drain. Shake the glass tip remaining in the sample cup into the solid waste container.**
3. **Thoroughly rinse the sample cup with deionized or distilled water.**

Accuracy, resolution, and practical detection limit (PDL)

Accuracy may be compromised if test results are outside the stated test ranges. The lower limit of the stated test range is the “practical detection limit (PDL),” defined as the lowest concentration at which less than ± 30% error is routinely obtained.

For the Chlorine ezSample Test Kit, the range, accuracy, PDL, and resolution are shown on the table below.

> **Note**: If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

Chlorine ezSample Test Kit: accuracy, resolution, and PDL (mg/L)

Range	Accuracy	PDL	Resolution
0–6	0.1	0.5	0.1

References

1. *Quality Criteria for Water (Red Book).* Washington, D.C.: U.S. Environmental Protection Agency; 1976.
2. *Stage 1 Disinfectants and Disinfection Byproducts Rule: A Quick Reference Guide*. U.S. Environmental Protection Agency, Office of Water; 2001.
3. Brungs WA, Middah DP. *Ambient Water Quality Criteria for Chlorine–1984.* U.S. Environmental Protection Agency; 1985.
4. *List of Drinking Water Contaminants & MCLs*. National Drinking Water Standards. U.S. Environmental Protection Agency; 2006.

Conductivity

What is conductivity?

Conductivity is the measure of water's ability to conduct electricity, depending on the concentration of dissolved ions in the water. Dissolved ions in natural waters consist mainly of some combination of the following: calcium (Ca^{2+}), magnesium (Mg^{2+}), sodium (Na^{+}), aluminum (Al^{3+}), potassium (K^{+}), carbonate (CO_3^{2-}), bicarbonate (HCO_3^{-}), phosphate (PO_4^{3-}), chloride (Cl^{-}), nitrate (NO_3^{-}), and sulfate (SO_4^{2-}).[1]

Conductivity is commonly measured using a specially designed probe and is expressed in units of microsiemens per centimeter (µS/cm). Conductivity is the reciprocal of resistivity—a measure of how resistive water is to conducting an electrical current. Thus, conductivity can also be expressed in units of mhos/cm as the reciprocal of resistance measured in ohms.[1] Mhos is the same as seimens and is the reciprocal of ohms. The term mhos comes from the term ohms, spelled backwards.

Why measure conductivity?

Measuring conductivity is a quick and easy way to estimate the amount of total dissolved solids (TDS) in natural waters, since most of these solids dissolve to form ions. Whereas total dissolved solids can include organic as well as inorganic molecules, conductivity depends only upon dissolved ions.

Conductivity measurements can also be a useful tool for monitoring the inflow of saline water in estuaries and identifying sources of pollution, such as mining or industrial waste or agricultural runoff. Conductivity is an important factor in soil analysis, as transport of micronutrients through plant roots can be affected by the conductivity of the soil. The salinity of soil dramatically affects the types of plants that can grow in soil, and can be measured indirectly with a Conductivity Sensor.

Municipal water supplies are monitored for conductivity, as increased dissolved solids in the water supply can create hard water, add scale build-up to plumbing, and change the taste of the water.

Conductivity analysis is important in aquariums and fish spawning habitats, as many species are sensitive to abrupt changes in the salinity of their environment.

Desirable conductivity levels, by designated use

Designated Use	Conductivity (µS/cm at 25°C)	Reference
Drinking water	Less than 750* (lower levels—400 µS/cm or lower—are better)	EPA Secondary Drinking Water Standard[3]
Freshwater aquatic life	150 to 500**	Ref[1]
Marine aquatic life	50,000	Ref[1]
Agricultural irrigation***	Less than 750*	Ref[3]

*As computed from 500 ppm TDS, using the 0.67 conversion factor;

**Many species can tolerate much higher levels[3]

***Many plants can grow at much higher levels (see the accompanying table in this section).

Dissolved solids hazard for irrigation water[3]

Effect	Conductivity (µS/cm)*
No adverse effects	Less than 750**
Adverse effects on sensitive crops	750–1500
Adverse effects on many crops unless carefully managed	1500–3000
Adverse effects for all plants except for some tolerant species, even with careful management	3000–7500

*As computed from 500 ppm TDS, using the 0.67 conversion factor

**less than 250 µS/cm is best[4]

More specific measurements may be necessary

A high conductivity determination may indicate the need to conduct more specific analyses of specific ions to determine whether the water contains high levels of an ion that would make it unsuitable for a designated use. For example, water that is high in sodium is unsuitable for drinking by individuals who have difficulty metabolizing high levels of sodium, such as some people suffering from cardiac disease or toxemia of pregnancy.[3]

Water that has a high level of aluminum (along with a low pH) may inhibit plant growth.

Water that has very high levels of nutrients, such as phosphate or nitrate, is said to have undergone eutrophication, which stimulates plant growth, followed by an undesirable drop in dissolved oxygen content when this excess plant life dies, supporting high levels of bacterial respiration. The pH of water that is very low in conductivity is very sensitive to added acids, such as occurs with acid rain.

What factors affect conductivity?

Conductivity is determined by several factors related to the concentration, degree of dissociation, valencies, and mobilities of ions in solution; thus the temperature of the solution also plays an important part.

The watershed of a body of water is the primary contributor of the ions that affect the conductivity of water. Melted snow running through granite, such as that in high mountain streams, has low levels of conductivity. Conversely, water that runs off from limestone or agricultural regions has higher levels of conductivity. Ground water, such as that pumped from wells, can have high conductivity levels because of long-time exposure to rock that contributes ions. Water in arid regions can have high conductivity because of high levels of evaporation. Infiltration of sea water can raise conductivity, as can inflows of certain mining or industrial wastes.

Temperature effects must be considered

Temperature affects conductivity; each increase of 1 °C causes an increase of about 2% in the conductivity. Therefore, when making comparisons between measurements, temperature should be controlled or accounted for. This is particularly important for measurements in mountain streams, where the temperature effect would be large.

What are desirable levels of conductivity?

The range of conductivity levels that are considered good depends upon the designated use of the water. For example, low conductivity is desirable in drinking water. Natural waters with low conductivity are very susceptible to changes in pH in response to added acids or bases, and they have low nutrient value for aquatic organisms. Irrigation water with high conductivity levels reduces plant growth and yield (see the tables above). Marine fish require water with high conductivity (salinity) to survive.

Additional information regarding conductivity and water quality is available from the cited references. [1–4]

Measuring Procedure

Plug the Conductivity Sensor into the data collection system �(2.1), and turn on the system.

Note: The Water Quality Sensor or the Salinity Sensor can also be used.

Calibration note

The Conductivity Sensor does not need to be calibrated for general studies. Press the button for the µS/cm range of your sample.

If it is found that the Conductivity Sensor factory calibration is inadequate, calibrate the Conductivity Sensor using a standard value �(2.5).

for 0–1,000 µS/cm

for 1,000–10,000 µS/cm

for 10,000–100,000 µS/cm

Taking a conductivity measurement

1. **If required for your data collection system, begin by building an experiment. Otherwise, set the display to show conductivity measurements.** �(2.2)
2. **Set your data collection system to Manual Sampling.** �(2.3)
3. **Immerse the end of the Conductivity or Salinity probe into the water to be measured. When the reading stabilizes, record the conductivity reading.** �(2.4)

   Note: Be sure to immerse the holes near the end of the sensor. Swirl the probe to be sure any air bubbles are removed from the end of the probe.
4. **Take additional measurements as needed, following the procedure above.**

More background information

More information about conductivity and its role in water quality may be found in the references listed below, or in the Recommended Reading and Resources section.

References

1. *Healthy Water, Healthy People Testing Kit Manual*. 2002, Bozeman, Montana: The Watercourse, International Project WET.
2. *National Secondary Drinking Water Regulations. Current Drinking Water Standards*. 2006. Environmental Protection Agency.
3. *Quality Criteria for Water (Red Book)*. 1976, Washington, D.C.: U.S. Environmental Protection Agency.
4. Stover, H. S. and S. L. Seager, *Environmental Chemistry: Air and Water Pollution*. (2nd ed.) 1976, Glenview, Illinois: Scott, Foresman, and Co.

Dissolved Carbon Dioxide

What is dissolved carbon dioxide (CO_2)?

Dissolved carbon dioxide is carbon dioxide gas (CO_2) that is held in water. It is normally measured in terms of ppm (mg/L). Dissolved CO_2 levels are commonly below 10 mg/L in natural surface waters, but groundwater levels can be up to several hundred mg/L.

Why measure dissolved carbon dioxide?

Dissolved CO_2 is required by aquatic plants during photosynthesis to create energy-rich molecules such as sugars and starches, as well as structural molecules such as cellulose. In the absence of dissolved CO_2, aquatic plants will turn yellow and become sick or die. However, too much CO_2 is toxic to fish, since a high concentration of dissolved CO_2 causes fish to have difficulty releasing CO_2 from their bodies, which is necessary for sustaining life. Fish exposed to high dissolved CO_2 levels become lethargic and do not feed; however, they recover after dissolved CO_2 levels drop.

High levels of dissolved CO_2 can lower the pH to unhealthy levels, particularly in water with low alkalinity, through the following chemical reaction:[1,2]

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + CaHCO_3^-$$

Measuring dissolved CO_2 can demonstrate its diurnal rhythm (see the graph of dissolved oxygen and dissolved carbon dioxide below) in lakes, ponds, and slow-moving rivers.

Additional reasons for measuring dissolved CO_2 include the following:

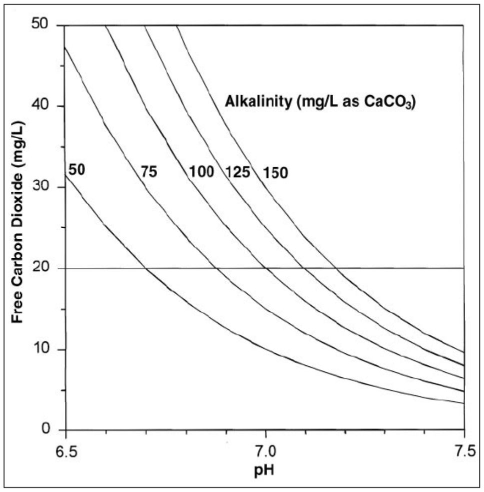
- Measurement of dissolved CO_2 can also signal the need for remedial action to reduce its concentration in fish tanks and ponds used for aquaculture.[1]
- Levels of dissolved CO_2 have been used to assess the relative health of marine estuaries.[2] Lakes and ponds overlying volcanic activity can undergo a significant build-up of dissolved CO_2 as a result of outgassing of CO_2 from magma and the formation of carbonic acid in the deep water. In 1986, a sudden release of CO_2 from the depths of Lake Nyos in Cameroon, Africa, resulted in human deaths,[3] signaling a need to monitor such lakes for buildup of dissolved CO_2.
- A new technology for treating waters contaminated with acid mine drainage involves the dissolution of limestone particles using pressurized CO_2, which is potentially hazardous to fish health and thus requires careful monitoring of dissolved CO_2 levels.[4]

What factors affect dissolved carbon dioxide levels?

Physical factors that affect dissolved CO_2 levels include temperature, pH, alkalinity, and the turbulence and surface-to-air ratio of the water. Cold water can hold more dissolved CO_2 than warm water, which is also true for dissolved oxygen (see the Dissolved Oxygen section45). Dissolved CO_2 levels are related to the pH, as previously discussed. The alkalinity affects dissolved CO_2 levels through the following chemical reaction:[2]

$$CaCO_3 + CO_2 + H_2O \rightleftharpoons Ca_2^+ + 2HCO_3^-$$

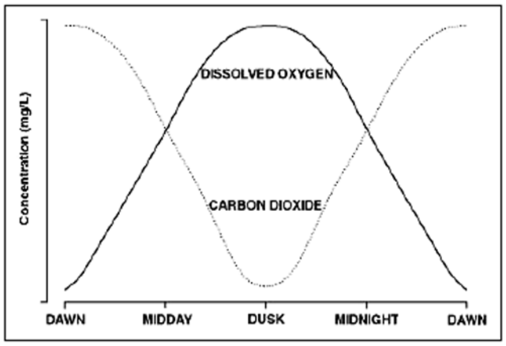
Dissolved CO_2 levels, pH, and alkalinity are interrelated.[1]



The "Free Carbon Dioxide" figure illustrates the relationship between dissolved CO_2, alkalinity, and pH. Note that at low levels of alkalinity, increased dissolved CO_2 levels cause greater decreases in pH than at higher levels of alkalinity, which buffer against changes in pH.

In fast-moving, shallow streams, the primary source of dissolved CO_2 is CO_2 gas from the air, which is readily dissolved in the turbulent waters until it reaches the saturation level; then gas exchange with the air keeps the dissolved CO_2 at a relatively constant level.[5] The large surface-to-air ratio of these streams also increases the exchange of dissolved CO_2 with CO_2 gas from the air.

Biological factors affecting dissolved CO_2 are more predominant in slow-moving or stagnant bodies of water. In lakes, ponds, and slow-moving rivers, dissolved CO_2 levels are related primarily to the process of photosynthesis by aquatic plants and phytoplankton—which causes a reduction in dissolved CO_2 levels during the day—and the process of cellular respiration by aquatic organisms—which causes an increase in dissolved CO_2 levels during the night, resulting in diurnal fluctuations in the dissolved CO_2 and dissolved oxygen levels as shown in the figure below.



In lakes, ponds, and slow-moving rivers, dissolved CO_2 and dissolved oxygen fluctuate inversely in a diurnal rhythm (once per 24 hours).

These fluctuations can become quite pronounced in water that has a large amount of plant and phytoplankton growth, such as under conditions of algae bloom. Even under normal conditions, the dissolved CO_2 concentration in a pond or lake can range from 0 mg/L during the late afternoon to 10 to 15 mg/L at dawn.[1]

What are ideal dissolved carbon dioxide levels?

Dissolved CO_2 seldom reaches sustained levels that are harmful to fish, especially at cooler water temperatures when levels of dissolved oxygen are higher, and it seldom drops to levels that do not support photosynthesis. In aquaculture, the danger level at which remedial action, such as supplemental aeration to promote outgassing of dissolved CO_2 to the air or even adding chemicals such as quicklime (CaO) to the water to reduce dissolved CO_2 levels, is about 20 mg/L.[1] This level can easily be reached during summertime algal blooms, which are also followed by algal die-off and increased bacterial respiration, causing depletion of dissolved oxygen levels and decreased pH levels. Under these conditions, it can be difficult to determine which factor is harming the aquatic organisms.

Levels of dissolved CO_2 that are well tolerated by fish vary according to the species. For example, trout are adversely affected at relatively low levels of dissolved CO_2, whereas catfish can tolerate high levels, up to 60 mg/L of dissolved CO_2 if the dissolved oxygen levels are high.[4,6]

Mobile organisms can move to areas in the water column that have lower dissolved CO_2 concentrations.[4] Therefore, when deep lakes develop high dissolved CO_2 levels in deeper water during calm conditions, mobile organisms can migrate to shallow water to escape to regions with less dissolved CO_2.

Further information is available from the references; also refer to the Recommended Reading and Resources section.

Measuring Procedure

The measuring procedure for dissolved carbon dioxide uses the CO_2 ezSample Field Titrator as detailed in the instruction card included with the kit.

> **Note:** If the water is turbid, it should be filtered before testing it. However, for most applications, the sediment can be allowed to settle for several hours instead being removed by filtration.
>
> **Note:** The test range of the dissolved CO_2 ezSample Field Titrator is 10–100 ppm (mg/ L) CO_2. Because the ampoules in the CO_2 ezSample Field Titrator kit have nonlinear scales, the accuracy of these kits varies with the analyte concentration. At the low end of the test range, the accuracy is ± 5%. At the high end of the range, the accuracy falls to ± 20%.
>
> **Note:** If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

References

1. Hargreaves J, Brunson M. *Carbon Dioxide in Fish Ponds.* Southern Regional Aquaculture Center. SRAC Publication No. 468, 1996.
2. Ford P. *Water Column Partial Pressure of Carbon Dioxide.* OzEstuaries, Geoscience Australia; 2006.
3. *Volcanic Lakes and Gas Releases.* U. S. Geological Survey/Cascades Volcano Observatory; 2001.
4. Ross R, Krise W, Redell L, Bennett R. *Effects of Dissolved Carbon Dioxide on the Physiology and Behavior of Fish in Artificial Streams.* Environ Toxicol. 2001;16:84–95.
5. *Dissolved Gases.* Water on the Web; 2004.
6. *Catfish: Water Quality*. Mississippi State University Extension Service; 2006.

Dissolved Oxygen

What is dissolved oxygen?

Dissolved oxygen (DO) is simply oxygen gas (O_2) that is dissolved in water. When an oxygen gas molecule encounters a water molecule, the polar water molecule induces a weak dipole on the oxygen gas molecule; thus the oxygen molecule becomes weakly attracted to the water molecule, forming a weak bond that holds the oxygen gas molecule in solution.[1]

Since kinetic energy can overpower these weak bonds, temperature is an important factor regarding the amount of DO that water can hold—the lower the temperature of liquid water, the more DO it can hold. The amount of DO that fresh water can hold (100% saturation at sea level) can vary from about 14.5 mg/L at 0 °C to 5.5 mg/L at 50 °C.

The partial pressure of oxygen is another important factor that determines how much DO water can hold. Therefore, the higher the air pressure (barometric pressure), the more DO can be held by liquid water. For example, water that is 10 °C can hold about 11 mg/L DO at sea level (barometric pressure, 760 mg Hg) but only about 9.3 mg/L DO at an altitude of 1 mile (630 mg Hg, for example, near Denver, Colorado), and only about 7.5 mg/L DO at an altitude of about 2 miles (510 mg Hg, for example, near Wilkerson Pass, Colorado). For an abbreviated table of the DO at 100% saturation, please see the table at the end of this section. For complete tables and for measuring DO in saline water, please see the User Manual for the Dissolved Oxygen Sensor.

Why measure dissolved oxygen?

Dissolved oxygen is a critical factor determining water quality, affecting the growth and health of aquatic communities and the adequacy of water to be used for drinking and municipal purposes.

Most aquatic organisms require dissolved oxygen to live, with some organisms requiring higher concentrations than others (see the table of desirable dissolved oxygen values, below, for examples).

Very high levels of DO can cause pipe corrosion, adding unwanted iron or other metals to the water, causing a deterioration in taste and clarity, which is a consideration for designated use of drinking water.[2] On the other hand, water that has very low levels of DO can cause the leaching of iron and manganese from sediments, which can affect the taste and color of water, in addition to staining plumbing fixtures or other materials that the water contacts.[2]

The presence of DO prevents the occurrence of anaerobic conditions that support the growth of anaerobic bacteria, which produce noxious or malodorous gases, such as methane (CH_4) or hydrogen sulfide (H_2S).

A special test for the presence of biologically available organic materials, called the 5-day biological oxygen demand (BOD_5) test, uses DO as an indicator. The BOD_5 test measures the amount of oxygen consumed by bacteria from the decomposition of organic matter.[4] Water that has a high level of biologically available organic materials, such as dead plant and animal tissue or ammonia, requires more chlorine to achieve sanitation. It will also fall significantly in DO concentration as time goes on, due to bacterial respiration.

Both DO and the BOD_5 are included in all of the water quality indexes.

What factors affect dissolved oxygen?

Many factors affect the concentration of dissolved oxygen in water, and can work in combination to produce unexpected results. It is important to understand the effect of each of these parameters on dissolved oxygen content.

Parameter	Effect
Temperature	Cold water can hold more DO than warm water can hold. Seasonal variations in water temperature will affect DO. The temperature of the water at the site of the sampling must be measured.
Atmospheric Pressure	Water can hold more DO at high pressure than it can at low pressure at the same temperature. Therefore the barometric pressure of the air at the site of the sampling must be taken. Even small changes in barometric pressure can have an effect on the amount of measurable DO.
Photosynthesis	Increases in photosynthesis cause increases in DO concentrations.
Respiration	Increases in respiration cause decreases in DO concentrations.
Water-mixing characteristics	Deep stagnant water develops regions of very low DO concentration at the bottom, whereas shallow, swiftly moving, turbulent water has high DO concentrations
Wind	Causes increased DO concentrations
Biologically available organic materials	Causes decreased DO concentrations
Nitrates	Fertilizers or animal waste cause decreased DO concentrations, or transient increases followed by decreases.
Salinity	Water with high salinity, such as ocean water, holds slightly less DO at 100% saturation than fresh water.
Chemicals	Sodium sulfite (Na_2SO_3), for example, is a chemical widely used in photographic development, industrial dyeing operations, and other uses, which causes dramatic reductions in DO concentrations.

What are ideal dissolved oxygen values?

The EPA states that natural waters should contain sufficient dissolved oxygen to maintain aerobic conditions in the water column, and except as affected by natural phenomena, at the sediment-water interface. For freshwater aquatic life, the minimum requirement is for 5 mg/L for maintenance of good fish populations, with some fish, such as salmon, requiring higher levels, especially for embryonic and young forms. These forms exist in the silt and gravel at the bottom of the waterway, so higher levels of DO are needed to penetrate into these areas for adequate oxygenation. Although adult aquatic animals often can live for short times at much lower levels of DO, they do not grow and thrive at these low levels.[2]

DO concentrations can be expressed both in mg/L (or ppm) and as percent of saturation. The water quality indices use percent saturation as a standard, with 100% being most desirable (see the Q-value figure below).[5]

Desirable dissolved oxygen (DO) values, by designated use[2,4]

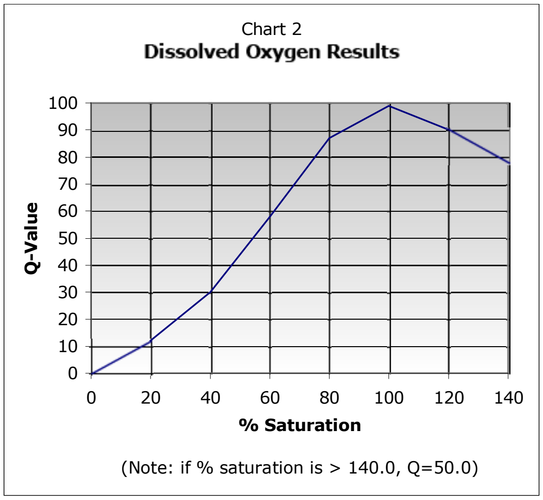
Designated Use	DO Value
Maintenance of good fresh water fish* populations	5.5 mg/L or higher (minimum, 5 mg/L)**
Maintenance of salmonids (salmon, trout)	6 mg/L or higher (minimum, 5 mg/L
Supporting reproduction and early life stages of salmonids	9 mg/L or higher (minimum, 8 mg/L)
Supporting diverse invertebrate life	5 mg/L or higher**
Drinking water	100% saturation or less

*Not salmon or trout.

**Many species can live for short periods of time at lower levels.

Water that has a DO concentration greater than 100% is termed supersaturated. Supersaturation with DO can be a symptom of eutrophication, that is, enrichment with nutrients such that algae and other plant life are stimulated to grow rapidly, with photosynthesis causing large amounts of O_2 to be released into the water. Since this rapid growth is followed by large-scale death and decay once the nutrients are used up, supersaturated waters can rapidly become deficient in DO or even anaerobic, causing fish kills and allowing the growth of anaerobic bacteria that excrete noxious or malodorous gases. Water that is supersaturated with DO is less desirable as drinking water because the supersaturated water is corrosive to pipes.[2]

Q-value curve of dissolved oxygen saturation levels used in calculating a water quality index[5]



(Note: if % saturation is > 140.0, Q=50.0)

The Q-value curve for DO that was established by the National Sanitation Foundation for their water quality index determination procedure illustrates that water that is significantly supersaturated with DO is as undesirable as water that low in DO concentration. The higher the Q-value, the better the water quality.

More background information

More information about dissolved oxygen and its role in water quality may be found in the references listed below; also refer to the Recommended Reading and Resources section.

Measuring Procedure

Setting up the Dissolved Oxygen Sensor

1. **Plug the Dissolved Oxygen Sensor into the data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for dissolved oxygen.

   Note: The Water Quality Sensor can also be used.

2. **Remove the storage bottle from the end of the probe, being careful not to touch the membrane at the end of the probe.**

   Note: You can simply unscrew the soaker bottle and remove the bottom part, leaving the lid attached to the probe.

3. **Calibrate the Dissolved Oxygen Sensor.** �(2.6)

   Note: It is necessary to calibrate the DO sensor at the exact temperature of the water sample and the barometric pressure at the sampling site. Failure to do this will result in inaccurate test results. Even slight variations in barometric pressure can affect results. It is not recommended that the atmospheric pressure be taken from local weather averages, such as are reported in newspapers. Use of a barometric sensor or weather sensor is strongly encouraged. For best results, use the complete DO Calibration Tables included with the sensor.

   Note: Probes of this kind have their electrodes housed in a fluid which contains oxygen. It is beneficial for these sensors to "burn off" this small amount of oxygen before they are calibrated. Plug in and power on the sensor and let it rest for a short amount of time. It would be useful to build a graph and watch the dissolved oxygen in the sensor tip get used up. When this process is complete, calibrate the sensor in the standard way.

Taking a Dissolved Oxygen Measurement

1. **If required for your data collection system, begin by building an experiment.** �(2.2)
2. **Set your data collection system to Manual Sampling.** �(2.3)
3. **Insert the Dissolved Oxygen Sensor into the solution.**
4. **Record the dissolved oxygen reading.** �(2.4)

   Important! Ensure that the Dissolved Oxygen Sensor has been maintained properly and that the membrane at the end of the probe is intact. If the membrane is torn, replace it before use. (Refer to the Dissolved Oxygen Sensor User Manual for details.)

Selected reference values for 100% saturation of dissolved oxygen by altitude or barometric pressure and temperature[1]

Altitude	Typical Barometric Pressure		DO at 100% Saturation, °C mg/L							
feet	in Hg	mm Hg (torr)	0	5	10	15	20	25	30	35
0	29.92	760	14.6	12.8	11.3	10.1	9.2	8.3	7.6	7.1
1,000	28.86	733	14.1	12.3	10.9	9.8	8.8	8.0	7.3	6.8
2,000	27.82	707	13.6	11.9	10.5	9.4	8.5	7.7	7.1	6.5
3,000	26.82	681	13.1	11.4	10.1	9.1	8.2	7.4	6.8	6.3
4,000	25.84	656	12.6	11.0	9.7	8.7	7.8	7.1	6.5	6.0
5,000	24.89	632	12.2	10.6	9.2	8.4	7.6	6.9	6.3	5.8
6,000	23.98	609	11.7	10.2	9.0	8.1	7.3	6.6	6.0	5.6
7,000	23.09	586	11.3	9.8	8.7	7.8	7.0	6.4	5.8	5.4
8,000	22.22	564	10.8	9.4	8.4	7.5	6.7	6.1	5.6	5.1
9,000	21.38	543	10.4	9.1	8.0	7.2	6.5	5.9	5.4	4.9
10,000	20.57	522	10.0	8.7	7.7	6.9	6.2	5.6	5.1	4.7
11,000	19.79	503	9.7	8.4	7.4	6.6	6.0	5.4	4.9	4.5

References

1. *PASPORT Dissolved Oxygen Sensor User Manual.*
2. *Quality Criteria for Water* (Red Book). 1976, Washington, D.C.: U.S. Environmental Protection Agency.
3. G.C. Delzer and S. W. McKenzie, *Five-Day Biochemical Oxygen Demand*, in National Field Manual for the Collection of Water-Quality Data, D. N. Myers and F. D. Wilde, Editors. 1999, U.S. Geological Survey.
4. Munson, B., R. Axler, C. Hagley, G. Host, G. Merrick, and C. Richards. *Dissolved Oxygen. Why Is It Important?* Water on the Web—Monitoring Minnesota Lakes on the Internet and Training Water Science Technicians for the Future—A National On-line Curriculum using Advanced Technologies and Real-Time Data. 2004, University of Minnesota–Duluth.
5. *National Sanitation Foundation Water Quality Index.*

Hardness

What is hardness?

Water hardness is caused by metal cations that are dissolved in water, mainly those of calcium (Ca^{2+}) and magnesium (Mg^{2+}) but also those of aluminum (Al^{3+}), barium (Ba^{2+}), iron (Fe^{2+}, Fe^{3+}), manganese (Mn^{2+}, Mn^{3+}), strontium (Sr^{2+}), and zinc (Zn^{2+}) when these are present in appreciable concentrations.[1] The primary negative ions that contribute to hardness are bicarbonate (HCO_3^-) and carbonate (CO_3^{2-}). These ions are important components of the total dissolved solids (TDS) (see also the section on conductivity). Hardness is commonly reported as an equivalent concentration of calcium carbonate ($CaCO_3$).

The concept of hardness comes from practical considerations. Hardness is measured by soap requirements for adequate lather formation and as an indicator of the rate of scale formation in hot water heaters and low-pressure boilers.[2] A commonly used classification is shown in the following table[1]:

Classification of water by hardness content*

$CaCO_3$ Concentration mg/L	Description
0–60	soft
61–120	moderately hard
121–180	hard
181 and greater	very hard

*Some organizations use other classification schemes

Hardness is related to TDS but is not exactly the same, since ions like sodium (Na^+) are important contributors to TDS but not to hardness. Generally, hard water has a high level of TDS, but water with a high level of sodium ions, and thus, a high level of TDS, is not necessarily hard.

Why measure hardness?

Hardness is a quality of water that particularly affects water that is designated for use as tap water or for industrial use. As mentioned in the previous section, tap water that is hard requires more soap to achieve cleansing lathers.

More seriously, the calcium and magnesium carbonates tend to precipitate out as adherent solids, or scale, on the surfaces of pipes and especially on the hot heat exchanger surfaces of boilers. The resulting scale buildup can impede water flow in pipes (see picture). In boilers, the deposits act as thermal insulation that impedes the flow of heat into the water, reducing heating efficiency and causing the metal to overheat, which in pressurized systems can lead to catastrophic failure.

Scale can shorten the useful life of water pipes and water heaters.[5] If hard tap water is identified, water softening procedures can be applied to lessen the negative impact of hard water. By measuring the amount of hardness, a determination of the chemical dosages for soda-lime water softening procedures or effective ion-exchange softening strategies can be determined.[2]

What factors affect hardness values?

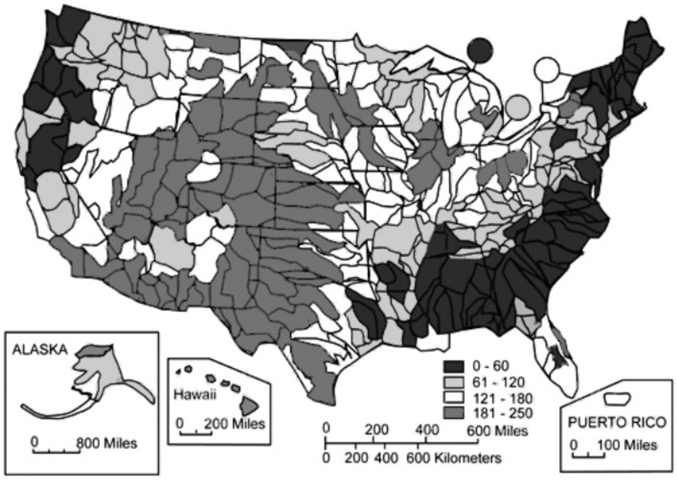
Most of the minerals that contribute to hardness come from the ground. When water passes over certain kinds of bedrock or soil, it dissolves minerals in it and carries the ions with it.[1] Some minerals that contribute to hardness as well as their ions when dissolved are shown listed on the table below.

Some minerals that contribute to hard water[4]

Mineral	Chemical Formula	Ions When Dissolved
Limestone/Chalk	$CaCO_3$	Ca^{2+}, CO_3^{2+}
Magnesite	$MgCO_3$	Mg^{2+}
Gypsum	$CaSO_4{\cdot}H_2O$	Ca^{2+}, SO_4^{2}
Dolomite	$CaCO_3{\cdot}H_2O$	Ca^{2+}, Mg^{2+}, CO_3^{2}

Areas with a lot of rock containing these minerals, also known as carbonate rock, often have hard water. On the other hand, areas having primarily granite rock, which does not contain these minerals, will have soft water. The map shows the results of hardness testing conducted by the USGS National Stream Quality Accounting Network in 1975.[3]

Concentration of hardness as calcium carbonate (mg/L)



The hardness of tap water may be influenced by the source of that water. For example, in one area in California, the water in one community is soft because it is taken from a river with headwaters in granite mountains, whereas the water in a community only 30 miles away is hard because it is taken from ground water in a valley area containing an abundance of carbonate rock.

Other sources of hardness include mining and rock quarry activities that expose rocks containing calcium and magnesium, as well as some types of pollution from industrial discharge, some types of cleaning agents, and human and animal waste.

What are ideal hardness values?

Water hardness, per se, does not affect health, so no public health standard has been established. The effects of hardness on fish and other aquatic life are related to the specific ions that contribute to the hardness, rather than the hardness itself, so again no standards have been established for freshwater aquatic life. However, in some instances, very soft water can contribute to the toxicity of copper and zinc for fresh water organisms.[2]

For industrial use, the maximum tolerable level of hardness is quite variable (see the table below) and is related to the cost of reducing hardness to the level required for the industrial process.[2]

Hardness is not a determination of concern for water used for irrigation. However, the concentration of the cations of calcium and magnesium that comprise hardness are a matter of concern.[2] Specific testing for these ions may be necessary when hard water is used for irrigation.

Further information can be found in the referenced material listed below.

Maximum hardness levels accepted by industry as a raw water source*

Industry	Maximum Concentration (mg/L as $CaC0_3$)
Electric Utilities	5,000
Textile	120
Pulp and paper	475
Chemical	1,000
Petroleum	900
Primary metal	1,000

*Requirements for final use within a process may be essentially zero, which requires treatment for concentration reduction; source EPA Redbook.[2]

Measuring Procedure

The measuring procedure for total hardness involves the Total Hardness ezSample Field Titrator as detailed in the instruction card that is included with the kit.

Note: If the water is turbid, it should be filtered before testing it. However, for most applications, the sediment can be allowed to settle for several hours instead being removed by filtration.

Note: The test range of the Total Hardness ezSample Field Titrator is 20–200 ppm (mg/ L) $CaCO_3$. Because the ampoules in the Total Hardness ezSample Field Titrator kit have nonlinear scales, the accuracy of these kits varies with the analyte concentration. At the low end of the test range, the accuracy is ± 5%. At the high end of the range, the accuracy falls to ± 20%

Note: If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

References

1. *Healthy Water, Healthy People Testing Kit Manual*. 2002, Bozeman, Montana: The Watercourse, International Project WET.
2. *Quality Criteria for Water (Red Book)*. 1976, Washington, D.C.: U.S. Environmental Protection Agency.
3. *Water Science for Schools*. U.S. Geological Survey. http://ga.water.usgs.gov/edu/characteristics.html (accessed Sept 25, 2010).
4. American Chemical Society, *ChemCom: Chemistry in the Community*. 2002, Freeman and Company: New York. p. 77.
5. Lower, S. *Hard water and water softening*. http://www.chem1.com/CQ/hardwater.html (accessed Sept 25, 2010).
6. Wilkes University Center for Environmental Quality. *Hard Water, Hardness*. http://www.water-research.net/hardness.htm (accessed Sept 25, 2010).

Iron

What is iron?

Iron is the fourth most abundant of the elements, by weight, that make up the earth's crust. Iron is common in many rocks and soils, especially in clay soils, where it usually is a major component. The primary forms of soluble iron in aquatic environments are the ferrous, or bivalent, ion (Fe^{2+}), and the ferric, or trivalent, ion (Fe^{3+}); however, the ferric ion is not very soluble in water. Iron also is bound up in organometallic compounds and colloidal forms.

Why measure iron?

Iron is an essential trace element required by both plants and animals. It is vital to oxygen transport by hemoglobin in all vertebrates and some invertebrates. Although iron is a required dietary element, most iron is consumed through food, not through water.

In some waters, iron may be a limiting factor for the growth of algae and other plants, especially in very alkaline conditions. However, excessive iron can adversely impact aquatic life. Iron has been shown to be lethal to invertebrates that serve as fish food and to the fish themselves. For example, in the United States, iron deposits in the Brule River, in the states of Michigan and Wisconsin, were found to have long-term effects on fish food organisms, even after the pumping of iron-bearing waters from deep shaft iron mines had ceased. Other studies have demonstrated lethal toxicity to carp, pike, and trout.

In the presence of dissolved oxygen, iron in water from mine drainage is precipitated as ferric hydroxide ($Fe(OH)_3$), a yellowish precipitate, or as ferric oxide (Fe_2O_3), a reddish precipitate. Both of these precipitates form as gels or flocs that can be detrimental to fish and other aquatic life. They can settle to form flocculent materials that cover stream bottoms, destroying bottom-dwelling invertebrates, plants, or incubating fish eggs. Ferric hydroxide flocs have been observed to coat the gills of white perch, minnows, and silverside fish. With time, these flocs can consolidate to form cement-like materials, thus consolidating bottom gravels into pavement-like areas that are unsuitable as spawning sites for nest-building fish. Trout and salmon are particularly affected, because their eggs require the protection of the spaces in gravel and a constant flow of oxygen-rich water. Settling iron flocs have also been reported to trap and carry diatoms downward in waters. It is feared that these settled iron flocs may have adverse effects on important commercial mussels and other shellfish resources.

Iron is an objectionable component in water supplies for either domestic or industrial use. Iron appreciably affects the taste of beverages at levels of 1.8 mg/L in spring water and at 3.4 mg/L in distilled water, and it can stain laundered clothes and plumbing fixtures. Dissolved iron can adversely affect many industrial processes, including the production of beverages, textiles, paper, leather products, and power.

High levels of dissolved irons in sprayed-on irrigation water can result in a build-up of iron oxides precipitated on leaves, causing growth stunting. Conversely, in alkaline soils, iron may be so insoluble that iron becomes unavailable to plants, causing chlorosis, or leaf-yellowing disease.

Regarding recreational use of water, suspended iron precipitates or deposits of yellow ochre or reddish iron oxides may interfere with swimming and may be aesthetically objectionable.

High concentrations of iron may signal the presence of industrial or mining waste or runoff.[1]

What factors affect iron levels?

Prime pollution sources of iron are industrial wastes, mine drainage waters, and iron-bearing groundwater.

pH affects iron levels, since iron is more soluble in acidic conditions and is practically insoluble in alkaline conditions. Acid rain can increase dissolved iron levels in natural waters, particularly in waters with low alkalinity levels (see also the Alkalinity section19).

The level of dissolved oxygen affects iron concentrations. In the presence of dissolved oxygen, iron rapidly oxidizes to the ferric ion that then precipitates to the bottom as ferric hydroxide or ferric oxide. In the absence of dissolved oxygen, the ferrous ion can persist in solution. This occurs in deep lakes that have an anaerobic layer near the bottom. If hydrogen sulfide (H_2S) is present, ferrous sulfide (FeS), a black compound, may form and precipitate, producing black mineral muds.[1]

What are ideal iron values?

The EPA has established a criterion of 0.3 mg/L for natural waters, and a Secondary Drinking Water Standard of 1 mg/L.[2] The European Inland Fisheries Advisory Commission recommended in 1964 that iron concentrations not exceed 1.0 mg/L in waters to be managed for aquatic life.[1]

Dissolved iron (0.32 mg/L) has been shown to be toxic to mayflies, stoneflies, and caddisflies, which are all important foods for fish. Iron was found to be toxic to carp at 0.9 mg/L when the water had a pH of 5.5. Iron concentrations of 1 to 2 mg/L were shown to be lethal to pike and trout. In an iron- polluted Colorado stream, neither trout nor other fish were found until the waters were diluted or the iron had precipitated to effect a concentration of less than 1.0 mg/L, even though the other water quality constituents measured were suitable for the presence of trout.[1]

There is no EPA Primary Drinking Water Standard[3] for iron, since dissolved iron is not toxic, but rather produces an objectionable taste when the Secondary Drinking Water Standard (1 mg/L) is exceeded. As previously mentioned, iron appreciably affects the taste of beverages at levels of 1.8 mg/L in spring water and at 3.4 mg/L in distilled water.

Measuring Procedure

The measuring procedure for iron uses the Water Quality Colorimeter and the Iron ezSample Test Kit as follows:

1. **Plug the Water Quality Colorimeter into the data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for ammonia.

2. **Calibrate the Water Quality Colorimeter.** �(3.1)

Taking an iron measurement

SAFETY PRECAUTIONS

- ❖ Wear safety glasses and protective gloves.
- ❖ Review the MSDS for each of the Iron ezSample reagents, and have them available for reference.

1. **Follow the included Test Procedure instructions for the Iron ezSample Test Kit for sample preparation.**

2. **If required for your data collection system, begin by building an experiment. Otherwise, set the display to show *iron* measurements.** �(2.3)

3. **Set your data collection system to Manual Sampling.** �(2.3)

4. **Place the prepared test ampoule in the Water Quality Colorimeter and cover it with the black cap.**

   Note: Handle the ampoule by the tip, and wipe the outside glass lens clean with a non-abrasive cleaning tissue.

   Task Result: The iron reading will automatically appear in the display. There may be slight fluctuations, so wait until the reading settles around a point.

5. **Record the iron reading.** �(2.4)

   Note: The units *mg/L* and *ppm* (parts per million) are equivalent.

Disposal of the test vial

1. **The ampoules may contain very low levels of potentially toxic substances (see the MSDS), so check with your local regulatory agency regarding disposal procedures.**
2. **Carefully pour the test water down the drain. Shake the glass tip remaining in the sample cup into the solid waste container.**
3. **Thoroughly rinse the sample cup with deionized or distilled water.**

Accuracy, resolution, and practical detection limit (PDL)

The accuracy of the Iron ezSample Test Kit is ±10% at 75% of full scale range. Accuracy may be compromised if test results are outside the stated test ranges. The lower limit of the stated test range is the "practical detection limit (PDL)," defined as the lowest concentration at which less than ±30% error is routinely obtained.

For the Iron ezSample Test Kit, the range, accuracy, PDL, and resolution are shown on the table below.

> **Note:** If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

Iron ezSample Test Kit: Accuracy, resolution, and PDL (mg/L)

Range	Accuracy	PDL	Resolution
0–7	0.1	1.0	0.1

References

1. *Quality Criteria for Water (Red Book)*. Washington, D.C.: U.S. Environmental Protection Agency; 1976.
2. *National Secondary Drinking Water Regulations: Guidance for Nuisance Chemicals*. U.S. Environmental Protection Agency; 1992.
3. *List of Drinking Water Contaminants & MCLs*. National Drinking Water Standards. U.S. Environmental Protection Agency; 2006.

Nitrate

What is nitrate?

Nitrate (NO_3^-) is an ion that is formed by the complete oxidation of ammonium ions (NH_4^+) or nitrite ions (NO_2^{2-}) by microorganisms in the soil or water; the nitrite ion is an intermediate product of this process. Nitrate is one of the seven molecular forms that are important in the nitrogen cycle. These seven forms are: molecular nitrogen gas (N_2), nitrous oxide (NO_2), organic nitrogen in amino and amine groups, ammonia or the ammonium ion (NH_3 or NH_4^+), nitrite, and nitrate.[4]

Nitrogen is essential to life on earth, as it is essential for plant growth and development. Nitrogen gas makes up 78 percent of the air we breathe, although our source for useful nitrogen comes from the food we eat. Nitrogen, in the various forms listed above, passes from the air to the soil, to all living things, and then back into the air through the process called the nitrogen cycle.

The relationship between ammonia, nitrite and nitrate is shown by the following reactions:

$$\underset{\text{(ammonia)}}{NH_4^+} + \underset{\text{(water)}}{2H_2O} \rightarrow \underset{\text{(nitrite)}}{NO_2^-} + \underset{\text{(hydrogen ions)}}{8H^+}$$

$$\underset{\text{(nitrite)}}{NO_2^-} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(nitrate)}}{NO_3^-} + \underset{\text{(hydrogen ions)}}{2H^+}$$

Why measure nitrate?

While nitrate itself is normally not toxic to humans, and fish can tolerate low levels of nitrate, nitrite is quite toxic to most aquatic animals and to humans. Where there are high levels of nitrate, there are probably high levels of nitrite. Nitrate is easier to measure, so it is used as a marker that may indicate the need for additional testing for nitrite. A high nitrate concentration may also signal the presence of pathogenic bacterial contamination.

Under certain circumstances, nitrate can be reduced in the body to nitrite in the gastrointestinal tract, and this nitrite can move into the bloodstream and react with hemoglobin to produce methemoglobin. Methemoglobin does not transport oxygen, which is the role of hemoglobin. This process occurs particularly in infants under 3 months of age because of their immature digestive tracts, causing a disease called "blue baby" syndrome (methemoglobinemia), which can be lethal. Serious and occasionally fatal poisonings of infants, due to blue baby syndrome, have occurred following ingestion of untreated well water shown to contain nitrate at concentrations greater than 10 mg/L, a situation frequently found in shallow farm and rural community wells, often as the result of inadequate protection from barnyard drainage or from septic tanks.[4]

Because of this potential hazard to infants, municipalities in the United States are required by the EPA to test tap water for nitrates once each year, and once each quarter if nitrate levels are present above 50 percent of the maximum contamination level (MCL).

Long-term exposure to nitrates above the MCL may cause diuresis, increased starchy deposits, and hemorrhaging of the spleen, and it has possible links to bladder and ovarian cancer and non-Hodgkin lymphoma.

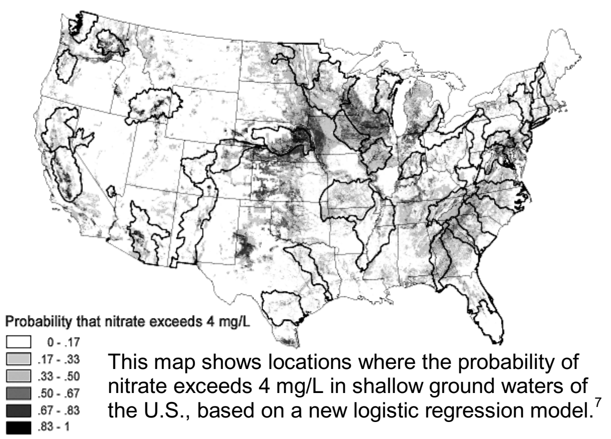
Regarding natural waters, most nitrogenous materials in natural waters tend to be converted to nitrate, so all sources of combined nitrogen, particularly organic nitrogen and ammonia, should be considered as potential nitrate sources.[1] As discussed in the Ammonia section, both ammonia and nitrite are toxic to fish and other aquatic life forms. Testing for nitrate can identify the occurrence of pollution-containing nitrates and the possible presence of toxic forms of nitrogen.

Natural waters polluted with nitrates (and phosphates), which commonly occur from runoff from farms and feed lots, undergo accelerated eutrophication, the process by which a body of water becomes enriched in dissolved nutrients. This results in algae blooms, followed by excessive bacterial respiration as bacteria break down the accumulated biomass of the algae, which consumes all the dissolved oxygen in the water, causing anoxic conditions that result in fish kills and the death of other aquatic organisms.

What factors affect nitrate values?

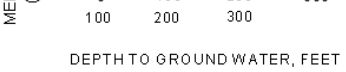
The major sources of nitrogen entry into water bodies are municipal and industrial waste waters, septic tanks, and feed lot discharges. The primary nitrogenous forms include nitrogenous fertilizer, industrial organics, industrial inorganics, explosives, and paper and pulp mills.[1] Runoff from farms following fertilization of crops, lawn fertilization, and landfills as well as atmospheric fallout and nitric oxide and nitrite discharges from automobile exhausts and other combustion processes are other sources of nitrates in natural waters.[4] The primary inorganic nitrates that may contaminate drinking water are potassium nitrate and ammonium nitrate, both of which are widely used as fertilizers. The United States Environmental Protection Agency determined that about 112 million pounds of nitrate were released into the waterways from 1991 to 1993.[1]

Nitrates have a high potential to migrate to ground water, since they are very soluble and do not bind to soils. Factors that can contribute to high levels of nitrates in ground water include high nitrogen input, well-drained soils, and less extensive woodland relative to cropland. Because nitrates do not evaporate, nitrates are likely to remain in the water until consumed by plants or other organisms. People who draw their water from shallow wells in high risk areas (see map) are especially vulnerable.[7]



This map shows locations where the probability of nitrate exceeds 4 mg/L in shallow ground waters of the U.S., based on a new logistic regression model.[7]

More than 43 million people in the United States are served by self-supplied domestic water systems, and almost all (98 percent) of this water comes from ground water. Ground water also is the source of water for about 90 million people who are served by public water-supply systems. Nitrate, from both natural sources and human activities, is possibly the most prevalent contaminant in ground water and can persist in shallow ground water for years under well-oxygenated conditions. About 13 percent of shallow wells sampled beneath agricultural and urban land-use study areas as part of the National Water-Quality Assessment (NAWQA) Program during 1992–99 exceeded the U.S. Environmental Protection Agency's drinking-water standard for nitrate (10 milligrams per liter).[7] The graph shows the increasing risk of nitrate concentrations exceeding EPA drinking water standards as the depth of the well decreases.[8]



Shallow ground water in relatively undeveloped areas of the United States contains about 1 milligram per liter (mg/L) of nitrate.[8]

Excess nitrate in ground water is a worldwide problem because of widespread use of nitrogenous fertilizers. For example, one study found that ground water in an area in China had some areas in which the nitrate concentration in ground water exceeded 20 mg/L.[6] In another example, because of this problem, England and Wales in 1996 adopted regulations that require monitoring vulnerable areas in those countries for nitrate levels, with provisions for action programs in the event that excessive nitrate concentrations are detected.[3]

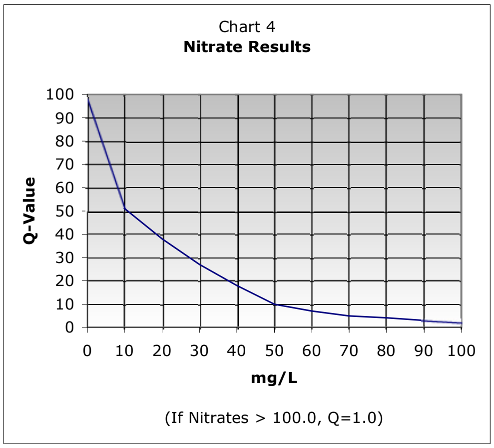
Natural attenuation of nitrate, the reduction in nitrate concentration by natural processes, may occur by plant uptake, mixing, and dilution with low-nitrate water, bacterial activity, or consumption in chemical oxidation–reduction reactions (also known as "redox reactions"). An example of the latter case is denitrification, in which the oxygen atoms of the nitrate (NO_3^-) are removed by chemical reduction to produce nitrogen gas (N_2), which is effectively nonreactive. Denitrification occurs in the absence of dissolved oxygen and in the presence of chemically reduced compounds such as organic carbon or iron sulfide minerals such as pyrite (FeS_2). Denitrification is usually mediated by bacteria, which derive energy from the reaction. Marine and estuarine silts and clays commonly contain reduced compounds that can act as a substrate for denitrification. In contrast, marine sands and fluvial (riverine) sands and gravels typically contain little organic carbon and are composed primarily of minerals such as quartz that are generally unreactive.[5]

Technologies that water suppliers use to reduce nitrate concentrations include ion exchange, reverse osmosis, electrodialysis, and distillation.[1,4]

What are ideal nitrate values?

For drinking water, the lower the nitrate levels, the better. The EPA has established a maximum contamination level (MCL) of 10 mg/L of nitrate.[2] In other countries, the standards for drinking water are higher; for example, in China the standard is 20 mg/L.[6] The EPA-established MCL for nitrite is even lower, at 1 mg/L.[2]

In natural waters, low levels of nitrate (between 1 and 3 mg/L) provide adequate nutrients for aquatic life while avoiding eutrophication. The EPS has not declared an MCL for nitrate in natural waters, since fish are not harmed by high levels of nitrate per se; however, high levels can contribute to eutrophication and the resulting depletion of dissolved oxygen, which can be fatal for fish and other aquatic organisms. Additionally, as mentioned previously, high levels of nitrate may indicate the presence of toxic levels of nitrite, that is, greater than about 5 mg/L for warm water fish and greater than about 0.1 mg/L for salmonid fish.[4] As the Q-value graph indicates, water quality drops off dramatically with increases in nitrates.



(If Nitrates > 100.0, Q=1.0)

Further information can be found in the referenced material listed below and in the Recommended Reading and Resources section.

Measuring Procedure

The measuring procedure for nitrate uses the Water Quality Colorimeter and the Nitrate ezSample Test Kit as follows:

1. **Plug the Water Quality Colorimeter into the data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for ammonia.

2. **Calibrate the Water Quality Colorimeter.** �(3.1)

Taking a nitrate measurement

SAFETY PRECAUTIONS

- ❖ Wear safety glasses and protective gloves.

- ❖ Review the MSDS for each of the Nitrate ezSample reagents, and have them available for reference.

1. **Follow the included Test Procedure instructions for the Nitrate ezSample Test Kit for sample preparation.**
2. **If required for your data collection system, begin by building an experiment. Otherwise, set the display to show Nitrate measurements.** �(2.2)
3. **Set your data collection system to Manual Sampling.** �(2.3)
4. **Place the prepared test ampoule in the Water Quality Colorimeter and cover it with the black cap.**

   Note: Handle the ampoule by the tip, and wipe the outside glass lens clean with a non-abrasive cleaning tissue.

   Task Result: The nitrate reading will automatically appear in the display. There may be slight fluctuations, so wait until the reading settles around a point.
5. **Record the nitrate reading.** �(2.4)

   Note: The units *mg/L* and *ppm* (parts per million) are equivalent.

Disposal of the test vial

1. **The ampoules may contain very low levels of potentially toxic substances (see the MSDS), so check with your local regulatory agency regarding disposal procedures.**
2. **Carefully pour the test water down the drain. Shake the glass tip remaining in the sample cup into the solid waste container.**
3. **Thoroughly rinse the sample cup with deionized or distilled water.**

Accuracy, resolution, and practical detection limit (PDL)

Note: The Nitrate ezSample Test Kit measures nitrate nitrogen. To convert the test results to nitrogen, multiply by 4.4 as shown on the table below.

The accuracy of the Nitrate ezSample Test Kit is ±10% at 75% of full scale range. Accuracy may be compromised if test results are outside the stated test ranges. The lower limit of the stated test range is the "practical detection limit (PDL)," defined as the lowest concentration at which less than ±30% error is routinely obtained.

For the Nitrate ezSample Test Kit, the range, accuracy, PDL, and resolution are shown on the table below.

Note: If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

Nitrate ezSample Test Kit: Accuracy, resolution, and PDL (mg/L)*

Range (nitrate)	Range (nitrate nitrogen)	Accuracy	PDL	Resolution
0–4.4	0–1 (Low)	0.05	0.2	0.01
4.4–11	1–2.5 (High)	0.1		0.1

*To convert nitrate-N to nitrate, multiply by 4.4. For example, a test result of 1 mg/L nitrate-nitrogen (nitrate-N) converts to 4.4 mg/L nitrate.

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pH

What is pH?

The pH test measures the concentration of hydrogen ions in water, assessing how acidic or basic the solution is. The pH of natural waters is a measure of the acid–base equilibrium achieved by the various dissolved compounds, salts, and gases.[2] The pH value is the negative log of the hydrogen ion concentration:

$$\text{pH} = -\log\,[\text{H}^+]$$

pH values range from 0 to 14, with 0 being the most acidic, 7 being neutral, and 14 being the most basic.

One pH unit equals a 10-times difference in acidity

Since pH is defined as the negative log of the hydrogen ion concentration [or rather more precisely, the concentration of hydronium ions (H_3O^+), which is the form taken by the hydrogen ion in aqueous solutions], lower pH values indicate higher hydrogen ion concentrations (higher acidity), and each decrement in a pH unit represents a 10-fold increase in acidity. Thus, a substance that has a pH of 6 is 10 times more acidic than a substance with a pH of 7; a substance with a pH of 5 is 100 times (10 × 10) more acidic than a substance with a pH of 7.

This relationship of a 10-fold change in acidity for every unit of pH is can be intuited from a consideration of pH values of common substances. Referring to the table, "pH of common solutions," you can relate the pH value to your own experience with some of these substances. Substances at the ends of the pH range (highly acidic and highly basic) are highly reactive substances that can cause damage to biological tissues and even to inorganic substances, such as metal pipes. While pH is an important factor in the reactivity of materials, it is not the only factor. Nevertheless, even very small changes in pH, and certainly rapid changes in pH, can upset the balance in a water system.

pH values of common solutions[1]

Solution	pH Value
Hydrochloric acid (concentrated)	0.0
Stomach acid	1.0 – 3.0
Lemon juice	2.2 – 2.4
Vinegar	2.4 – 3.4
Cola	2.6
Grapefruit juice	3.0 – 3.2
Acid rain	4.0 – 5.5
Natural rain	5.6 – 6.2
Milk	6.3 – 6.7
Pure deionized water	7.0
Sea water	7.0 – 8.3
Baking soda	8.4
Milk of magnesia	10.5
Household ammonia	11.9
Sodium hydroxide (concentrated)	13.0 – 14.0

Why measure pH?

pH is an important component of water quality, affecting the solubility of metals and other substances including nutrients. Extremely low and high pH values are incompatible with life for a variety of reasons that include potentiating harmful levels of certain metals, such as cadmium and mercury, and promoting corrosion of living tissues. pH is a component of the Water Quality Index and the EPA has adopted a secondary standard for pH regarding drinking water.[1]

What factors affect pH?

The carbonate system, which is composed of carbon dioxide (CO_2), carbonic acid (H_2CO_3), bicarbonate ion (HCO_3^{-}), and carbonate ion (CO_3^{-2}), is the principle system that regulates pH in natural waters.[2] Therefore, photosynthesis and respiration can affect the pH of water.

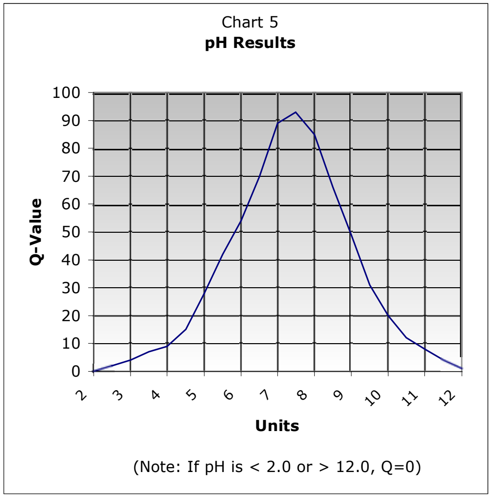
Chart 5
pH Results

(Note: If pH is < 2.0 or > 12.0, Q=0)

Acidifying influences, those that lower the pH, include respiration by plants, animals, and bacteria, acid rain resulting from the burning of fossil fuels, and runoff from mines, chemical plants, and other non-natural sources.

In any case, as the pH changes from neutral to either more acidic or more alkaline, the water quality, and corresponding Q-value, deteriorates. The table of Q-values for pH (at left) demonstrates this relationship.

Temperature alone does not appreciably affect pH levels, but it does affect the functioning of the sensor electrode. For best results, measure the pH at a temperature close to that of the calibration solutions.

Buffers affect how dramatically the pH changes following the addition of acidic or basic substances

The extent to which pH values will change in response to added acids or bases is dependent upon the amount of buffering materials dissolved in the water. Water that contains higher levels of dissolved carbonates, derived from erosion of limestone and other carbonate rocks, has a higher buffering capacity and is thus more resistant to changes in pH. This higher buffering capacity is reflected in higher conductivity, total dissolved solids, and alkalinity values.

Because of the buffering system present in sea water, which contains salts that contribute to its alkalinity, the naturally occurring variability of pH is generally less than that for fresh water.[2]

What are ideal pH values?

The range of pH values considered desirable depends upon the designated use of the water. Some examples are listed in the table, "Desirable pH values, by designated use." Biological systems are generally more sensitive to damaging effects of changes in pH than non-biological systems.

Desirable pH values, by designated use

Designated Use	pH Value	Reference
Drinking water	6.5 – 8.5	EPA Secondary Drinking Water Standard[1]
Freshwater aquatic life	6.5 – 9.0	EPA[2]
Marine aquatic life	6.5 – 8.5	EPA[2]
Domestic water supplies (welfare)	5 – 9	EPA[2]
Crop irrigation	any*	EPA[2]
Industrial cooling	5.0 – 8.9	EPA[2]
Industrial processing	3.0 – 11.7	EPA[2]

*The pH of irrigation water rapidly changes to approximately that of the soil to which it is applied.

Different organisms thrive at different pH ranges. Some examples are listed in the table, "pH values compatible with life, by organism." A pH value that is 5 or less or that is 9 or more is generally harmful to most aquatic organisms; however, some bacteria thrive at pH values lower or higher than that. Certain organisms, such as trout and many larvae, are more sensitive to changes in pH than other aquatic organisms.

pH values compatible with life, by organism[3]

Organism	pH Value
Bacteria	2 – 13
Plants (algae, rooted, etc.)	6.5 – 12
Carp, suckers, catfish, some insects	6 – 10
Bass, bluegill, crappie	6.5 – 9.5
Snails, mussels, clams	7 – 10
Trout, mayfly nymphs, stonefly nymphs, caddisfly larvae	6.5 – 8.5

Normal pH values of sea water are 8.0 to 8.2 at the surface, decreasing to 7.7 to 7.8 with increasing depth. Some marine communities are more sensitive to pH changes than others, with plankton and bottom-dwelling invertebrates being generally more sensitive than fish. Marine communities in shallow, biologically active waters in tropical or subtropical areas can tolerate or adapt to the large diurnal changes in pH (ranges from pH 7.3 to pH 9.5 in a day) that occur because of photosynthetic activity.

However, mature and larval forms of oysters may be adversely affected by values as low as pH 6.5 or as high as pH 9.5.[2]

Unnatural, rapid changes in pH can be harmful

Even if the pH is in the desirable range for a given use, a rapid change that is not part of the naturally occurring variation can be harmful to organisms. Small changes in pH can cause large shifts in metallic complexes, such as those of copper, zinc, cadmium, aluminum, and lead, which may affect the toxicity of these metals. For example, the pH should not be changed more than 0.2 pH units outside of the naturally occurring variation in deep ocean waters.[2]

Therefore, when monitoring pH, the change in pH over time or throughout an area is an important consideration.

Measuring Procedure

Setting up the pH Sensor

1. **Plug the pH Sensor into the data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for dissolved oxygen.

   Note: The Water Quality Sensor can also be used.

2. **Remove the storage bottle from the end of the probe.**

   Note: You can simply unscrew the soaker bottle and remove the bottom part, leaving the lid attached to the probe.

3. **Calibrate the pH Sensor �(2.7) if required, or use the factory calibration for non-standard comparisons.**

Taking a pH measurement

1. **If required for your data collection system, begin by building an experiment.** �$^{(2.2)}$
2. **Set your data collection system to Manual Sampling.** �$^{(2.3)}$
3. **Insert the pH Sensor into the solution.**
4. **Record the pH reading.** �$^{(2.4)}$

More background information

More information about pH and its role in water quality may be found in the references listed below, and in the Recommended Reading and Resources section.

See also the Resources section for vendors of pH buffer solution standards.

References

1. *National Secondary Drinking Water Regulations. Current Drinking Water Standards.* U.S. Environmental Protection Agency; 2006.
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Phosphate

What is phosphate?

Phosphate (PO_4^{3-}) is an ionic (water-soluble) form of phosphorous. Phosphorus can be measured as total phosphorus (TP) or as soluble reactive phosphate (SRP), which is also called phosphate (PO_4^{3-}) or orthophosphate (ortho-P). Phosphate represents the fraction of TP that is available to organisms for growth and is the fraction that is usually measured in water quality studies. Phosphate, PO_4^{3-}, is measured in terms of mg/L (parts per million, ppm) or $\mu g/L$ (parts per billion, ppb). Phosphate phosphorous, PO_4^{3-}P, is equivalent to a phosphate measurement divided by 0.3261, which is the ratio of the atomic weights of phosphorus to phosphate (30.97 g/mol / 94.97 g/mol).

Phosphate is one of the major nutrients required for plant and animal growth and is essential for life. Phosphate is used by all aerobic organisms (organisms that require oxygen to live) in the production of energy-storage molecules, and it is a constituent of cell membranes and other intracellular molecules and structures, including DNA.

Why measure phosphate?

Phosphate is generally the rate-limiting eutrophication reagent. That is, in most circumstances, of all the nutrients required for freshwater plant growth, phosphate is the one that is used up first. Therefore, when there is no more phosphate, plant growth stops. Conversely, when phosphate is present, plant growth is stimulated. The presence of excessive amounts of phosphate stimulates excessive plant growth.

Excessive plant growth in aquatic environments stimulated by excessive nutrient content, called accelerated or cultural eutrophication, interferes with water use and is considered a nuisance. Algal growths impart undesirable tastes and odors to the water, interfere with water treatment, become aesthetically unpleasant, alter the chemistry of the water supply, and contribute to the phenomenon of accelerated or cultural eutrophication. Eutrophication often leads to anoxic conditions that occur when plant material is broken down by bacteria, and the dissolved oxygen is consumed during bacterial respiration. Most of the eutrophication problems are associated with lakes or reservoirs, rather than streams and rivers.[1]

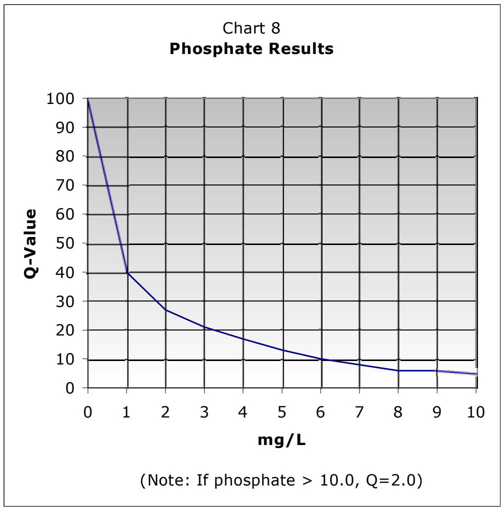
Chart 8
Phosphate Results

(Note: If phosphate > 10.0, Q=2.0)

The impact of phosphate levels on water quality is illustrated by the phosphate Q-value curve for determining a water quality index as shown in this graph (see also the Water Quality Index section).

The Q-value curve for phosphate that is used in the calculation of a water quality index illustrates that water quality declines dramatically as the concentration of phosphate increases.

Of all the elements required for aquatic plant growth, phosphorus is the most easily controlled, and measurement is the first step in that control.[1] The presence of high levels of phosphates may indicate a point source of pollution that could be better managed.

What factors affect phosphate values?

Phosphate comes originally from phosphate rock, or phosphorite, a non-detrital sedimentary rock. The phosphate rock beds lie within a few feet of the surface, and mining is accomplished by use of hydraulic water jets and a washing operation that separates the phosphate from waste materials, a process similar to strip mining. Runoff from such operations can cause elevations in aquatic phosphate levels.

Phosphates are present in human and animal waste including bird guano, so untreated sewage causes increases in phosphate levels. Phosphate detergents used to be significant sources of aquatic phosphate, but environmental concerns have prompted the development of nonphosphate-containing laundry detergents. Nevertheless, some phosphate cleaners and other domestic chemical preparations containing phosphate are still in use and can enter natural waters through urban runoff. Some industries, such as potato processing, have wastewaters high in phosphates. Crop, forest, urban, and idle land contribute varying amounts of phosphates from surface runoff of rainfall, effluent from drainage pipes, or return flow from irrigation. Fertilizers, cattle feedlots, concentrations of domestic duck or wild duck populations, tree leaves, and fallout from the atmosphere are all contributing sources.[1]

Once phosphates are consumed by algae and other aquatic plants, they are removed from the water. When the plants decompose, some of the phosphates are released into the water again through bacterial catabolism. However, much of the phosphates in decomposing plant matter settles to the bottom, becomes consolidated with other material on the bottom, and is bound permanently so it will not recycle into the system, with the net effect of a reduction in phosphate levels in the water.[1]

What are ideal phosphate values?

Most relatively uncontaminated lakes are known to have surface waters that contain from 0.03 to 0.09 mg/L phosphate. However, in lakes or reservoirs, phosphate levels as low as 0.08 mg/L may stimulate excessive or nuisance growths of algae and other aquatic plants during the spring when nutrients are cycling to the surface. Streams or other flowing water are somewhat less susceptible to accelerated or cultural eutrophication, so a desired goal for them is a concentration of phosphate of less than 0.3 mg/L. In areas where streams enter lakes or reservoirs, the desired phosphate level is less than 0.15 mg/L.[1] However, it should be kept in mind that phosphates in streams and rivers may end up in a lake or reservoir downstream.

Phosphate is not regulated by the U.S. EPA because it is not toxic to humans or animals, and it is not considered a nuisance chemical in drinking water.

Further information is available from the references below or from the Recommended Reading and Resources section.

Measuring Procedure

The measuring procedure for phosphate uses the Water Quality Colorimeter and the Phosphate ezSample Test Kit as follows:

1. **Plug the Water Quality Colorimeter into the data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for ammonia.

2. **Calibrate the Water Quality Colorimeter.** �(3.1)

Taking measurements

SAFETY PRECAUTIONS

- Wear safety glasses and protective gloves.
- Review the MSDS for each of the Phosphate ezSample reagents, and have them available for reference.

1. **Follow the included Test Procedure instructions for the Phosphate ezSample Test Kit for sample preparation.**
2. **If required for your data collection system, begin by building an experiment. Otherwise, set the display to show Phosphate measurements.** �(2.2)
3. **Set your data collection system to Manual Sampling.** �(2.3)
4. **Place the prepared test ampoule in the Water Quality Colorimeter and cover it with the black cap.**

   Note: Handle the ampoule by the tip, and wipe the outside glass lens clean with a non-abrasive cleaning tissue.

   Task Result: The phosphate reading will automatically appear in the display. There may be slight fluctuations, so wait until the reading settles around a point.

5. **Record the phosphate reading.** �(2.4)

   Note: The units *mg/L* and *ppm* (parts per million) are equivalent.

Disposal of the test vial

1. **The ampoules may contain very low levels of potentially toxic substances (see the MSDS), so check with your local regulatory agency regarding disposal procedures.**
2. **Carefully pour the test water down the drain. Shake the glass tip remaining in the sample cup into the solid waste container.**
3. **Thoroughly rinse the sample cup with deionized or distilled water.**

Accuracy, resolution, and practical detection limit (PDL)

The accuracy of the Phosphate ezSample Test Kit is ±10% at 75% of full scale range. Accuracy may be compromised if test results are outside the stated test ranges. The lower limit of the stated test range is the "practical detection limit (PDL)," defined as the lowest concentration at which less than ±30% error is routinely obtained. For the Phosphate ezSample Test Kit, the range, accuracy, PDL, and resolution are shown on the table below.

Phosphate ezSample Test Kit: Accuracy, resolution, and PDL (mg/L)

Range	Accuracy	PDL	Resolution
0–1 (low)	0.05	0.05	0.01
1–8 (high)	0.1	1	1

Note: If your test result is at the maximum of the test range, dilute the sample with distilled water and re-test, then make the correction for the dilution. For example, dilute 10 mL of sample with 10 mL of distilled water, and then multiply the test result by 2 for the final concentration.

References

1. *Quality Criteria for Water (Red Book)*. Washington, D.C.: U.S. Environmental Protection Agency; 1976.
2. National Science Foundation. *Water Quality Index*. 2004.

Salinity

What is salinity?

Salinity is a measure of the concentration of dissolved solids in a solution of water. There are several ways to express salinity, including percent (parts per hundred), parts per thousand (ppt), milligrams per liter (mg/L), decisiemen per meter (dS/m), or grams per kilogram (g/kg). The most common way to express salinity is in parts per thousand, which is approximately equal to grams of salt per liter of solution, and the table below shows the equivalent values for average seawater salinity.

Comparison of Units of Salinity for Seawater

3.5%		(parts per hundred)
35	ppt	(parts per thousand)
35	g/kg	(grams per kilogram of water)
35,000	ppm	(parts per million)
35,000	mg/L	(mg per liter)
54.69	dS/m	(decisiemens per meter)

A salinity measurement is markedly different from that of total dissolved solids (TDS) because it is specific to ionic materials found in water. Where TDS can include micromolecules of organic compounds as well as salts, the measure of salinity only includes those ionic compounds that can conduct a current. The most common salt in seawater is sodium chloride (NaCl), but magnesium chloride ($MgCl_2$), sodium sulfate (Na_2SO_4), calcium chloride ($CaCl_2$), potassium chloride (KCl), sodium biocarbonate ($NaHCO_3$), potassium bromide (KBr), hydrogen borate (H_3BO_3), strontium chloride ($SrCl_2$), and sodium fluoride (NaF) are also found.[1]

Water is considered saline if it contains moderate or relatively high amounts of dissolved salts. The term is most often employed to describe water that would, if evaporated fully, leave behind salts incorporating sodium, calcium, or magnesium. Salinity is the degree to which a water is saline.[6] In the 1970s, salinity was redefined as the conductivity ratio of a water sample to a standard potassium chloride (KCl) solution.

Why measure salinity?

Salinity measurement is one of the most important factors in determining the types of plants or animals that can live in or near a body of water. The affects of salinity are felt strongly in the agricultural industry, where the proper landscape irrigation is dependent upon water quality. If the salinity content of irrigation water is too high, salts build up in the soil and starve root systems of water. Plants affected by highly saline irrigation waters will exhibit dwarfing and yellowing of leaves or may lose leaves altogether.

Highly saline soils cause an osmotic effect that adversely affects less salt-tolerant plants. Water naturally moves from less salty areas to areas containing higher concentrations of salt (osmosis). The more salt that accumulates in soil, the more energy a plant must exert to pull water through the root system and into the plant. Once the salt concentration becomes so high that the plant is unable to retrieve water from the soil, injury or death will occur.

An accumulation of salts in soil can additionally cause changes in the physical properties of the soil. Irrigation waters that are high in sodium can cause clay to break down in a process known as deflocculation. As a result of deflocculation, soil becomes more compact, reducing the soil's water-carrying capacity by decreasing its aeration.[2]

Water Type	Salinity Value (ppt)
Fresh water	<0.5
Brackish water	0.5 to 30
Saline water	30 to 50
Seawater average	35
Brine	>50

The salt content of bodies of water likewise dictates the organisms that live in that body of water. Many aquatic plants and animals can tolerate a wide range of salinities but often prefer a specific salinity concentration in which to reproduce.

Fish, like plants, depend upon osmosis to regulate their internal fluids. This process is known as osmoregulation. Fish in saline waters must prevent loss of excess water due to osmosis by drinking sea water and excreting or secreting the excess salts. Freshwater fish, on the other hand, must work to maintain the salt concentration of their internal fluids and must continuously excrete excess water.[3]

The Salinity Sensor measures salinity, conductivity, and temperature. The Salinity Sensor's range for conductivity is higher than the range in the Conductivity Sensor, which makes the Salinity Sensor more useful for testing brackish or estuarine water. The sensor determines salinity based on electrical conductivity. The sensor carries out a built-in calculation to compensate for the change in conductivity due to temperature change based on the Practical Salinity Scale (PSS).[6]

What factors affect salinity values?

Like conductivity, salinity is determined by the concentration of ions in solution and the mobility of those ions. Warmer water has a higher ability to dissolve more salts than cooler water, thus, saltier water can be found often in tropical regions. However, there are a number of other factors that contribute to the salinity concentration of ocean waters. Evaporation in these areas is high due to the amount of solar energy emitted over these waters. The fresh water in these tropical oceans evaporates, while the salts are left behind, causing an increase in salinity.

Factors That Cause an Increase in Salinity	Factors That Cause a Decrease in Salinity
Evaporation of fresh water concentrates salts left behind	Large river systems emptying into the ocean
Runoff from roadbeds that have been salted to remove ice	Periodic releases of reservoir water
Warmer water temperatures allow more salt to be carried in solution	Cooler water temperatures allow less salt to be held in solution
Accumulation of sea ice leaves seawater saltier	Melting of sea ice reduces salinity concentration
Rainfall if the soil runoff leaches salts	Rainfall if the soil runoff does not leach salts

Why then are some of the most saline waters found off the coast of northern Canada? Multiple areas off the coast of northern Canada have salinity concentrations of up to 41.5 ppt. Like evaporation, the formation of sea ice increases salinity. Sea ice forms as a result of freezing freshwater, leaving the salts behind. In far northern Canada and the Greenland coasts, where ice packs are thick and abundant, this process of sea ice formation serves to raise the salinity of the surrounding ocean. Conversely, the melting of sea ice will provide an influx of freshwater that will cause a decrease in salinity concentration. A number of other natural processes, including precipitation, runoff from land, and icebergs melting, all contribute to a decrease in salinity concentration as well.[4]

While solubility as a result of temperature is a factor when using the PASPORT Salinity Sensor, it is not as important as temperature's effect on the mobility of the ions. When measuring salinity, the temperature of the sample must be referenced to 25 °C, the standard reference temperature for salinity, conductivity, and total dissolved solids measurements. As the temperature of a sample increases, the mobility of the ions increases. This mobility is measured as an increase in conductivity, which, if not corrected for temperature, can be reported as a false high value of salinity. The Salinity Sensor compensates for temperature differences and corrects both the salinity and conductivity to readings consistent with a 25 °C reading.

What are desirable levels of salinity?

Desirable levels of salinity are dependent upon the species of plant or animal and the intended use of the water in question and are often expressed in terms of conductivity. Drinking water, for example, must be low in salinity, as highly saline water can cause dehydration in both humans and animals. Water given to livestock and poultry must be below the 5.0 to 8.0 dS/m range. Water beyond this range will cause diarrhea, decreased growth, birth defects, and eventually death.[5] Animals, especially livestock, have the ability to adapt to gradually changing levels of salinity; however, rapid changes in salinity will cause sickness or death.[7]

Freshwater and marine aquatic organisms live and breed in a wide range of salinity levels. Brine shrimp, for example, prefer 30 to 50 ppt, while goldfish prefer 0 to 5 ppt.[6] These ranges are, of course, species specific, and many freshwater fish may be capable of tolerating much higher levels of salt.

Water for use in irrigation should generally be less than 7.5 dS/m. Some plants and crops, however, are quite tolerant of salts. Certain ion concentrations, though, must be watched. Boron, for example, only needs to be slightly higher than the safe concentration to be very harmful to plants.[1]

How are units of conductivity converted to units of salinity?

To convert dS/m to ppm (parts per million), multiply the conductivity value in dS/m by 640. To convert from ppm to dS/m, divide by 640. Recall that 1 ppm is 1,000 times less than 1 ppt.[8]

Measuring Procedure

Setting up the Salinity Sensor

1. **Plug the Salinity Sensor into the data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for salinity.

2. **Calibrate the Salinity Sensor.** �(2.9)

Taking a salinity measurement

1. **If required for your data collection system, begin by building an experiment.** �(2.3) **Otherwise, set the display to show Salinity measurements.** �(2.3)

2. **Set your data collection system to Manual Sampling.** �(2.3)

3. **Immerse the end of the Salinity probe into the water to be measured. When the reading stabilizes, record the salinity measurement.** �(2.4)

   Note: Be sure to immerse the holes near the end of the sensor. Swirl the probe to be sure any air bubbles are removed from the end of the probe.

4. **Take additional measurements as needed, following the procedure above.**

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Stream Flow

What is stream flow?

Stream flow, or discharge, is the volume of water that moves over a designated point over a fixed period of time. It is often expressed as cubic feet per second (ft^3/s) or cubic meters per second (m^3/s). Stream flow is affected seasonally by increasing or decreasing rainfall, and by snow melt. Stream flow can also be affected by the introduction of debris into the stream, beaver dams, or vegetation growth or removal. When the velocity of the stream changes, organisms that are adapted to that stream are affected. Additionally, stream flow has an effect on other water quality parameters such as dissolved oxygen and temperature.

Calculating stream flow involves solving the following equation:

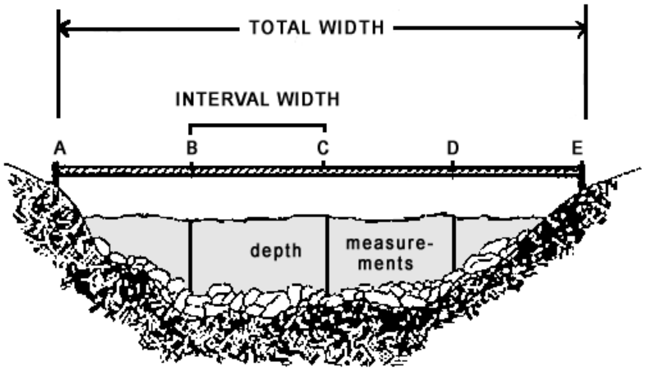
$$\text{Flow} = A \times V$$

where

A is the average cross-sectional area of the stream (stream width multiplied by average water depth; ft^2 or m^2) (see figure[1])

V is the velocity (ft/s or m/s)

Measuring the cross-sectional area



To measure the cross-sectional area[1]:

1. **Determine the average depth along the transect by marking off with twist ties equal intervals along a string stretched across the stream.**
2. **Measure the depth at each interval.**
3. **Calculate the average depth by adding together the depth measurements and dividing by the number of intervals.**

   Note: The number of intervals will be the number of depth measurements plus 1. You add 1 to the denominator to account for the 0 depth at the shores. In this example, you would divide the sum of the 3 measurements by 4.

4. **Calculate the cross-sectional area by multiplying the total width times the average depth.**

Why measure stream flow?

Stream flow is important because of its impact on water quality and on the living organisms and habitats in the stream. Large, swiftly flowing rivers can receive pollution discharges and be little affected, whereas small streams have less capacity to dilute and degrade wastes.

Stream velocity, which increases as the volume of the water in the stream increases, determines the kinds of organisms that can live in the stream (some need fast-flowing areas; others need quiet pools). It also affects the amount of silt and sediment carried by the stream. Sediment introduced to quiet, slow-flowing streams will settle quickly to the stream bottom. Fast-moving streams will keep sediment suspended longer in the water column. Fast-moving streams generally have higher levels of dissolved oxygen than slow streams because they are better aerated.[1]

Monitoring stream flow provides information regarding the available resources for aquatic organisms, especially fish, as well as available resources for human use. The Endangered Species Act, salmon recovery efforts, and increased focus on water resource management have fueled the need for accurate and timely stream flow data. Adjusting water quality measurements for stream flow can enhance the ability to detect long-term changes in water quality.[2]

The U. S. Geological Survey (USGS) continuously monitors stream flow across the United States, and real-time data is available on the USGS Web site.[3] Some stream-flow monitoring programs integrate the locally collected stream flow data with that collected by the USGS.[2]

What factors affect stream flow values?

The flow of a stream is directly related to the amount of water moving off the watershed into the stream channel. It is affected by weather, increasing during rainstorms and decreasing during dry periods. It also changes during different seasons of the year, decreasing during the summer months when evaporation rates are high and shoreline vegetation is actively growing and removing water from the ground.

Water withdrawal for irrigation purposes can seriously deplete water flow, as can industrial water withdrawals. Dams used for electric power generation, particularly facilities designed to produce power during periods of peak need, often block the flow of a stream and later release it in a surge.[1]

What are ideal stream flow values?

There are no ideal stream flow values. Rather, stream flow data should be considered in the context of other water quality measurements. However, changes in flow affect the water quality, aquatic organisms, and availability of water for human use, as well as the potential for flooding. Additionally, some organisms such as salmon require a certain stream flow to complete their reproductive cycles.

Further information is available from the references; also refer to the Recommended Reading and Resources section.

Measuring Procedure

Setting up the Flow Rate/Temperature Sensor

1. **Plug the Flow Rate/Temperature Sensor into your data collection system.** �(2.1)

   Task Result: On some sensors, a digits display of flow rate values automatically opens.

2. **The Flow Rate/Temperature sensor does not require calibration.**

3. **Extend the telescoping tube of the probe as needed. The tube will extend to 7 feet (2.13 meters).**

Taking a stream flow measurement

SAFETY PRECAUTIONS

- ❖ If standing on a bank or the shore, take care to stand on solid, non-slippery ground. The extended pole and the strong flow of the river or stream can set you off balance.
- ❖ If standing on a bridge or in a boat, do not lean over the edge. Wear life jackets when necessary.
- ❖ Do not let the data collection device fall into the water. Use a lanyard or tether to secure it.

1. **Take the first measurement.**
   a. Immerse the end of the Flow Rate/Temperature probe into the water just below the surface. Point the arrow on the impeller in the direction of stream flow.

      Task Result: The flow rate and temperature values appear on the display.

   b. Observe the display. Record the velocity and temperature of the water when the value stabilizes.
   c. If the reading drops to zero, remove the probe and clean debris from the impeller.
2. **Take additional measurements as needed, following the procedure above.**

References

1. U. S. Environmental Protection Agency. *Water: Our Waters: Rivers & Streams: Monitoring & Assessment: Stream Flow*. http://water.epa.gov/type/rsl/monitoring/vms51.cfm (accessed Sept 29, 2010).
2. Department of Ecology, State of Washington. *River and Stream Flow Monitoring*. http://www.ecy.wa.gov/programs/eap/flow/shu_main.html (accessed Sept 29, 2010).
3. U. S. Geological Survey. *WaterWatch—Current Water Resources Conditions*. http://waterwatch.usgs.gov/ (accessed Sept 29, 2010).

Temperature

What is temperature?

Temperature is the physical property of a system that underlies the common notions of "hot" and "cold"; something that is hotter has greater temperature. Temperature arises from the random microscopic motions of the atomic and subatomic constituents of matter because of their kinetic energy. Temperature is related to the average energy of these microscopic motions.

Temperature is measured with thermometers that may be calibrated to a variety of temperature scales. The scientific community uses Celsius and the thermodynamic (absolute) temperature scale (kelvin). In common usage in some countries is the Fahrenheit scale.

Why measure temperature?

The temperature of water affects the way people use water and the abilities of aquatic organisms to live, grow, and reproduce. Temperature is a catalyst, a depressant, an activator, a restrictor, a stimulator, a controller, and a killer, and it is one of the most important and influential water quality characteristics for life in water.[2, 3]

Since nearly all aquatic organisms are “cold-blooded,” the temperature of the water regulates their metabolism and ability to survive and reproduce effectively. Temperature affects the ability of water to hold dissolved oxygen (see the section on dissolved oxygen for more information). The temperature of the water regulates the species composition and activity of aquatic life in any location. For example, one study showed that diatoms predominated in temperatures from 20 °C to 25 °C, green algae predominated at 25 °C to 30 °C, and blue-green algae predominated at 30 °C to 35 °C. A few degrees elevation in average monthly temperature can appreciably alter the species composition of an aquatic community through changes in inter-species relationships.[2]

Each species has a favored range of water temperatures for its various life processes (for examples, see the table below[2]). Although aquatic organisms can become acclimated to slow, natural changes in water temperature, rapid unnatural changes, such as those caused by industrial thermal pollution or large cool water releases from dams, can be lethal for them, even in the temperature range that organisms can tolerate under natural conditions. Thus for aquatic life, the rate and amount of change in temperature is as important as the present temperature of the water.

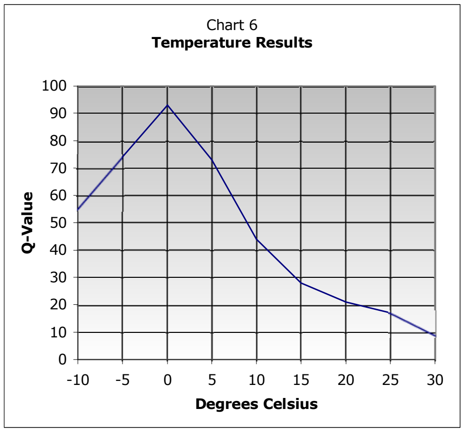
What factors affect temperature?

Natural influences on water temperature include the ambient air temperature, evaporation, solar radiation, water depth, flow rate, and the amount of mixing in the water column. Examples of man-made influences include the practice of clear-cutting to the water’s edge, release of heated water from industrial applications, and release of cold-impounded water into streams. One study showed clear-cutting the forest around a stream resulted in a 7.8 °C change in temperature (from 13.9 °C to 21.7 °C), but leaving a 100-foot buffer region of uncut vegetation along the stream’s edge prevented the temperature change.[1]

What are ideal temperature values?

Desirable temperatures depend on the designated use of the water. Designated uses include recreational use—both with and without full body contact, drinking water, industrial processing, freshwater aquatic life, and marine aquatic life.

Depending on the amount of activity by the swimmer, comfortable temperatures range from 20 °C to 30 °C. Most people can tolerate short durations of lower and higher temperatures. For example, for a 30-minute period, most individuals can tolerate temperatures of 10 °C or 35 °C without harm.[2]



Temperature also affects the self-purification phenomenon in water bodies and therefore the aesthetic and sanitary qualities that exist. Increased temperatures accelerate the biodegradation of organic material, which may lead to total depletion of dissolved oxygen and the growth of anaerobic bacteria that release noxious and malodorous gases. At temperatures below freezing, a similar pattern exists, where colder water decreases the ability of organisms to perform necessary functions, and the quality of the water diminishes.

Temperature affects the amount of chlorination required to sanitize water for drinking purposes. The effectiveness of coagulation and filtration is reduced in water that is 5 °C, and water that is 10 °C requires three times as much chlorine to disinfect it than water that is 20 °C.[2]

More information is known about the upper temperature limits for various species than is known for the lower limits. Different upper limits apply to the various life processes of fish, with the upper limits for reproduction being lower than those for growth or survival (for examples, see the table below). Aquatic organisms can tolerate short periods of temperatures higher than those that are desirable to them, and the mobile forms can migrate to areas with more suitable temperatures. The number and distribution of bottom-dwelling (benthic) organisms decrease as water temperatures increase, with the upper tolerance limit for a balanced benthic population structure being approximately 32 °C.[2] It is known that a large assortment of organisms dwell in extremely hot environments such as thermal pools and ocean floor volcanic vents called black or white “smokers”. These organisms have adapted to extreme temperatures.

Through the natural changes in climatic conditions, the temperatures of water bodies fluctuate daily as well as seasonally. These changes do not eliminate indigenous aquatic populations, but they affect the existing community structure and the geographical distribution of species. Such temperature changes are necessary to induce the reproductive cycles of aquatic organisms and to regulate other life factors.[2]

Maximum weekly average temperatures for various fish species for various life processes[2]

	Temperature [°C (°F)]			
Species	**Maximum Tolerated***	**For Growth***	**For Embryo Survival****	**For Spawning****
Atlantic salmon	23 (73)	20 (68)	7 (45)	5 (41)
Bigmouth buffalo	—	—	27 (81)	17 (63)
Black crappie	—	27 (81)	—	—
Bluegill	35 (95)	32 (90)	34 (93)	25 (77)
Brook trout	24 (75)	19 (66)	13 (55)	9 (48)
Carp	—	—	33 (91)	21 (70)
Channel catfish	35 (95)	32 (90)	29 (84)	27 (81)
Coho salmon	24 (75)	18 (64)	13 (55)	10 (50)
Emerald shiner	—	30 (86)	28 (82)	24 (75)
Freshwater drum	—	—	26 (79)	21 (70)
Lake herring (Cisco)	25 (77)	17 (63)	8 (46)	3 (37)
Largemouth bass	34 (93)	32 (90)	27 (81)	21 (70)
Northern pike	30 (86)	28 (82)	19 (66)	11 (52)
Rainbow trout	24 (75)	19 (66)	13 (55)	9 (48)
Sauger	—	25 (77)	21 (70)	10 (50)
Smallmouth bass	—	29 (84)	—	17 (63)
Smallmouth buffalo	—	—	21 (70)	17 (63)
Sockeye salmon	22 (72)	18 (64)	13 (55)	10 (50)
Striped bass	—	—	24 (75)	18 (64)
Treadfin shad	—	—	34 (93)	18 (64)
White bass	—	—	26 (79)	17 (63)
White crappie	—	28 (82)	23 (73)	18 (64)
White sucker	—	28 (82)	20 (68)	10 (50)
Yellow perch	—	29 (84)	20 (68)	12 (54)

*As calculated using the formula specified by the EPA in the *Red Book*[2]

**As reported from field observations and reported in the EPA *Red Book*[2]: "—" means not calculated or observed.

Note: In general, the *maximum tolerated temperature* > *maximum temperature for growth* > *maximum temperature for embryo survival* > *maximum temperature for spawning*.

Measuring Procedure

Setting up the Temperature Sensor

Plug the Temperature Sensor into your data collection system �(2.1).

Task Result: On some data collection systems, a digits display will appear, by default, for temperature.

Taking measurements

1. **If required for your data collection system, begin by building an experiment. �(2.2)**
2. **Set your data collection system to Manual Sampling. �(2.3)**
3. **Record the temperature reading. �(2.4)**
4. **It is possible to display both Fahrenheit and Celsius values or kelvin. When building the data page, select two displays and change one display from the default temperature to the temperature of choice. �(2.2)**

References

1. *Healthy Water, Healthy People Testing Kit Manual*. 2002, Bozeman, Montana: The Watercourse, International Project WET. http://projectwet.org/water-resources-education/water-quality-education/ (accessed Sept 30, 2010).
2. *Quality Criteria for Water (Red Book)*. 1976, Washington, D.C.: U.S. Environmental Protection Agency.
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Thermocline

What is a thermocline?

A thermocline (sometimes called a metalimnion) is a layer within a body of water or air where the temperature changes rapidly with changing depth. Thermoclines can occur in both shallow and deep ponds and lakes and in the ocean.

Almost all sunlight is absorbed near the surface of the water, which heats up. Wind and waves circulate the water in the surface layer, distributing heat within it, and the temperature may be quite uniform near the surface. Below this uniform layer, however, the temperature drops rapidly. This area of rapid temperature transition is the thermocline.

Below the thermocline, the temperature may continue to drop with depth but far more gradually. In the Earth's oceans, 90% of the water is below the thermocline. This deep ocean consists of layers of equal density, being poorly mixed, and may be as cold as 0 °C to 3 °C.[1]

Temperature affects the density of water, so warm, less dense water does not mix well with cold, more dense water. (The effect is similar to floating oil on water.) The area between these water layers is the thermocline, where the water temperature rapidly transitions from that of the warmer layer to that of the cooler layer.

Why measure thermoclines?

Measuring thermoclines helps in understanding the ecosystem of a pond or lake. Thermoclines affect the mix of nutrients and dissolved oxygen in the water, which in turn affects the behavior of aquatic animals as they seek suitable environments to support their life processes.

Bass fishermen are aware of thermoclines, since bass will tend to stay where the most dissolved oxygen is, which is typically at the top of a thermocline when one is present in a lake. Bass are less frequently found below the thermocline, where there is less dissolved oxygen.

In the open ocean, the thermocline has been important in submarine warfare because it can reflect active sonar. This is possible because cold and warm water columns have different refractive indexes, causing the same effect that can be observed when hot air rises off the tarmac at airports or desert roads, producing mirages. Scuba divers may encounter thermoclines and see the visual distortion they produce.

If a thermocline is present in a water body, it is necessary to collect water samples from above it and below to understand the water quality of the entire body of water.

What factors affect thermoclines?

The principle factors that cause thermoclines to form are the following:

- Sunlight can only penetrate and heat the top of a water column, and the amount of heat from sunlight delivered to the water is inversely proportional to the depth;
- Wind and wave action cause the surface of the water column to mix and become relatively uniform in temperature, rather than slowly declining in temperature as would be expected if the water was perfectly still, whereas water at greater depth is not affected by wind and waves; and
- The cooler the water, the denser it is, down to 4 °C; however, temperatures of liquid water lower than 4 °C become progressively less dense until ice is formed.

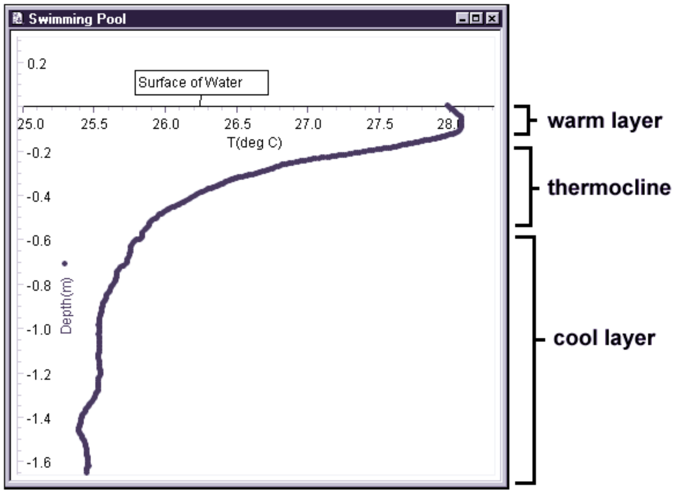
This last factor is fundamental to the "turnover" of the water column that occurs in the fall and spring. As winter approaches, the temperature of the surface water will drop until it approaches 4 °C, about the same as the temperature below the thermocline. When the entire body of water is at or close to 4 °C, "fall turnover" begins; the thermocline disappears (or, rather, it reaches the surface), and the water from the bottom of the lake can mix freely with the water from the top. This process is aided by wind or any other force that agitates the water.

As the water mixes, dissolved oxygen is carried to the bottom and nutrients from the bottom are carried to the top of the water column. In Arctic and Antarctic waters and along some deep coastlines, this effect (also known as upwelling) enriches surface nutrients and produces blooms of phytoplankton, making these very rich and diverse ecosystems. Since ice is even less dense than water at 0 °C, ice that forms on the surface stays there, insulating the liquid water underneath from the cold air temperatures. Thus liquid water is maintained under the ice, even though the surface air temperatures may be well below 0 °C.

In the spring as the surface warms, the ice melts. When the surface water reaches 4 °C, the free mixing of water occurs again, resulting in the "spring turnover." Dissolved oxygen and nutrients are again distributed throughout the water column. As the surface of the water column warms, the denser colder water sinks below it and the thermocline becomes reestablished.[2]

Bass fishermen know that during spring and fall turnover, bass can be found feeding at the surface. During the summer months they will stay close to the top of the thermocline, where the water is cooler than at the surface and thus holds more dissolved oxygen.[2]

Example of a thermocline curve measured in a swimming pool



The thermocline will be found at greater depths as more and more surface water warms. Deep lakes will allow the thermocline to exist at greater depths than shallower lakes. In a shallow body of water under conditions of little agitation, the thermocline may be quite near the surface (see picture). In deep lakes and the ocean, the thermocline may be located many meters from the surface and extend for hundreds of meters.[3] In any case, the curve of a thermocline graph of temperature versus depth has a similar characteristic shape.

What are ideal thermocline values?

There is no particular "best" depth for thermoclines to become established. For the vitality of an aquatic ecosystem, it is vital that thermoclines disappear periodically so the nutrients and dissolved oxygen can be fully distributed throughout the water column.

A deeper thermocline in the ocean (often observed during El Niño years) limits the amount of nutrients brought to shallower depths by upwelling processes, greatly impacting the year's fish crop.

Measuring Procedure

Setting up the Thermocline Sensor

Plug the thermocline sensor into your data collection system. �(2.1)

Task Result: A choice of measurements for the thermocline sensor appears.

Calibration note

It is not necessary to calibrate the thermocline sensor.

Taking measurements

1. **If required for your data collection system, begin by building an experiment.** �(2.2)
2. **Set your data collection system to Manual Sampling.** �(2.3)
3. **Insert the thermocline probe into the water.**
4. **Record the thermocline reading.** �(2.4)

More background information

More information about thermoclines and its role in water quality may be found in the references listed below; also refer to the Recommended Reading and Resources section.

References

1. Department of Atmospheric Sciences. *Thermocline*. 1999, University of Illinois. Weather World 2010 Project. http://ww2010.atmos.uiuc.edu/(Gh)/wwhlpr/thermocline.rxml (accessed Sept 30, 2010).
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3. University Corporation for Atmospheric Research. *Temperature of Ocean Water*. 2003, Windows to the Universe. University of Michigan. http://www.windows2universe.org/earth/Water/temp.html (accessed Sept 30, 2010).

Total Dissolved Solids

What are total dissolved solids?

Total dissolved solids (TDS) are the mineral and salt impurities dissolved in water and are measured in parts per million. Small organic solids (micromolecules) can also be a part of measurable TDS, but sugar, alcohol, pesticides, and large organic solids (macromolecules) are not included as total dissolved solids because they do not conduct a current.[1] TDS can affect the way drinking water appears, but toxic levels depend on the species of animals in or using the water. TDS is used as an indicator for chemical contaminants, and to map water movement in underground systems, or to build runoff models or basin drainage models.

Total dissolved solids consist of nutrient runoff from storm drains or road salts, and some organic compounds. Sources of TDS include agricultural runoff, residential runoff, and industrial point-source pollution. Micromolecules are constantly leached from soils and end up in the water supply. If the micromolecules are smaller than 2 micrometers and are water soluble, they are classified as dissolved solids. If they are larger than 2 micrometers and stay suspended in the water column, they are called suspended solids.

Determining TDS

Total dissolved solids can be determined in a number of ways, some more precise than others. With the gravimetric method, a sample of water is boiled until all liquid is evaporated, leaving a residue on the bottom of the beaker. The mass of the residue is then weighed using an analytical balance. This method, however, takes a great deal of time and has its own measure of error.

If there are ions present, a second “ballpark” method exists that uses the electrical conductivity reading of a sample to estimate the total dissolved solids. The conductivity reading of the sample is multiplied by a conversion factor that is based on the type of dissolved solids in the water. The conversion factor can vary between 0.40 and 0.96; a value of 0.65 is used as an approximation if the dissolved solids are not known.

TDS (ppm) = Conductivity (µS/cm) × conversion factor

Unlike the salinity portion of the Salinity Sensor, the conductivity reading takes into account every ionic compound dissolved into a sample, not just the ionic compounds associated with sea water, and has built-in temperature compensation. A conductivity measurement (in units of µS/cm) taken at 25 °C can be quickly converted to a good estimate of TDS in natural waters (in units of parts per million, ppm) by multiplying it by the conversion factor.

Conversion factors

Type of Water	Total Dissolved Solids (ppm)	Conversion Factor
Freshwater	0 – 2,200	0.7
Brackish water	2,200 – 8,300	0.6
Saline water	> 8,300	0.5

Why determine total dissolved solids values?

Measuring conductivity is a quick and easy way to estimate the amount of total dissolved solids (TDS) in natural waters, since most of these solids dissolve to form ions. TDS calculations are important when determining appropriate uses for water. Determination of the conductivity (and thus of the approximate concentration of TDS) gives information about the water quality as it relates to certain designated uses, such as for drinking water and for agricultural irrigation. The amount of mineral and salt impurities in water can affect the amount of salt build-up in agricultural land, the corrosion of pipes, and the toxicity of drinking water.

Calculating TDS can also be a useful tool for monitoring the inflow of saline water in estuaries and identifying sources of pollution, such as mining or industrial waste or agricultural runoff.

Total dissolved solids is therefore a parameter considered in water quality indexes (refer to the section on Water Quality Index).

What factors affect total dissolved solids values?

Like salinity and conductivity, TDS is determined by the concentration of ions in solution and the mobility of those ions. Temperature, evaporation, ice formation, and many of the other factors contributing to salinity fluctuations also affect total dissolved solids. Agricultural runoff, pollution, nutrient runoff, and a number of other sources associated with human activity also account for higher TDS concentrations.

Of all of these influences, the temperature of the water affects conductivity, and thus the measurement of total dissolved solids, the most. Increased temperatures cause ions to move more quickly; the probe reads this increase in activity as a higher concentration of ions. Conductivity measurements should be taken at 25 °C or a temperature compensation factor must be applied to the measurement. The conductivity reading of pure water can rise as much as 4.55% for every 1 °C deviation from 25 °C, while the reading of NaCl solutions rises 2.12%.[1]

The Conductivity Sensor has a built-in calculation to compensate for the change in conductivity due to temperature changes. If the Salinity Sensor is used, the compensation factor must be applied to the measurement.

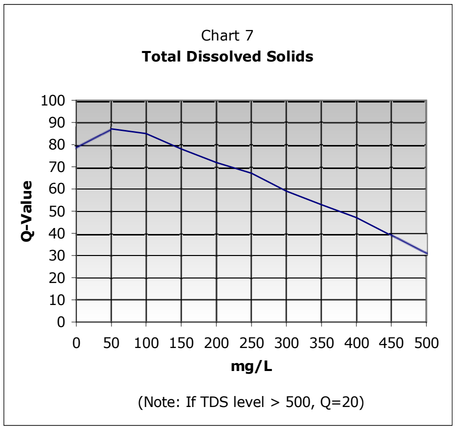
What are desirable levels of total dissolved solids?

TDS is most important to water quality when it concerns designated uses and has been listed by the EPA as a secondary ground water and drinking water contaminant. Secondary contaminants cause aesthetic, technical, and cosmetic effects. Water high in total dissolved solids may have an unpleasant taste, odor, or color and may also have a laxative effect beyond the tolerance level. High levels of TDS will also cause corrosion and mineral deposits in pipes, causing some industries to lower their maximum allowable levels.[2]

Normal and maximum TDS levels vary with the use and the source

	TDS Levels (mg/L)	
Designated Use	**Normal Ranges**	**Upper Limit**
Human consumption[2]	25–100	500
Livestock and fish consumption[4]	Varies with species	1000
Fresh water[5]	Depends on location	1500
Brackish water[5]	1500	5,000
Sea water[5]	nearly 35,000	100,000
Lakes and streams[5]	50–250	35,000
Rivers[5]	100–20,000	Depends on conditions
Groundwater[5]	Depends on location	Can be higher than 20,000

Generally, as the level of total dissolved solids increases, the Q-value declines and water quality diminishes. The only exception to this curve is when TDS is very low (less than 50 mg/L). A small amount of dissolved nutrients in a water body is healthy for cell development and function in organisms.



(Note: If TDS level > 500, Q=20)

Measuring Procedure for Conductivity

1. **Plug the Conductivity Sensor or Salinity Sensor into the data collection system.** �(2.1)
2. **Calibrate the Conductivity Sensor,** �(2.5) **or Salinity Sensor.** �(2.9)

Finding TDS

1. **If required for your data collection system, begin by building an experiment.** �(2.3) **Otherwise, set the display to show Conductivity measurements.** �(2.3)
2. **Set your data collection system to Manual Sampling.** �(2.3)
3. **Immerse the end of the probe into the sample of water to be measured. When the reading stabilizes, record the conductivity reading.** �(2.4)

   Note: Be sure to immerse the holes near the end of the sensor. Swirl the probe to be sure any air bubbles are removed from the end of the probe.

4. **Take additional measurements as needed, following the procedure above.**
5. **From the conductivity values, apply the conversion factor to calculate TDS (see "Determining TDS" above).**

References

1. Holmes-Farley, R. What is TDS? *Reefkeeping*. 2004, 4. www.reefkeeping.com/issues/2004-04/rhf/feature/ (accessed Oct 1, 2010).
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Turbidity

What is turbidity?

Turbidity is a measure of the cloudiness of water, indicating the presence of suspended solids and colloidal material, such as clay, silt, sand, sediment, minerals, microorganisms, and chemicals. Technically, turbidity is an optical property of the water based on the amount of light reflected by suspended particles.

Turbidity is measured with a nephelometer, which uses a light beam and photoelectric cell to electronically measure the amount of light scattered by suspended particles, in nephelometric turbidity units (NTU).[1]

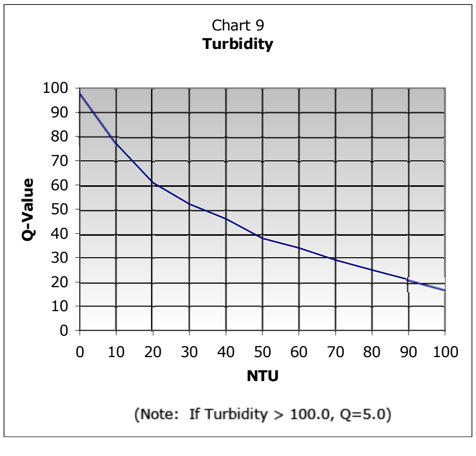
Turbidity is not the same as clarity. Clarity is a property of water that is affected not only by suspended solids but also by chemicals that color the water. Clarity is measured using a Secchi disk, which is a disk with black and white areas on it. Clarity is quantified in terms of the maximum number of meters that the Secchi disk can be lowered into the water and still be seen. A rule of thumb is that light can penetrate water 2 to 3 times the maximum depth at which a Secchi disk can be seen.[4]

Turbidity is also not the same as total suspended solids because different types of suspended solids reflect light differently. Total suspended solids are determined by filtering the water and measuring the mass of the solid material that is filtered out of the water after it has been dried. However, measuring turbidity is a quick way to estimate the amount of total suspended solids.[1, 4]

Why measure turbidity?

Turbidity is one of the most common water quality problems. It causes significant problems, no matter what the designated use (see the table on detrimental effects of excessive turbidity below). Turbidity is regulated under the EPA's Primary Drinking Water Standards, and it is of such concern that daily measurement is required for treated drinking water.[2]

Turbidity indicates a possible presence of contamination with pathogenic organisms. Turbid water is more difficult to sanitize, since suspended matter provides areas for the pathogens that are sheltered from the sanitizing chemicals, such as chlorine.



Turbidity resulting from eutrophication, or the presence of high levels of nutrients such as nitrates, may progress to anoxic water that supports the growth of anaerobic bacteria, with the resultant release of noxious or malodorous gases.

Turbidity blocks the penetration of sunlight through the water, resulting in reduced plant and algae growth and reduced dissolved oxygen concentrations and food sources for fish and other organisms. The large amounts of silt and clay in some turbid water settle on the bottom, blocking bottom-dwelling organisms, eggs, and developing embryos and larvae from essential dissolved oxygen, causing their death. The suspended solids can clog or damage fish gills, suffocating them.[3]

Turbid water absorbs solar radiation, causing increased temperatures in the surface water and the formation of a thermal layer, with the reduced mixing in the water column. This causes lower levels of dissolved oxygen at the bottom of the water column.[1]

Turbidity can cause unsafe swimming conditions, hiding submerged objects and hindering rescue efforts.

Highly turbid water used for irrigation can clog pipes, coat the ground with a water-impermeable layer, and coat leaves, impeding photosynthesis, stunting growth, and causing aesthetic problems on leafy green vegetables. Sediment from turbid waters fills and clogs irrigation reservoirs and channels.[3]

Measuring turbidity is an easy, reliable way to monitor the effectiveness of preventative or remedial measures to combat erosion. In one study, turbidity was used as an indicator to demonstrate the effectiveness of leaving a 100-foot area next to a stream uncut during the clear-cutting of a watershed.[1] In another study, turbidity was monitored over several years to demonstrate the effectiveness of improved irrigation procedures on the lower Yakima River in Washington State.[5]

Detrimental effect of excessive turbidity by designated use

Designated Use	Detrimental Effect
Drinking water	Requires more chlorine to sanitize, aesthetically unpleasant
Swimming, recreation	Hides submerged objects, hinders rescue efforts, more likely to contain pathogens and requires more chlorine to sanitize, aesthetically unpleasant.
Aquatic animal life	Causes reduced dissolved oxygen or anoxic water; smothers and kills eggs, embryos, larvae and some bottom-dwelling organisms; clogs and damages gills.
Aquatic plant life	Reduces the amount and depth of penetration of sunlight, reducing plant growth, causing reduced dissolved oxygen concentrations and food for aquatic animals; smothers plants growing on the bottom.
Irrigation	Sediment clogs pipes and fills reservoirs; coats ground with water-impermeable layer; coats plant leaves, impeding photosynthesis; reduces marketability of leafy vegetables.

What factors affect turbidity values?

Factors that affect turbidity can be classified into natural and human-caused, and they can cause either an increase or decrease in turbidity, as follows:

Natural factors that increase turbidity

- Torrential rains, causing erosion, even in natural areas
- Increased temperature, providing an environment for increased growth of algae, plankton, and bacteria
- Rapid flow of water keeping suspended solids mixed into the water
- Breakdown of the thermal layer of a lake in the fall, causing mixing in the water column and algal blooms at the surface

Natural factors that reduce turbidity

- Inhibition of erosion by well-established plant life
- Filtration through roots, mulched material, gravel, and sand
- Low or no flow, allowing suspended solids to settle out of the water

Human-caused factors that increase turbidity

- Accelerated eutrophication due to the addition of nutrients to water, often through run-off from farms, golf courses, and urban areas
- Improper cultivation and irrigation practices on farms
- Clear-cutting to the water's edge in watersheds
- Improperly constructed logging roads
- Improper disposal of water from mining operations

Human-caused factors that decrease turbidity

- Preventative and remedial practices in watersheds, such as leaving buffer zones by streams, proper construction of logging roads, proper cultivation, irrigation, and mining practices
- Allowing water to stand undisturbed over a period of time can reduce turbidity (for example, water impounded in a reservoir allows some of the clay, silt, and sand to settle out of it)
- Filtration, as in water treatment
- Chemical sanitation, reducing the growth of suspended organisms

What are ideal turbidity values?

For almost every designated use, the less turbid the water, the better. The standards are highest for drinking water. Recommendations may refer to absolute turbidity values or to increased turbidity values over natural baseline values (see the following table). One exception is that turbid water absorbs pesticides, which may be useful in some situations.[3] In another exception, very low turbidity may indicate water that has insufficient nutrients to support a healthy natural aquatic community.[1]

Aquatic organisms are less bothered by occasional natural increases in turbidity (such as during spring floods) than they are by sustained elevations in turbidity, such as those in areas of increased human-caused erosion. For example, some species of fish can tolerate very high levels of turbidity, even more than 10,000 NTU for several days, but months of exposure to turbidities as low as 500 NTU can be fatal. Long-term exposure to turbidity levels as low as 10 NTU to 50 NTU can inhibit growth and reproduction.[4] One study showed marked variations between fish species regarding tolerance for turbidity; one species (smelts) had a 50% mortality rate after exposure for 24 hours to water with turbidities of 1700 NTU to 3000 NTU, while turbidities of 17,500 NTU to 21,000 NTU were required for 50% mortality for another species (inanga).[6] Young organisms are usually more sensitive to turbidity than older ones.

Some fish species exhibit avoidance behavior or refuse to spawn at turbidities as low as 20 NTU.[6]

Maximum recommended turbidity levels by designated use

Designated Use	Maximum Turbidity Level (NTU)	Reference
Drinking water	0.3–5 (depends upon circumstances)	Ref[2]
Swimming, recreation	50 NTU	Ref[7]
Aquatic life	1500 NTU (depends upon the species)	Ref[6]
Aquatic life	Increase of 8 NTU over a natural background of 80 NTU or less; Increase of 10% when a natural background is > 80 NTU	Ref[7]
Irrigation, terrestrial life, industrial use	Increase of 10 NTU when natural background is < 50 NTU; Increase of 20% when natural background is > 50 NTU	Ref[7]

Measuring Procedure

Setting up the Turbidity Sensor

1. **Plug the Turbidity Sensor into your data collection system.** �(2.1)

   Task Result: On some data collection systems, a digits display will appear, by default, for turbidity.

2. **Calibrate the Turbidity Sensor.** �(3.2)

Preparing samples

1. **Collect a representative sample from the water body in a clean container with a lid.**
2. **Remove any floating debris, and if necessary, filter though a coarse filter to remove any large particles.**
3. **Invert the container several times to redistribute any sediment.**
4. **With a pipet, put at least 6 mL of the test water from the middle of the sample into the cuvette (filling the cuvette).**

   Note: Handle the cuvette with a lab tissue. Wipe finger prints or water from the outside of the cuvette.

5. **Screw on the cap of the cuvette.**

Taking a turbidity measurement

1. **If required for your data collection system, begin by building an experiment.** �(2.2)
2. **Set your data collection system to Manual Sampling.** �(2.3)
3. **Insert the sample cuvette into the Turbidity Sensor and cover it with the cap.**
4. **Record the turbidity reading.** �(2.4)

More background information

More information about turbidity and its role in water quality may be found in the references listed below; also refer to the Recommended Reading and Resources section.

References

1. *Healthy Water, Healthy People Testing Kit Manual*. 2002, Bozeman, Montana: The Watercourse, International Project WET. http://projectwet.org/water-resources-education/water-quality-education/ (accessed Sept 30, 2010).
2. *National Drinking Water Standards*. U. S. Environmental Protection Agency. 2002.
3. *Quality Criteria for Water (Red Book)*. 1976. Washington, D.C.: U.S. Environmental Protection Agency.
4. *Turbidity*. Water on the Web. http://www.waterontheweb.org/under/waterquality/turbidity.html (accessed Oct 1, 2010).
5. *Washington: Lower Yakima River. Changes in Irrigation Practices Reduce Turbidity*. Nonpoint Source Program Success Story 2005. http://www.epa.gov/owow_keep/NPS/Success319/state/wa_yakima.htm (accessed Oct 1, 2010).
6. Rowe, D., Smith, J., and Williams, E., *Lethal Turbidities and Native Fish*. Water & Atmosphere (National Institute of Water and Atmospheric Research, New Zealand), 2004. 12(2): p. 26–27.
7. Singleton, H. *Summary of Water Quality Guidelines for Turbidity, Suspended, and Benthic Sediments*. 2001. The Ministry of Environment, Land and Parks, British Columbia, Canada. http://www.niwa.co.nz/news-and-publications/publications/all/wa/12-2/native (accessed Oct 1, 2010).

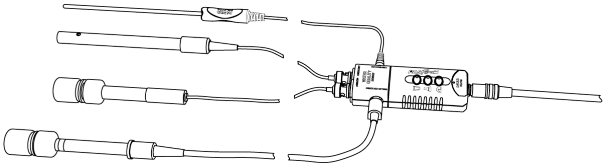
Additional Options

Water Quality Sensor

Overview

The PASCO Water Quality Sensor is a MultiMeasure Sensor™ that comprises the following sensors:

- Temperature
- Conductivity
- pH
- Dissolved Oxygen



The Water Quality Sensor enables the collection of multiple types of water quality data simultaneously from the same sample. The sensor rapidly alternates reading each sensor, so that there is no electrical interference between sensors, such as would occur when using separate sensors in the same sample. Using the Water Quality Sensor therefore allows the collection of these four parameters simultaneously.

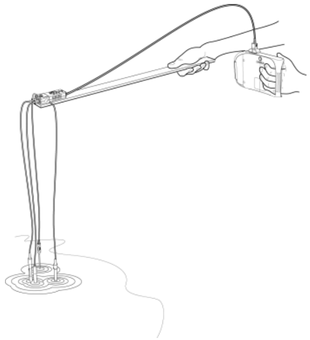
Using the Water Quality Sensor

Plug the Water Quality Sensor into a data collection system. �(2.1) Depending on the data collection system you are using, the parameters will automatically display in different ways. A digits display of 2 parameters (dissolved oxygen and conductivity) may pop up automatically, or you may need to build an experiment to display the parameters. Depending upon the data collection system, up to 8 parameters may display at once. �(2.2)

The cord on the dissolved oxygen sensor is quite long to enable you to measure the amount of dissolved oxygen at the bottom of a water column as well as at the top and midway. If you only want to sample the dissolved oxygen at the same level as the other sensors, simply coil up the cord to the desired length and secure it with a twist wire or tape.

If you want to sample water farther out from the bank. Obtain a long pole or a meter stick. Tape the sensor onto the end of the pole or stick, and allow the probes to hang down from the sensor. Use a sensor extension cable to attach the sensor to the data collection system. Hold the pole or stick at the other end and extend the pole over the body of water as far as you can reach. Dip the probes into the water and wait for the data to stabilize before you capture it.

> **Note:** Do not allow the sensor to touch the water! Only the probes should go into the water.



For many general water quality studies in which you only wish to see relative changes, the sensors do not need to be calibrated. When you plug the Water Quality Sensor into the data collection system and turn on the power, you are ready to start measuring. When you immerse the sensors into the water sample, the display will show the 4 measurements simultaneously. For true values, or to share data with a government agency or water resources board, the sensors must be calibrated before data collection begins. �(2.5), �(2.6), �(2.7)

Calibrating the Water Quality MultiMeasure Sensor

When you are conducting comparative studies, you may need to calibrate the sensors to standardize the measurements. Calibrating the Water Quality Sensor is similar to calibrating any single sensor that is connected to the data collection system. Once the Water Quality Sensor is plugged in, select the particular sensor you want to calibrate and use the Tech Tip for that sensor. �(2.5), �(2.6), �(2.7)

GPS/GIS

GPS Position Sensor Overview

Getting good sensor data is only part of a complete water quality investigation. The analysis of the data and the cataloging of information should also be part of a large-scale study.

In some cases, the analysis of the data can be accomplished by comparing it to national standards or by graphing changes over time or depth. In other cases, the study of daily, weekly, or yearly fluctuations will yield more information about water quality, as will mapping the data to look at spatial relationships.

Mapping water quality data begins with locating the sample site and transferring the data to a map. Cataloging the data requires that the information be entered into a spreadsheet or a database for further analysis. The combination of GPS and GIS can accomplish all of these goals.

The PASPORT GPS Position Sensor uses signals received from the Global Positioning System (GPS) to determine its latitude, longitude, and altitude anywhere on earth. It is designed to be used alone or in conjunction with the data collected from other PASPORT sensors.

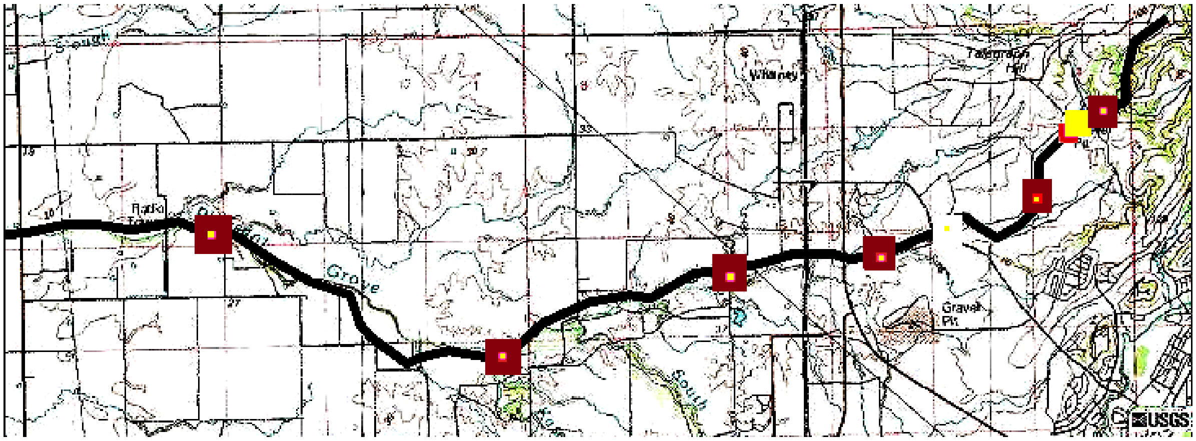
Whenever it is powered on, the GPS Position Sensor continuously receives data from GPS satellites, even if data is not being recorded by the data collection system. When sampling, the data collection system will collect data from all the sensors plugged into the interface as well as from the GPS Position Sensor. Data can be collected continuously or as manual samples. In addition to collecting sensor data, you can also time stamp your data for creating long-term monitoring projects.

GIS Function

My World GIS™ is a Geographic Information System (GIS). With My World GIS, students can analyze water quality data based on spatial relationships. Consider the following example showing samples collected along a creek with the GPS Position Sensor and the Water Quality Sensor:

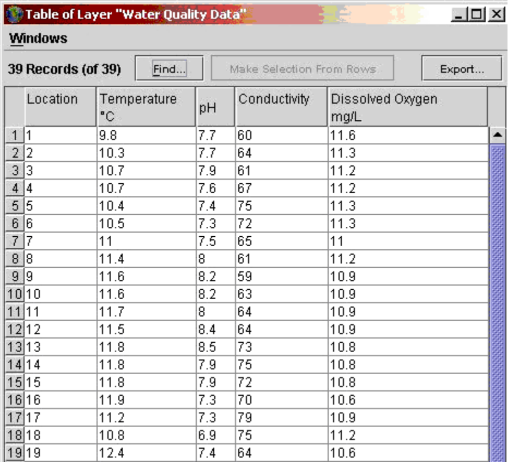
pH data mapped over an aerial photo of the Pleasant Grove Creek area

The same pH data mapped over a topographic map of the area



By mapping the data spatially, connections can be made to the physical surroundings.

A table display of the data (shown at right shows the connection to all collected data and exemplifies My World GIS as a database for storing water quality information for long-term studies.



39 Records (of 39)

	Location	Temperature °C	pH	Conductivity	Dissolved Oxygen mg/L
1	1	9.8	7.7	60	11.6
2	2	10.3	7.7	64	11.3
3	3	10.7	7.9	61	11.2
4	4	10.7	7.6	67	11.2
5	5	10.4	7.4	75	11.3
6	6	10.5	7.3	72	11.3
7	7	11	7.5	65	11
8	8	11.4	8	61	11.2
9	9	11.6	8.2	59	10.9
10	10	11.6	8.2	63	10.9
11	11	11.7	8	64	10.9
12	12	11.5	8.4	64	10.9
13	13	11.8	8.5	73	10.8
14	14	11.8	7.9	75	10.8
15	15	11.8	7.9	72	10.8
16	16	11.9	7.3	70	10.6
17	17	11.2	7.3	79	10.9
18	18	10.8	6.9	75	11.2
19	19	12.4	7.4	64	10.6

Transferring the data to My World GIS is a simple two-step process. First download the data to a USB flash drive, then open the data set from within My World GIS. For additional information about using GPS and GIS together, see the related pages in the PASCO scientific website.

Using Coordinate Markers

Both the Xplorer GLX and the SPARK Science Learning System have the ability to create coordinate markers for use in applications where returning to a very specific location at a later date is important.

For information on using coordinate markers, see Tech Tip 2.8.

Teaching Strategies

Teaching Tips

Teaching Strategies for Water Quality Studies (case study)

The approaches you might use to teach water quality concepts are as varied as your imagination. However, it might help to hear about strategies used by one teacher, Roger Palmer.[1] Palmer is highly experienced teaching Earth and Environmental Sciences to junior high and high school students using a combination of in-class lab work, field trips to collect data on-site, field trips to local water treatment facilities, collaborations with local and regional parks services, and collaborations with local TV stations. He makes the following points about successful water quality studies with students.

Start simply, testing water samples in the school laboratory setting

For starters, you will want students to become acquainted with using sensor and test-kit technologies in the more controlled environment of the classroom. This means that you will be going out the day before and collecting water samples (see the section on Water Sampling Equipment for suggested equipment for these sample collections) from a local waterway for use in the classroom. As long as you fill your storage containers with water, cap them, and put them on ice, they will maintain their character sufficiently for classroom investigations overnight or even for a couple of days. Alternatively, test tap water, since students will be interested to know more about it.

Special conditions:

1. For samples intended for testing dissolved oxygen, avoid unnecessary or prolonged contact with air.
2. Very turbid water can be allowed to settle overnight—rather than filtering it—and it will be suitable for most analyses in the classroom. Just agitate the sample before you measure its turbidity.

Have students test only one or two parameters, for example dissolved oxygen and conductivity, in their first lab experience. It may take a couple of lab periods before everyone learns how to use the technology, so planning for repetition through testing different types of water samples or using different sensors or test kits at each lab period will provide the cumulative experience that will result in successful testing and concept development. See the section below, "Ideas for 'getting your feet wet' with water quality studies," for a listing of simple activity ideas for each sensor.

Don't worry too much about calibrating equipment at this stage. The sensors out-of-the-box will provide a measurement that is within an acceptable range for teaching purposes. Later, when students gain more experience and confidence using the sensors, the concept of calibration can be introduced.

> **Note:** However, the Water Quality Colorimeter should be calibrated before use, because the calibration button might accidentally be pressed, throwing the calibration off. Since calibration of the Water Quality Colorimeter is an easy procedure, this should not pose a problem. Additionally, if you are located at a high elevation, you will need to calibrate the Dissolved Oxygen Sensor for representative results.

Take advantage of the multiple measurements that students from different classes will make on the same samples with the same equipment

Inevitably, students will obtain a wide range of values for a parameter from the same water sample. This variability can be instructive. Examining all the data, outliers can be identified and discarded (as well as discussed regarding the reasons for their occurrence, which might include sensor malfunction or errors in testing methodology), and a reasonable hypothesis as to the "real" value can be arrived at. This can be a rich opportunity to elaborate on concepts such as experimental error, sample decay, precision and accuracy, and experimental uncertainty.

Encourage the development of a story or the identification of patterns

Provide a context regarding the water source. Take photos or videos of the collection site to show the students in the classroom, and point out features that might influence or be influenced by the quality of the water.

Divide the class up such that small groups of students or individual students are working on a particular aspect of the whole picture. For example, students working in pairs could examine one of the water quality parameters each, while other student pairs could collect information about the watershed for the water source; others could document the weather at the time of the water sampling and testing; still others could document the aquatic vegetation and animal life.

Serial observations enhance story or pattern development. One scenario would be to plan four different lab sessions at different times of the year, revisiting the same sites. Alternatively, different sections of a waterway might be examined at different lab periods.

Then in a culminating event, these various aspects can be combined so that patterns will be revealed, providing meaning to the water quality testing experience.

Foster emerging student interest as the year progresses

You will probably find that a small number of your students are especially interested in the water quality testing labs you organize. As the year progresses, you may be able to involve these students in water sample collection excursions after school.

Look for opportunities to engage community resources

Potential community resources include:

- Water treatment facilities
- City, state, or national park managers
- City planning officials
- University faculty
- Research centers
- Local TV or radio stations
- Regional or national initiatives, such as the Globe Project, Jason Project, Project Wet, and WaterWatch Project

For example, the local water treatment plant may allow some of your students to come in and test some of their water samples on their equipment. After some experience in the school lab with sensor and test kit technology, your students would benefit largely from the opportunity to see the testing equipment in an industrial setting. For example, Palmer was able to arrange for several of his students to test their water samples at the local water treatment plant, and the operators showed them how to use all of their equipment, including the high pressure liquid chromatography (HPLC) device and spectrophotometer.

The local TV station may be engaged to film the culminating activity that students put together with the total set of data they have collected during the year. For example, Palmer's classes reported on their project for the local TV station's Evening News for three years in a row.

A local park service may invite your students to present their culminating report at a public information session. For example, Palmer lined up a scheduled "campfire" session held on Friday nights during the summer by a local park as a venue for his students to report their findings.

Try to incorporate at least one field trip for on-site sampling into your water quality studies

Palmer points out that field study can provide the following benefits:

- Builds group identity
- Supports kinesthetic learning
- Models the community of learners called for in the standards
- Always becomes cross-curricular
- Stretches you as an instructor
- Fosters fun in learning
- Clarifies the authenticity of projects, which become self-evident to students
- Gives us a sense understanding about our watersheds, producing a sense of connection to them and responsibility for them
- Instills a sense of wonder about your local area

Additionally, he points out that multiple concepts can be taught from the field, including the following:

- **Geology**—land forms, soil types, surface geology, bedrock geology, hydrology, natural history
- **Earth**—meteorology, rocks and minerals, river made landforms, mineral cycles, material hardness, energy and water budget
- **Biology**—taxonomy, ecology, invasive species, population dynamics, biotic interactions, biomes, macroinvertebrates, human health
- **Chemistry**—water quality, spectroscopy, heat capacity, gas pressure, solutions, concentrations, solubilities, acid base reactions, stoichiometry, chemistry involved in nutrient cycles, density, conductivity, battery chemistry
- **Physics**—reflection, absorption, scattering, refraction of light, speeds of sound and light, stream velocity, GPS, animal tracking, velocity calculations, electronics of sensing equipment, projectile calculations, design of habitat boxes, friction
- **Environmental**—land use, population studies, human footprint/impact on earth, pollution, alternate energy, water treatment, natural resource management, succession, plant location preferences, economics

Note: Remember that field trips do not necessarily involve buses. There may be a suitable water body within walking distance of your school.

Pre-trip planning

- Try to identify a suitable water source that is within walking distance of your school. An ideal site might be a shallow, relatively clean river that has a low bridge crossing for easy water sampling
- Arrange transportation, if necessary
- Gather information about the natural history of the area from the local park service, local history from the local library, and topology from a good online map, to enrich the experience
- Prepare students with a check list of items each is required to bring
- Get appropriate parental consent

SAFETY PRECAUTIONS

- Visit the site ahead of time and look for potential problems, including dangerous footing, poisonous snakes or invertebrates, water hazards, potential for getting lost, and plan to prepare students accordingly.
- Use snapping lanyards to attach the data collection systems to the students. This will prevent the data collection systems from falling into the water. If a lanyard cannot be attached to the data collection system, develop a technique with your students to prevent the device from accidentally falling into the water. If the device should happen to fall into the water, call Teacher Support immediately for help in drying and salvaging the system. (Inside the U.S. call 800-772-8700; outside the U.S. call 916-786-3800.)

Student notebooks

Every student should have one (even if you don't use them at first). As your sophistication grows, include:

- Aerial photographs
- Data entry forms
- Interpretive photos
- Species list
- Sketches of the site, organisms, etc.

Start small and leave lots of room between ideas to grow.

Packing for the trip

Some suggested equipment for your field trip:

- Data collection systems (fully recharged), sensors, test kits
- Student notebooks
- Water sampling device (refer to the section below on Water Sampling Equipment)
- Plastic ample storage bottles with caps (For samples to be brought back to the school lab for testing)
- Ice chest and ice (for storing collected water samples)
- GPS Position Sensor
- Wading boots
- Seining or kick nets for collecting macroinvertebrates
- Telescoping poles for long-reach sampling (away from the bank)
- Binoculars
- Field microscope
- Camera, extra batteries, extra film (if necessary)
- Digital camera or camcorder
- Sunscreen
- Mosquito repellent.
- Water bottles
- Non-perishable snacks
- First aid kit with pocket knife

- Extra clothing if contact with water is anticipated
- Rain gear, if needed

Follow tried and true on-site strategies

Some suggestions for managing students in the field include the following:

- Make a base to bring samples back for testing, analyzing
- If samples can be brought back to school lab, do it!
- For multiple outings, get into a packing routine
- Establish the clean-up routine for coming off the field
- Use a GPS Position Sensor to enable mapping of data to increase its meaning for students (data can be uploading into My World GIS™ back at school)
- Keep it simple-focus each student pair on one or two activities
- Emphasize the need to develop a story—encourage collaboration between student pairs or groups in doing so
- Measure other aspects of the environment, such as temperature, humidity, barometer, rainfall, wind, insolation, light intensity, etc., and challenge students to incorporate these findings into the overall story

Organize a summer school session focused on field experiences

A summer session may be the ideal way to teach water quality studies. A dedicated course allows you to get students into the field for multiple sequential days in a total-immersion experience. For example, Palmer organized summer sessions that were 2-weeks in length comprising 9 or 10 day trips. Because the course involved 10-hour days, from 7 a.m. through 5 p.m., his students were able to complete a full course of study in the 2-week period. The repetition involved was a real confidence builder for students, Palmer says, and the experience intensified the advantages listed above regarding field studies.[1]

Additional ideas

Have students start a club

Allow fundraising to pay for overnight field studies (camping at a local water resource) and participate in community service and field service opportunities. Most communities have some sort of restoration project going on in conjunction with schools. National Parks, State Park agencies and your local municipal park agency can recommend programs to become involved with. Also look into local naturalist organizations.

Water Sampling Equipment

To obtain representative water samples, you need to collect your samples away from the shoreline and below the surface of the water. Special sampling devices are available that are made specifically for this purpose. Their shape is cylindrical with closable ends, weights, and attached long ropes. Some examples include the following:

Available through Forestry Suppliers, Inc.:

- Wildco Student Water Bottle Kit
- Wildco Alpha Horizontal and Vertical Water Bottles
- LaMotte D.O. Water Sampler

Also handy to have are long poles with a device on the end to hold a plastic water bottle. You can make these with supplies from the local hardware store, or you can order them ready-made from Forestry Suppliers or other companies specializing in such gear; some examples are as follows:

- Swing Sampler
- Conbar Telescopic Dipper

Refer to the Recommended Reading and Resources section for contact information.

Ideas for "Getting Your Feet Wet" with Water Quality Studies

Perhaps the most difficult aspect of water quality studies is to get started using them with your classes. Once you have some experience with them, you'll be hooked, and your imagination will provide all the ideas you need.

Start by picking one or two activities that involve only one or two sensors or test kits. Here are some ideas for simple activities that will help you and your students get started and gain confidence using the sensors and test kits.

1. Test one or two parameters in a variety of drinking water samples

Collect drinking water (students can help) from as many sources as possible, including:

- your tap water
- well water
- water from neighboring communities that might draw their drinking water from a different source than your community, or that might have a different chlorination or fluoridation policy than your community
- bottled mineral water-several types, if possible
- bottled distilled water
- carbonated water

Collect two samples from two different locations that seem likely to show some differences, to compare the values for one or two of the following:

- conductivity
- chlorine
- fluoride
- dissolved carbon dioxide
- iron (if your community is one with appreciable iron in the water)
- pH
- total hardness

2. Test one or two parameters in a variety of swimming pool waters

If your community has a number of swimming pools, collect water (students can help) from several of them. Test one or two of the following that seem likely to show some differences:

- alkalinity
- chlorine
- conductivity
- dissolved oxygen
- pH

- phosphate
- total hardness

3. Compare conductivity, total hardness, and total dissolved solids

In a sample of water that you have found (through previous activities) to have high conductivity (swimming pool water or well water might be examples) measure the conductivity and total hardness. Then determine the weight of the dissolved solids by weighing an empty beaker, filling it with the water sample, boiling it dry, and measuring the beaker again. Find relationships between the three measurements.

> **Note:** You may need to boil down several beakers of water to obtain an appreciable difference in mass.

4. Compare temperatures at different points on a water body

If you have a convenient water body on or near school grounds, determine the temperature at different points, including shallow, deep, stagnant, swiftly flowing, sunny, and shady spots. Experiment with mapping this data using the GPS Position Sensor and My World GIS, if your school has access to these technologies.

5. Find the thermocline in a swimming pool

If you have access to a swimming pool, pick a time when the pool filter has been off for a day or two, no one has been swimming in it, and the day is quite warm. Plug in the Thermocline sensor to the data collection system and display a graph. Slowly drop the Thermocline Sensor into the deep end while recording the temperature and depth. See if you find a depth where a relatively sharp drop in temperature occurs. This is the thermocline.

6. Compare DO and temperature profiles in a local lake

Measure the DO and the temperature in a local lake or reservoir using a water sampler (see the section on Water Sampling Equipment for ordering information). Take samples at 1- to 2-meter intervals. Plot out the temperature and DO profiles and find the thermocline. See of any relationship between the thermal gradient and the DO gradient can be observed. Make repeated visits over the course of the year to observe the Winter and Summer profiles and the Fall and Spring overturns.

7. Conduct Izaak Walton League SOS macroinvertebrate counts

Study the macroinvertebrate populations for indications of water quality. The Stream Study is a macroinvertebrate study project developed through the University of Virginia and is based on material developed by the Save Our Streams Program of the Izaak Walton League of America.

8. Explore point-source pollution in your community

If your community has a notable problem with point-source pollution into a river or stream, collect water samples (using appropriate protective gear) from places above, at, and below the point source, bring the samples into class, and test for one or two parameters (using appropriate protective gear). Examples of possible point-source pollution include:

- cattle or dairy farm
- row-crop farm
- heavy industrial complex
- culvert that collects runoff from city streets
- golf courses

- construction sites
- zoos and animal preserves
- sewage treatment plant
- power plant
- mining operation

Measure one or two parameters that you expect might be affected.

If you have the GPS Position Sensor and My World GIS, map the collection points relative to the point-source of the pollution that you find.

9. Explore the effect of rainfall runoff on a local river

If your community has a suitable river or stream, collect water samples before and after a storm involving heavy rainfall and measure the turbidity, conductivity, hardness, and pH.

See if you can determine whether acid rain might be influencing the pH or whether a significant diluting has occurred. See if conductivity, turbidity, and hardness values are affected and develop hypotheses to explain any changes noted.

> **Note:** This is always a fun trip to the stream students have gotten to know so well over the course of the year. When they see it in flood stage, they are usually quite impressed with the changes in its characteristics and are excited to make the new measurements.

10. Explore the effect of algal bloom on water quality

If your community has a pond or lake that undergoes seasonal algal blooms, take samples before, during, and after the bloom and measure the effect on one or two of the following:

- BOD_5
- dissolved oxygen
- dissolved CO_2
- nitrate
- pH
- phosphate
- turbidity

References

1. Roger Palmer presented much of this information at the 2006 NSTA Conference in Dallas, TX.

Ordering Information

Ordering Information

Water Quality Parameter	**PASCO Measuring Technology**	**Measurable Ranges (mg/L; ppm)**
Alkalinity	Alkalinity ezSample Field Titrator Kit	10 – 100 mg/L (ppm)
Ammonia	ezSample Snap Vial – Ammonia*	0.2 – 3 mg/L (ppm)
Biological Oxygen Demand	PASPORT Dissolved Oxygen Sensor or PASPORT Water Quality MultiSensor	0 – 20 mg/L (ppm)
Chlorine	ezSample Snap Vial – Chlorine *	0.5 – 6 mg/L (ppm)
Conductivity	PASPORT Conductivity Sensor or PASPORT Water Quality MultiSensor	0 – 100,000 µS/cm
Dissolved Carbon Dioxide	Dissolved CO_2 Gas ezSample Field Titrator Kit	10 – 100 mg/L (ppm)
Dissolved Oxygen	PASPORT Dissolved Oxygen Sensor or PASPORT Water Quality MultiSensor	0 – 20 mg/L (ppm)
Hardness	Total Hardness ezSample Field Titrator Kit	20-200 mg/L (ppm)
Iron	ezSample Snap Vial – Iron*	1.5 – 8.0 mg/L (ppm)
Nitrate	ezSample Snap Vial – Nitrate*	0.25 – 2 mg/L (ppm)
pH	PASPORT pH Sensor or PASPORT Water Quality MultiSensor	0 – 12 pH units
Phosphate	ezSample Snap Vial – Phosphate*	0.20 – 8 mg/L (ppm)
Salinity	PASPORT Salinity Sensor	1 – 55 ppt 1,000 – 100,000 µS/cm
Stream Flow	PASPORT Flow Rate/Temperature Sensor	0 – 3.5 m/s
Temperature	PASPORT Temperature Sensor or PASPORT Flow Rate/Temperature MultiSensor or PASPORT Water Quality MultiSensor	-35°C – 135°C or -10°C – 50°C or -35°C – 135°C
Thermocline	PASPORT Thermocline Sensor	0 – 10.5 m (346 ft) 0-100 °C
Total Dissolved Solids	PASPORT Conductivity Sensor or PASPORT Water Quality MultiMeasure Sensor or PASPORT Salinity Sensor	0 – 100,000 µS/cm 0.5 – 55 ppt 1,000 – 100,000 µS/cm
Turbidity	PASPORT Turbidity Sensor	0 – 400 NTU

Additional Measurement Sensors	
Global Position	PASPORT GPS Sensor
	PASPORT Water Quality Colorimeter – to use with the ezSample kits

* Requires the PASPORT Water Quality Colorimeter

Standard solutions available from PASCO

Item
pH Sensor: pH Buffer Set (pH 4, 7, and 10)
Turbidity Sensor: Replacement 100 NTU Cal Standard

For additional information and to order, visit our Web site: **www.pasco.com**

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Recommended Reading Resources

Recommended Reading

General Water Quality, Methods, and Standards

Variously dated. *National field manual for the collection of water-quality data:* U.S. Geological survey techniques of water-resources investigations. 9 Chaps. A1–A9. *Available from:* http://water.usgs.gov/owq/FieldManual/

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Other Informative Water Quality Web Sites

Water on the Web

"About Us: Our primary goal is to train students to understand and solve real-world environmental problems."

Available from: http://www.waterontheweb.org/aboutus/index.html

Izaak Walton League SOS macro invertebrate counts

Information about using macro invertebrates as indicators of water quality.

Available from: http://people.virginia.edu/~sos-iwla/Stream-Study/StreamStudyHomePage/StreamStudy.HTML

Hach

Source for reference standard solution preparations and fecal coliform test kits.

Available from: http://www.hach.com

SK Science Kits and Boreal Labs

Source for fecal coliform test kits.

Available from: http://sciencekit.com/fecal-coliform-test-kit/p/IG0021167/

Forestry Suppliers

Source of water sampling equipment.

Available from: http://www.forestry-suppliers.com/

Appendix

Appendix A: SPARK Science Learning System™ Tech Tips

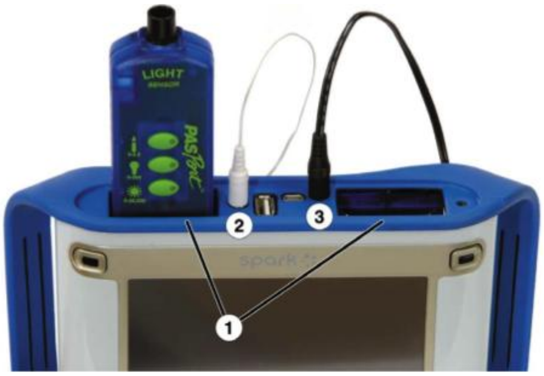
The number following the symbol "�" at the end of a procedure step refers to the numbered Tech Tip below which provides directions for operating the SPARK Science Learning System.

2.1 Connecting a sensor to the SPARK

- If you have a blue PASPORT sensor, plug it into either of the PASPORT ports on the SPARK.
- If you have a temperature probe (fast-response or stainless steel), plug it into the temperature port.
- If you have a voltage probe, plug it into the voltage port.
- If you have a black PASPORT colorimeter, water quality colorimeter or turbidity meter, you will need a sensor extension cable to connect the sensor to the SPARK.

Task Result: The SPARK detects the sensor and adds it to your experiment.

1. PASPORT ports. 2. Temperature port. 3. Voltage port.



2.2 Building an experiment

1. Connect a sensor to the SPARK. A live data screen will appear.
2. Touch **Build**.
3. Choose the parameter you want to observe in the menu to the left of the screen. Touch the name of that parameter.
4. Touch the **Digits Display** 1.23 button. Continue to add parameters in this manner until you have added all parameters relevant to the experiment.

   Note: Six digits-displays can be shown at once.
5. Touch **OK**. The display(s) will appear on a new page.

2.3 Putting the SPARK into manual sampling mode

1. Touch the **Sampling Options** button.

   Result: The Sampling Option screen opens.

2. Touch **Manual**.
3. Touch **OK** to close the Sampling Options screen.

Task result: The SPARK is now ready to record manually sampled data.

2.4 Recording a set of manually sampled data

If the SPARK is in manual sampling mode, complete these steps to record a data set:

1. Touch the **Start** button. ▶

   Alternatively, you can press one of the **Record** buttons instead of touching the on-screen **Start** button.

   Result: The SPARK creates a new data set. Live data appear in the data displays. The **Record** buttons start flashing to indicate that the SPARK is ready to be triggered.

2. When you are ready to trigger the recording of a data point, touch the **Keep** button ✓

   Alternatively, you can press one of the **Record** buttons instead of touching the on-screen **Keep** button.

   Record buttons

   Result: The SPARK records a single value from each measurement.

3. Repeat Step 2 as many times as necessary to record all of the data you want in the data set.

4. When the entire set has been recorded, touch the **Stop** button.

Result: The data set closes.

Note: If you accidentally stop the data collection early (by touching the **Stop** button instead of the **Keep** button), you will need to start over again from the beginning.

2.5 Calibrating the Conductivity Sensor or Water Quality Sensor

Depending upon your goals, the factory calibration of the Conductivity Sensor may be inadequate. Conductivity measurements are usually used to provide an estimate of total dissolved solids (TDS) or to ascertain whether an additional, more specific measurement of a particular ion is needed. If calibration is required, you will need a conductivity standard in the approximate range and at the same temperature as that of the unknown solution you will be testing.

Obtaining a conductivity standard

You can make your own conductivity standard solutions or you can purchase them from a supply company such as Hach Company or Cole-Parmer. For most purposes, a standard solution that you make will be sufficient. However, for high-accuracy measurements, use a commercially prepared and standardized solution appropriate to your measuring situation. See the “Preparation of standard solutions” table below.

Preparation of standard solutions

% weight (approx.)	Mass of NaCl (mg)	TDS (ppm or mg/L)	Conductivity (μS/cm) at 25°C
0.001	10	10	21.4
0.01	100	100	210
0.1	1000	1000	1,990
1.0	10,000	10,000	17,600
10.0	100,000	100,000	140,000

Put the mass of NaCl that will yield a conductivity value close to the range you will be measuring into a 1-L flask. Add 500 mL of distilled water and stir until dissolved. Then bring the volume to 1-L with distilled water.

Materials Required
Conductivity standard
Small beakers (3)
Wash bottle (1)
Distilled water

Calibrating the sensor

1. Place a sample of distilled water into a beaker and a sample of standard solution into another beaker.
2. Soak the conductivity electrode in the beaker of distilled water for 10 minutes.
3. Plug the Conductivity Sensor into the SPARK.
4. Touch the Conductivity reading on the live data page and touch **Show**.
5. Touch the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.

6. Touch **CALIBRATE SENSOR**.

   Result: The Calibrate Sensor: screen opens.

7. Touch the **Sensor** box and select the **Conductivity Sensor**.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor from the list.

8. Select **Conductivity (µS)** from the **Measurement**: box.
9. Use the **1 Point (Adjust Slope Only)** calibration. Touch **NEXT**.
10. Touch the **Standard Value** box.
11. Using the keypad, type in the standard reference value (µS/cm) that corresponds to that of the reference sample you are using.

    Note: Be sure your reference sample is at 25 degrees Celsius. If not, adjust the reference value by 2% for each 1-degree deviation.

12. Insert the Conductivity Probe into the standard solution, being sure to immerse the holes near the end of the probe.
13. Touch the **READ FROM SENSOR** button. The sensor reading will fill the **Sensor Value** box.
14. Touch **OK**. The display will return to the **Calibrate Sensor:** screen.
15. Touch **OK** to return to the **Experiment Tools** screen. Touch **OK** again to return to the experiment.

2.6 Calibrating the Dissolved Oxygen Sensor or Water Quality Sensor

Calibration of the Dissolved Oxygen (DO) Sensor must be performed at or near the temperature and barometric pressure of the natural water body being tested.

To find the most accurate reference value, you will need the barometric pressure at the test site. You can use a PASCO Barometer/Low Pressure Sensor to determine the local barometric pressure, or you could call the local weather station to find the barometric pressure, being sure to ask for the *station pressure*, not the *altimeter* or *sea-level pressure*.

You will also need to know the temperature of the water you will be testing. Then you can use an online calculator such as the one provided by Frostburg State University to calculate the 100% saturation value (in terms of mg/L DO). (Alternatively, you can create your own equations to make the calculation, using information from the Internet or from the PASPORT Dissolved Oxygen Sensor User Manual).

You may want to make a table of values that bracket the temperature you expect to find at the test site before heading into the field.

If you are measuring dissolved oxygen in saline water, consult the PASPORT Dissolved Oxygen Sensor User Manual. This manual contains detailed charts of correction factors for measurements taken at various levels of salinity.

1. Plug the Dissolved Oxygen Sensor into the data collection system.
2. Place approximately 5 mL of deionized water into the soaker bottle of the DO Sensor.
3. Touch the Dissolved Oxygen reading on the live data page and touch **Show**.
4. Touch the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.
5. Touch **CALIBRATE SENSOR**.

   Result: The **Calibrate Sensor**: screen opens.
6. Touch the **Sensor** box and select the **Dissolved Oxygen Sensor**.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor from the list.
7. In the Measurement box, select **Dissolved Oxygen (mg/L)**.
8. Use the **1 Point (Adjust Slope Only)** calibration. Touch **NEXT**.
9. Adjust the end of the probe such that it is just above the top of the water in the soaker bottle.
10. Touch the **Standard Value:** box. Using the keypad, enter the value from the solubility table for the temperature of the sample and barometric pressure.
11. Shake the bottle vigorously for a few seconds, and then gently shake off any water clinging to the end of the probe.
12. Touch the **READ FROM SENSOR** button.
13. Touch **OK**. The display will return to the Calibrate Sensor: screen.

14. Touch **OK** to return to the Experiment Tools screen. Touch **OK** again to return to the experiment.

2.7 Calibrating the pH Sensor or Water Quality Sensor

You will need buffer solutions of pH 4 and pH 10, and deionized water.

1. Plug the pH Sensor into the SPARK.
2. Touch the pH reading on the live data page and touch **Show**.
3. Touch the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.
4. Touch **CALIBRATE SENSOR**.

   Result: The Calibrate Sensor: screen opens.
5. Touch the **Sensor** box and select the **pH Sensor**.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor from the list.
6. Select **pH** from the **Measurement:** box.
7. Use the **2-point (Adjust Slope and Offset)** calibration. Touch **NEXT**.
8. Place the pH probe into the pH 4 buffer solution and wait for about 1 minute.
9. Under **Calibration Point 1,** touch **Standard Value** and enter **4** (the pH of the buffer solution).
10. Under **Calibration Point 1**, touch **Read From Sensor**.

    Result: The **Standard Value** and **Sensor Value** for Point 1 will be filled with the standard value (4) and the reading from the sensor, respectively.
11. Rinse the probe with deionized water.
12. Place the pH probe into the pH 10 buffer solution and wait for about 1 minute.
13. Under **Calibration Point 2**, touch **Standard Value** and enter **10** (the pH of the buffer solution).
14. Under **Calibration Point 2**, touch **Read From Sensor**.

    Result: The Standard Value and Sensor Value for Point 2 will be filled with the standard value (10) and the reading from the sensor, respectively.
15. Touch **OK**. The display will return to the Calibrate Sensor: screen.

16. Touch **OK** to return to the **Experiment Tools** screen. Touch **OK** again to return to the experiment.

2.8 Creating a Coordinate Marker using the GPS Sensor

1. Connect the GPS Sensor to the SPARK.
2. Touch **Build** on the live data page. Build an experiment that shows latitude and longitude on a line graph. Touch **OK**.
3. Set the SPARK to manual sampling mode. �(2.3)
4. Walk to the location you would like your coordinate marker to represent.
5. When you reach the location, record a data point and then end the data set.
6. Set the SPARK back to periodic sampling mode.
7. Begin collecting data. As you walk around, your second data set will be in the same scale and on the same graph as your original point. You can use the first data set as a reference and return to that point later from any location.

2.9 Calibrating the Salinity Sensor

You will need a 35 ppt sodium chloride solution and deionized water.

1. Connect the Salinity Sensor to the interface (SPARKlink or SPARK connected to the computer). Open SPARKvue.
2. Touch the salinity reading on the live data page and touch **Show**.
3. Touch the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.
4. Touch **CALIBRATE SENSOR**.

   Result: The Calibrate Sensor: screen opens.
5. Touch the **Sensor** box and select the Salinity Sensor.
6. Select **Salinity** from the **Measurement:** box.
7. Use the **1 point (Adjust Slope Only)** calibration. Touch **NEXT**.
8. Place the salinity probe into the 35 ppt sodium chloride solution and wait for about 1 minute.
9. Under **Calibration Point 1,** touch **Standard Value** and enter 35 ppt.
10. Under **Calibration Point 1**, touch **Read From Sensor**.

Result: The **Standard Value** and **Sensor Value** for Point 1 will be filled with the standard value (35 ppt) and the reading from the sensor, respectively.

11. Rinse the probe with deionized water.
12. Touch **OK**. The display will return to the Calibrate Sensor: screen.
13. Touch **OK** to return to the Experiment Tools screen. Touch **OK** again to return to the experiment.

Additional Sensor Calibrations (On-board Calibration)

3.1 Calibrating the Water Quality Colorimeter

1. Plug the PASPORT Water Quality Colorimeter into the data collection system �(2.1) and turn on the system.

   By default, two ammonia readings will become visible, one for low-range ammonia readings [Ammonia (L)] and one for high-range ammonia readings [Ammonia (H)].

2. Insert the ammonia ezSample snap vial calibration ampoule into the ampoule holder and cover it with the black cap. Be sure to cover the ampoule completely with the cap.

   Note: Handle the ampoule by the tip and wipe the outside with a clean, non-abrasive cleaning tissue.

   Important: To avoid errors when calibrating or measuring, do not use the Water Quality Colorimeter around bright lights.

3. Press the green Calibration button on the Water Quality Colorimeter.

   Result: The green light illuminates to indicate calibration is in progress.

4. Wait for the green light to turn off and then remove the calibration ampoule.
5. To check your calibration, reinsert the calibration ampoule and cover it with the black cap. The value should be approximately 0. If it is not, repeat the calibration with special care.

   Note: Because of the calculations involved in the ammonia calibration curve, the calibrated value may not be exactly 0, but may be approximately 0. However, the PASPORT Water Quality Colorimeter will nevertheless produce values within the accuracy range listed in the instruction sheet for the Ammonia ezSample Snap Vial.

Note: All calibration readings are stored in flash memory inside the PASPORT Water Quality Colorimeter. When you unplug the Water Quality Colorimeter and reconnect it, the Water Quality Colorimeter retains the last calibration reading. However, since it is possible the calibration button may be inadvertently pushed, it is important for the most reliable results to do the calibration procedure immediately before taking the measurement.

Note: A blinking red light on the Calibration button means either (a) the dark count is too high—stray light is entering the ampoule holder or (b) the sensor measurement is out of range. The blinking red light turns off when the reading is within the normal range.

3.2 Calibrating the Turbidity Sensor

1. Place a cuvette filled with distilled water in the turbidity sensor and close the lid.
2. Press the **Calibration** button on the sensor.

   Result: The green light illuminates to indicate calibration is in progress.

3. When the light starts blinking, replace the cuvette with the standard 100 NTU cuvette (included with the sensor) and close the lid.
4. Press the button again.

   Result: The green light illuminates to indicate calibration is in progress.

5. When the light turns off, the calibration process is complete

Appendix B: SPARKvue™ Tech Tips

The number following the symbol "�" at the end of a procedure step refers to the numbered Tech Tip below which provides directions for operating SPARKvue.

2.1 Adding a sensor to your SPARKvue experiment

1. Connect an interface such as a SPARKlink or a SPARK to your computer.
2. Connect the sensor to a port on the interface.
3. Open SPARKvue.

Task Result: SPARKvue detects the sensor.

2.2 Building an experiment

1. Connect a sensor to the interface (SPARKlink or SPARK connected to the computer). Open SPARKvue.
2. Click **Build**.
3. Choose the parameter you want to observe in the menu to the left of the screen. Click the name of that parameter.
4. Click the **Digits Display** 1.23 button. Continue to add parameters in this manner until you have added all parameters relevant to the experiment.

   Note: Six digits displays can be shown at once.

5. Click **OK**. The display(s) will appear on a new page.

2.3 Putting SPARKvue into manual sampling mode

1. Click the **Sampling Options** button.

   Result: The **Sampling Option** screen opens.

2. Click **Manual**.
3. Click **OK** to close the Sampling Options screen.

Task result: SPARKvue is now ready to record manually sampled data.

2.4 Recording a set of manually sampled data

If SPARKvue is in manual sampling mode, complete these steps to record a data set:

1. Click the **Start** button. ▶

   Result: SPARKvue creates a new data set. Live data appear in the data displays.

2. When you are ready to trigger the recording of a data point, click the **Keep** button. ✓

3. Repeat Step 2 as many times as necessary to record all of the data that you want in the data set.

4. When the entire set has been recorded, click the **Stop** button. ■

   Result: The data set closes.

   Note: If you accidentally stop the data collection early (by touching the **Stop** button instead of the **Keep** button), you will need to start over again from the beginning.

2.5 Calibrating the Conductivity Sensor or Water Quality Sensor

Depending upon your goals, the factory calibration of the Conductivity Sensor may be inadequate. Conductivity measurements are usually used to provide an estimate of total dissolved solids (TDS) or to ascertain whether an additional, more specific measurement of a particular ion is needed. If calibration is required, you will need a conductivity standard in the approximate range and at the same temperature as that of the unknown solution you will be testing.

Obtaining a conductivity standard

You can make your own conductivity standard solutions or you can purchase them from a supply company such as Hach Company or Cole-Parmer. For most purposes, a standard solution that you make will be sufficient. However, for high-accuracy measurements, use a commercially prepared and standardized solution appropriate to your measuring situation. See the “Preparation of standard solutions” table below.

Preparation of standard solutions

% weight (approx.)	Mass of NaCl (mg)	TDS (ppm or mg/L)	Conductivity (μS/cm) at 25°C
0.001	10	10	21.4
0.01	100	100	210
0.1	1000	1000	1,990
1.0	10,000	10,000	17,600
10.0	100,000	100,000	140,000

Put the mass of NaCl that will yield a conductivity value close to the range you will be measuring into a 1-L flask. Add 500 mL of distilled water and stir until dissolved. Then bring the volume to 1-L with distilled water.

Materials Required
Conductivity standard
Small beakers (3)
Wash bottle (1)
Distilled water

Calibrating the sensor

1. Place a sample of distilled water into a beaker and a sample of standard solution into another beaker.
2. Soak the conductivity electrode in the beaker of distilled water for 10 minutes.
3. Connect the Conductivity Sensor to the interface (SPARKlink or SPARK connected to the computer). Open SPARKvue.
4. Click the Conductivity reading on the live data page and click **Show**.
5. Click the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.
6. Click **CALIBRATE SENSOR**.

   Result: The **Calibrate Sensor**: screen opens.
7. Click the **Sensor** box and select the **Conductivity Sensor**.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor from the list.
8. Select **Conductivity (µS)** from the **Measurement:** box.
9. Use the **1 Point (Adjust Slope Only)** calibration. Click **NEXT**.
10. Click the **Standard Value:** box.
11. Using the keypad, type in the standard reference value (µS/cm) that corresponds to that of the reference sample you are using.

    Note: Be sure your reference sample is at 25 degrees Celsius. If not, adjust the reference value by 2% for each 1-degree deviation.
12. Insert the Conductivity Probe into the standard solution, being sure to immerse the holes near the end of the probe.
13. Click the **READ FROM SENSOR** button. The sensor reading will fill the **Sensor Value** box.
14. Click **OK**. The display will return to the **Calibrate Sensor:** screen.
15. Click **OK** to return to the **Experiment Tools** screen. Click **OK** again to return to the experiment.

2.6 Calibrating the Dissolved Oxygen Sensor or Water Quality Sensor

Calibration of the DO sensor must be performed at or near the temperature and barometric pressure of the natural water body being tested.

To find the most accurate reference value, you will need the barometric pressure at the test site. You can use a PASCO Barometer/Low Pressure Sensor to determine the local barometric pressure, or you could call the local weather station to find the barometric pressure, being sure to ask for the *station pressure*, not the *altimeter* or *sea-level pressure*.

You will also need to know the temperature of the water you will be testing. Then you can use an online calculator such as the one provided by Frostburg State University[1] to calculate the 100% saturation value (in terms of mg/L DO). (Alternatively, you can create your own equations to make the calculation, using information from the Internet or from the PASPORT Dissolved Oxygen Sensor User Manual)..

You may want to make a table of values that bracket the temperature you expect to find at the test site before heading into the field.

If you are measuring dissolved oxygen in saline water, consult the PASPORT Dissolved Oxygen Sensor User Manual. This manual contains detailed charts of correction factors for measurements taken at various levels of salinity.

1. Connect the Dissolved Oxygen Sensor to the interface (SPARKlink or SPARK connected to the computer). Open SPARKvue.
2. Place approximately 5 mL of deionized water into the soaker bottle of the DO Sensor.
3. Click the Dissolved Oxygen reading on the live data page and click **Show**.
4. Click the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.
5. Click **CALIBRATE SENSOR**.

   Result: The **Calibrate Sensor**: screen opens.
6. Touch the **Sensor** box and select the Dissolved Oxygen Sensor.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor from the list.
7. In the **Measurement** box, select **Dissolved Oxygen (mg/L)**.
8. Use the **1 Point (Adjust Slope Only)** calibration. Click **NEXT**.
9. Adjust the end of the probe such that it is just above the top of the water in the soaker bottle.
10. Click the **Standard Value:** box. Using the keypad, enter the value from the solubility table for the temperature of the sample and barometric pressure.

11. Shake the bottle vigorously for a few seconds, and then gently shake off any water clinging to the end of the probe.
12. Click the **READ FROM SENSOR** button.
13. Click **OK**. The display will return to the Calibrate Sensor: screen.
14. Click **OK** to return to the Experiment Tools screen. Click **OK** again to return to the experiment.

2.7 Calibrating the pH Sensor or Water Quality Sensor

You will need buffer solutions of pH 4 and pH 10, and deionized water.

1. Connect the pH Sensor to the interface (SPARKlink or SPARK connected to the computer). Open SPARKvue.
2. Click the pH reading on the live data page and click **Show**.
3. Click the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.

4. Click **CALIBRATE SENSOR**.

   Result: The Calibrate Sensor: screen opens.

5. Click the **Sensor** box and select the pH Sensor.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor from the list.

6. Select **pH** from the **Measurement:** box.
7. Use the **2 Point (Adjust Slope and Offset)** calibration. Click **NEXT**.
8. Place the pH probe into the pH 4 buffer solution and wait for about 1 minute.
9. Under Calibration Point 1, click **Standard Value** and enter **4** (the pH of the buffer solution).
10. Under Calibration Point 1, click **Read From Sensor**.

    Result: The Standard Value and Sensor Value for Point 1 will be filled with the standard value (4) and the reading from the sensor, respectively.

11. Rinse the probe with deionized water.
12. Place the pH probe into the pH 10 buffer solution and wait for about 1 minute.
13. Under **Calibration Point 2**, click **Standard Value** and enter **10** (the pH of the buffer solution).
14. Under **Calibration Point 2**, click **Read From Sensor**.

Result: The Standard Value and Sensor Value for Point 2 will be will be filled with the standard value (10) and the reading from the sensor, respectively.

15. Click **OK**. The display will return to the Calibrate Sensor: screen.
16. Click **OK** to return to the **Experiment Tools** screen. Click **OK** again to return to the experiment.

2.9 Calibrating the Salinity Sensor

You will need a 35 ppt sodium chloride solution and deionized water.

1. Connect the Salinity Sensor to the interface (SPARKlink or SPARK connected to the computer). Open SPARKvue.
2. Click the salinity reading on the live data page and click **Show**.
3. Click the **Experiment Tools** button.

   Result: The Experiment Tools screen opens.

4. Click **CALIBRATE SENSOR**.

   Result: The Calibrate Sensor: screen opens.

5. Click the **Sensor** box and select the Salinity Sensor.
6. Select Salinity from the **Measurement:** box.
7. Use the **1 point (Adjust Slope Only)** calibration. Click **NEXT**.
8. Place the salinity probe into the 35 ppt sodium chloride solution and wait for about 1 minute.
9. Under **Calibration Point 1,** click **Standard Value** and enter 35 ppt.
10. Under **Calibration Point 1**, click **Read From Sensor**.

    Result: The **Standard Value** and **Sensor Value** for Point 1 will be filled with the standard value (35 ppt) and the reading from the sensor, respectively.

11. Rinse the probe with deionized water.
12. Click **OK**. The display will return to the Calibrate Sensor: screen.
13. Click **OK** to return to the Experiment Tools screen. Click **OK** again to return to the experiment.

Additional Sensor Calibrations (On-board Calibration)

3.1 Calibrating the Water Quality Colorimeter

1. Plug the PASPORT Water Quality Colorimeter into the data collection system �(2.1) and turn on the system.

   By default, two ammonia readings will become visible, one for low-range ammonia readings [Ammonia (L)] and one for high-range ammonia readings [Ammonia (H)].

2. Insert the ammonia ezSample snap vial calibration ampoule into the ampoule holder and cover it with the black cap. Be sure to cover the ampoule completely with the cap.

   Note: Handle the ampoule by the tip and wipe the outside with a clean, non-abrasive cleaning tissue.

   Important: To avoid errors when calibrating or measuring, do not use the Water Quality Colorimeter around bright lights.

3. Press the green Calibration button on the Water Quality Colorimeter.

   Result: The green light illuminates to indicate calibration is in progress.

4. Wait for the green light to turn off and then remove the calibration ampoule.

5. To check your calibration, reinsert the calibration ampoule and cover it with the black cap. The value should be approximately 0. If it is not, repeat the calibration with special care.

   Note: Because of the calculations involved in the ammonia calibration curve, the calibrated value may not be exactly 0, but may be approximately 0. However, the PASPORT Water Quality Colorimeter will nevertheless produce values within the accuracy range listed in the instruction sheet for the Ammonia ezSample Snap Vial.

   Note: All calibration readings are stored in flash memory inside the PASPORT Water Quality Colorimeter. When you unplug the Water Quality Colorimeter and reconnect it, the Water Quality Colorimeter retains the last calibration reading. However, since it is possible the calibration button may be inadvertently pushed, it is important for the most reliable results to do the calibration procedure immediately before taking the measurement.

Note: A blinking red light on the Calibration button means either (a) the dark count is too high—stray light is entering the ampoule holder or (b) the sensor measurement is out of range. The blinking red light turns off when the reading is within the normal range.

3.2 Calibrating the Turbidity Sensor

1. Place a cuvette filled with distilled water in the turbidity sensor and close the lid.
2. Press the **Calibration** button on the sensor.

   Result: The green light illuminates to indicate calibration is in progress.

3. When the light starts blinking, replace the cuvette with the standard 100 NTU cuvette (included with the sensor) and close the lid.
4. Press the button again.

   Result: The green light illuminates to indicate calibration is in progress.

5. When the light turns off, the calibration process is complete.

Appendix C: Xplorer GLX™ Tech Tips

The number following the symbol "�" at the end of a procedure step refers to the numbered Tech Tip below which provides directions for operating the Xplorer GLX.

2.1

Connecting a sensor to the GLX

- If you have a PASPORT sensor, plug it into one of the PASPORT ports on the GLX.
- If you have a temperature probe (fast-response or stainless steel) plug it into one of the temperature ports on the side of the GLX.

  Note: The temperature ports are labeled with the icons 🌡1 and 🌡2.

Task Result: The Xplorer GLX detects the sensor and adds it to your experiment.

2.2

Building an Experiment: Showing Data Displays

1. Connect a sensor to the Xplorer GLX.

   Result: A default measurement from the sensor will appear in the Digits Display.

2. In the Digits Display, the number of sensor measurements can be changed to display up to 8 measurements.
3. To change any measurements on the display, press ✓. The field for the first measurement will be highlighted.
4. Using the arrow keys on the Xplorer GLX, navigate to the field whose measurement will be changed.
5. Press ✓. A menu will appear containing all of the measurements for all sensors connected to the Xplorer GLX.
6. Using the arrow keys on the Xplorer GLX, navigate to the measurement that will be shown and press ✓.

2.3

Putting the Xplorer GLX into manual sampling mode

1. From the Home screen, press F4 to open the **Sensors** screen.
2. Press F1 to open the **Mode** screen.
3. Using the down arrow, highlight “Manual” and press ✓.

Task Result: The Xplorer GLX is now ready to record manually sampled data.

2.4 Recording a set of manually sampled data

If the Xplorer GLX is in manual sampling mode, complete these steps to record a data set:

1. From the Home screen, use the arrow keys and press ✓ to open the digits display.
2. Press ▶. The Xplorer GLX is now ready to collect data.
3. When the reading you are taking has stabilized, press ⚑.

   Result: The Xplorer GLX records a data point.

   Repeat as necessary until all data points have been collected.
4. Press ▶ to stop data collection.

2.5 Calibrating the Conductivity Sensor or Water Quality Sensor

Depending upon your goals, the factory calibration of the Conductivity Sensor may be inadequate. Conductivity measurements are usually used to provide an estimate of total dissolved solids (TDS) or to ascertain whether an additional, more specific measurement of a particular ion is needed. If calibration is required, you will need a conductivity standard in the approximate range and at the same temperature as that of the unknown solution you will be testing.

Obtaining a conductivity standard

You can make your own conductivity standard solutions or you can purchase them from a supply company such as Hach Company or Cole-Parmer. For most purposes, a standard solution that you make will be sufficient. However, for high-accuracy measurements, use a commercially prepared and standardized solution appropriate to your measuring situation. See the “Preparation of standard solutions” table below.

Preparation of standard solutions

% weight (approx.)	Mass of NaCl (mg)	TDS (ppm or mg/L)	Conductivity (μS/cm) at 25°C
0.001	10	10	21.4
0.01	100	100	210
0.1	1000	1000	1,990
1.0	10,000	10,000	17,600
10.0	100,000	100,000	140,000

Put the mass of NaCl that will yield a conductivity value close to the range you will be measuring into a 1-L flask. Add 500 mL of distilled water and stir until dissolved. Then bring the volume to 1-L with distilled water.

Materials Required
Conductivity standard
Small beakers (3)
Wash bottle (1)
Distilled water

Calibrating the sensor

1. Place a sample of distilled water into a beaker and a sample of standard solution into another beaker.
2. Soak the conductivity electrode in the beaker of distilled water for 10 minutes.
3. Plug the Conductivity Sensor into the Xplorer GLX.
4. From the Home screen, press F4 to open the **Sensors** screen.

   Note: If you have more than one sensor connected, use the arrow keys to select the appropriate sensor across the top of the GLX screen.

5. Press F4 again to open the **Sensors** menu.
6. Using the arrow keys, select **Calibrate** and press ✓.
7. The first box should read **Conductivity Sensor**. If it does not, press the ✓ button until it appears.
8. The second box should read **Conductivity (µS/cm)**. If it does not, press the ✓ button until it appears.
9. Using the arrow keys, highlight **Pt 2 (µS/cm)** and press ✓. Input the standard reference value using the keypad.

   Note: Be sure that your reference sample is at 25 degrees Celsius. If not, adjust the reference value by 2% for each 1-degree deviation.

10. Insert the Conductivity Probe into the standard solution, being sure to immerse the holes near the end of the probe.
11. Press F4 and **Read Pt 2** when the reading stabilizes.
12. Press **OK** to exit the Calibration screen. Press ⌂ to return to the Home screen.

2.6 Calibrating the Dissolved Oxygen Sensor or Water Quality Sensor

Calibration of the DO sensor must be performed at or near the temperature and barometric pressure of the natural water body being tested.

To find the most accurate reference value, you will need the barometric pressure at the test site. You can use a PASCO Barometer/Low Pressure Sensor to determine the local barometric pressure, or you could call the local weather station to find the barometric pressure, being sure to ask for the *station pressure*, not the *altimeter* or *sea-level pressure*.

You will also need to know the temperature of the water you will be testing. Then you can use an online calculator such as the one provided by Frostburg State University to calculate the 100% saturation value (in terms of mg/L DO). (Alternatively, you can create your own equations to make the calculation, using information from the Internet or from the PASPORT Dissolved Oxygen Sensor User Manual).

You may want to make a table of values that bracket the temperature you expect to find at the test site before heading into the field.

If you are measuring dissolved oxygen in saline water, consult the PASPORT Dissolved Oxygen Sensor User Manual. This manual contains detailed charts of correction factors for measurements taken at various levels of salinity.

1. Plug the Dissolved Oxygen Sensor into the Xplorer GLX.
2. Place approximately 5 mL of deionized water into the soaker bottle of the DO Sensor.
3. From the Home screen, press F4 to open the Sensors screen.

   Note: If you have more than one sensor connected, use the arrow keys to select the appropriate sensor across the top of the GLX screen.

4. Press F4 again to open the **Sensors** menu.
5. Using the arrow keys, select **Calibrate** and press ✓
6. The first box should read **Dissolved Oxygen Sensor**. If it does not, press the ✓ button until it appears.
7. The second box should read **Dissolved Oxygen (mg/L)**. If it does not, press the ✓ button until it appears.
8. Using the arrow keys, select the **Pt 2 Standard Value** box and press ✓.
9. Using the keypad, enter the value from the solubility table for the temperature of the sample and barometric pressure.
10. Adjust the end of the probe such that it is just above the top of the water in the soaker bottle.
11. Shake the bottle vigorously for a few seconds, and then gently shake off any water clinging to the end of the probe.
12. Press F4 to read **Pt 2**.
13. Press **OK** to exit the Calibration screen. Press ⌂ to return to the Home screen.

2.7 Calibrating the pH Sensor or Water Quality Sensor

You will need buffer solutions of pH 4 and pH 10, and deionized water.

1. Plug the pH Sensor into the Xplorer GLX.
2. From the Home screen, press F4 to open the **Sensors** screen.

   Note: If you have more than one sensor connected, use the arrow keys to select the appropriate sensor across the top of the GLX screen.

3. Press F4 again to open the **Sensors** menu.
4. Using the arrow keys, select **Calibrate** and press ✓
5. The first box should read **pH Sensor**. If it does not, press the ✓ button until it appears.
6. The **Calibration Type** box should read **2 Point**. If it does not, highlight this box, press ✓, and select **2 Point**.
7. Place the pH probe into the pH 4 buffer solution and wait for about 1 minute.
8. Select **Pt 1,** press ✓, and enter **4** (the pH of the buffer solution).
9. Press F3 to read **Pt 1**.

   Result: The Standard Value and Sensor Value for Point 1 will be filled with the standard value (4) and the reading from the sensor, respectively.

10. Rinse the probe with deionized water.
11. Using the arrow keys, scroll down and select **Pt 2.** Press ✓, and enter **10** (the pH of the buffer solution).
12. Press F4 to read **Pt 2**.

    Result: The Standard Value and Sensor Value for Point 2 will be filled with the standard value (10) and the reading from the sensor, respectively.

13. Press **OK** to exit the Calibration screen. Press ⌂ to return to the Home screen.

2.8 Creating a Coordinate Marker using the GPS Sensor

1. Connect the GPS Sensor to the Xplorer GLX.
2. From the Home screen, press F1 to open the **Graph**.
3. Press F3 to open the **Tools** menu.
4. Using the arrows, select the **Coordinate Marker** tool and press ✓.

5. A small window will appear. Enter the latitude and longitude into the boxes.

   Note: Latitude and longitude must be in decimal degrees and must be represented as negative or positive as opposed to a directional representation (north, south, east and west).

6. A hash mark will appear on the graph. This is your coordinate marker.
7. Begin collecting data. As you walk around, your data set will be in the same scale and on the same graph as your coordinate marker. You can use this as a reference to return to that point later from any location.

2.9 Calibrating the Salinity Sensor

You will need a 35 ppt sodium chloride solution and deionized water.

1. Connect the Salinity Sensor to the Xplorer GLX
2. From the Home screen, press F4 to open the Sensors screen.

   Note: If you have more than one sensor connected, use the arrow keys to select the appropriate sensor across the top of the GLX screen.

3. Press F4 again to open the **Sensors** menu.
4. Using the arrow keys, select **Calibrate** and press ✓ •EE
5. The first box should read **Salinity Sensor**. If it does not, press the ✓ button until it appears.
6. The second box should read **Salinity (ppt)**. If it does not, press the ✓ button until it appears.
7. Using the arrow keys, highlight **Pt 2 (ppt)** and press ✓. Input the standard reference value using the keypad.
8. Insert the Salinity Probe into the standard solution.
9. Press F4 and read **Pt 2** when the reading stabilizes.
10. Press **OK** to exit the Calibration screen. Press ⌂ to return to the Home screen.

Additional Sensor Calibrations (On-board Calibration)

3.1 Calibrating the Water Quality Colorimeter

1. Plug the PASPORT Water Quality Colorimeter into the data collection system �(2.1) and turn on the system.

   By default, two ammonia readings will become visible, one for low-range ammonia readings [Ammonia (L)] and one for high-range ammonia readings [Ammonia (H)].

2. Insert the ammonia ezSample snap vial calibration ampoule into the ampoule holder and cover it with the black cap. Be sure to cover the ampoule completely with the cap.

   Note: Handle the ampoule by the tip and wipe the outside with a clean, non-abrasive cleaning tissue.

   Important: To avoid errors when calibrating or measuring, do not use the Water Quality Colorimeter around bright lights.

3. Press the green Calibration button on the Water Quality Colorimeter.

   Result: The green light illuminates to indicate calibration is in progress.

4. Wait for the green light to turn off and then remove the calibration ampoule.

5. To check your calibration, reinsert the calibration ampoule and cover it with the black cap. The value should be approximately 0. If it is not, repeat the calibration with special care.

   Note: Because of the calculations involved in the ammonia calibration curve, the calibrated value may not be exactly 0, but may be approximately 0. However, the PASPORT Water Quality Colorimeter will nevertheless produce values within the accuracy range listed in the instruction sheet for the Ammonia ezSample Snap Vial.

   Note: All calibration readings are stored in flash memory inside the PASPORT Water Quality Colorimeter. When you unplug the Water Quality Colorimeter and reconnect it, the Water Quality Colorimeter retains the last calibration reading. However, since it is possible the calibration button may be inadvertently pushed, it is important for the most reliable results to do the calibration procedure immediately before taking the measurement.

Note: A blinking red light on the Calibration button means either (a) the dark count is too high—stray light is entering the ampoule holder or (b) the sensor measurement is out of range. The blinking red light turns off when the reading is within the normal range.

3.2 Calibrating the Turbidity Sensor

1. Place a cuvette filled with distilled water in the turbidity sensor and close the lid.
2. Press the **Calibration** button on the sensor.

   Result: The green light illuminates to indicate calibration is in progress.

3. When the light starts blinking, replace the cuvette with the standard 100 NTU cuvette (included with the sensor) and close the lid.
4. Press the button again.

   Result: The green light illuminates to indicate calibration is in progress.

5. When the light turns off, the calibration process is complete.

Appendix D: DataStudio® Tech Tips

The number following the symbol "�" at the end of a procedure step refers to the numbered Tech Tip below which provides directions for operating DataStudio.

2.1 Adding a sensor to your DataStudio experiment

1. If it is not already connected, connect a PASPORT interface (such as a USB Link, PowerLink, or Xplorer GLX) to your computer. Some interfaces also need to be connected to the AC power or turned on.
2. Connect the sensor to the PASPORT interface. You can plug the sensor directly into the interface or use an extension cable between the sensor and interface.

Task Result: DataStudio detects the sensor and adds it to your experiment.

2.2 Showing Data Displays

1. The sensors available for the experiment you have built will be shown to the left of the screen in the Data menu.
2. Click and drag the sensor measurement (pH, for example) to the type of display you want (graph, for example). When you release the measurement, a window will appear, displaying that measurement in the specified display type.

2.3 Putting DataStudio into manual sampling mode

1. Click **SETUP**.
2. In the experiment setup window, click **Sampling Options**.
3. Select **Keep data values only when commanded.**
4. If you are going to enter manual data, such as keyboard data, type the name of the data you will be manually entering in the **Name** box.
5. Click **OK**.

   Result: DataStudio is now ready to record manually sampled data.

2.4 Recording a set of manually sampled data

If DataStudio is in manual sampling mode, complete these steps to record a data set:

1. Click **Start**.

   Result: A new data set is started and the **Start** button is replaced by the **Keep** button.

2. When you are ready to record a point, click **Keep**.

3. If DataStudio prompts you to enter keyboard data, you must enter a number, not text.

Repeat as necessary until all data points have been collected.

4. When you have recorded the *entire* data set, click the **Stop** button.

2.5 Calibrating the Conductivity Sensor or Water Quality Sensor

Depending upon your goals, the factory calibration of the Conductivity Sensor may be inadequate. Conductivity measurements are usually used to provide an estimate of total dissolved solids (TDS) or to ascertain whether an additional, more specific measurement of a particular ion is needed. If calibration is required, you will need a conductivity standard in the approximate range and at the same temperature as that of the unknown solution you will be testing.

Obtaining a conductivity standard

You can make your own conductivity standard solutions or you can purchase them from a supply company such as Hach Company or Cole-Parmer. For most purposes, a standard solution that you make will be sufficient. However, for high-accuracy measurements, use a commercially prepared and standardized solution appropriate to your measuring situation. See the “Preparation of standard solutions” table below.

Preparation of standard solutions

% weight (approx.)	Mass of NaCl (mg)	TDS (ppm or mg/L)	Conductivity (µS/cm) at 25°C
0.001	10	10	21.4
0.01	100	100	210
0.1	1000	1000	1,990
1.0	10,000	10,000	17,600
10.0	100,000	100,000	140,000

Put the mass of NaCl that will yield a conductivity value close to the range you will be measuring into a 1-L flask. Add 500 mL of distilled water and stir until dissolved. Then bring the volume to 1-L with distilled water.

Materials Required
Conductivity standard
Small beakers (3)
Wash bottle (1)
Distilled water

Calibrating the sensor

1. Place a sample of distilled water into a beaker and a sample of standard solution into another beaker.
2. Soak the conductivity electrode in the beaker of distilled water for 10 minutes.

3. Plug the Conductivity Sensor into a PowerLink or USBLink and connect the interface to the computer.
4. Click **Setup**.
5. In the Experiment Setup window, click **Calibrate Sensors**.
6. The first box should read Conductivity Sensor. If it does not, click the down arrow and select **Conductivity Sensor**.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor in the first box.

7. The second box should read Conductivity (µS/cm). If it does not, click the down arrow and select Conductivity (µS/cm).
8. Under Calibration Type, click the radio button next to 2 Point (Adjust Slope and Offset).
9. Input the standard reference value for Calibration Point 2 using the keyboard.

   Note: Be sure your reference sample is at 25 degrees Celsius. If not, adjust the reference value by 2% for each 1-degree deviation.

10. Insert the Conductivity Probe into the standard solution, being sure to immerse the holes near the end of the probe.
11. When the reading stabilizes, click **Read From Sensor**.
12. Click **OK**.

2.6 Calibrating the Dissolved Oxygen Sensor or Water Quality Sensor

Calibration of the DO sensor must be performed at or near the temperature and barometric pressure of the natural water body being tested.

To find the most accurate reference value, you will need the barometric pressure at the test site. You can use a PASCO Barometer/Low Pressure Sensor to determine the local barometric pressure, or you could call the local weather station to find the barometric pressure, being sure to ask for the *station pressure*, not the *altimeter* or *sea-level pressure*.

You will also need to know the temperature of the water you will be testing. Then you can use an online calculator such as the one provided by Frostburg State University to calculate the 100% saturation value (in terms of mg/L DO). (Alternatively, you can create your own equations to make the calculation, using information from the Internet or from the PASPORT Dissolved Oxygen Sensor User Manual).

You may want to make a table of values that bracket the temperature you expect to find at the test site before heading into the field.

If you are measuring dissolved oxygen in saline water, consult the PASPORT Dissolved Oxygen Sensor User Manual. This manual contains detailed charts of correction factors for measurements taken at various levels of salinity.

1. Plug the Dissolved Oxygen Sensor into a PowerLink or USBLink and connect the interface to the computer.
2. Place approximately 5 mL of deionized water into the soaker bottle of the DO Sensor.
3. Click **Setup**.
4. In the experiment setup tab, click **Calibrate Sensors**.
5. The first box should read **Dissolved Oxygen Sensor**. If it does not, click the down arrow and select **Dissolved Oxygen Sensor**.

   Note: If you are using the Water Quality Sensor, select Water Quality Sensor in the first box.

6. The second box should read **Dissolved Oxygen (mg/L)**. If it does not, click the down arrow and select **Dissolved Oxygen Sensor (mg/L)**.
7. Under **Calibration Type**, click the radio button next to **2 Point (Adjust Slope and Offset)**.
8. Using the keyboard, enter the standard value from the solubility table for the temperature of the sample and barometric pressure into **Calibration Point 2**.
9. Adjust the end of the probe such that it is just above the top of the water in the soaker bottle.
10. Shake the bottle vigorously for a few seconds, and then gently shake off any water clinging to the end of the probe.
11. Click **Read From Sensor**.
12. Click **OK**.

2.7 Calibrating the pH Sensor or Water Quality Sensor

You will need buffer solutions of pH 4 and pH 10, and deionized water.

1. Plug the pH Sensor into a PowerLink or USBLink and connect the interface to the computer.
2. Click **Setup**.
3. Click the **Calibrate Sensors** tab.
4. The first box should read **pH Sensor**. If it does not, click the down arrow and select **pH Sensor**.

   Note: If you are using the Water Quality Sensor, select **Water Quality Sensor** in the first box.

5. The second box should read **pH**. If it does not, click the down arrow and select **pH**.
6. Under **Calibration Type**, click the radio button next to **2 Point (Adjust Slope and Offset)**.

7. Place the pH probe into the pH 4 buffer and wait for the reading to stabilize. The reading is indicated under **Present Sensor Measurement**.
8. Enter the pH value of the buffer (in this case, **4**) into the **Standard Value** box for **Calibration Point 1**. Then click **Read From Sensor**.
9. Rinse the probe with deionized water and place it into the pH 10 buffer solution. Wait for the reading to stabilize.
10. Enter the pH value of the buffer (in this case, **10**) into the **Standard Value** box for **Calibration Point 2**. Then click **Read From Sensor**.
11. Click **OK**.

2.9 Calibrating the Salinity Sensor

You will need a 35 ppt sodium chloride solution and deionized water.

1. Plug the Salinity Sensor into a PowerLink or USBLink and connect the interface to the computer.
2. Click **Setup**.
3. Click the **Calibrate Sensors** tab.
4. The first box should read Salinity Sensor. If it does not, click the down arrow and select **Salinity Sensor**.
5. The second box should read Salinity (ppt). If it does not, click the down arrow and select **Salinity** (ppt).
6. Under Calibration Type, click the radio button next to **1 Point (Adjust Slope Only)**.
7. Place the Salinity probe into the standard solution and wait for the reading to stabilize. The reading is indicated under **Present Sensor Measurement**.
8. Enter the salinity concentration of the standard into the **Standard Value** box for Calibration Point 1. Then click **Read From Sensor**.

   Result: The **Standard Value** and **Sensor Value** for Point 1 will be filled with the standard value (35 ppt) and the reading from the sensor, respectively.

11. Rinse the probe with deionized water.
12. Click **OK**. The display will return to the Experiment Setup screen.
13. Close the window to return to the experiment.

Additional Sensor Calibrations (On-board Calibration)

3.1 Calibrating the Water Quality Colorimeter

1. Plug the PASPORT Water Quality Colorimeter into the data collection system �$^{(2.1)}$ and turn on the system.

   By default, two ammonia readings will become visible, one for low-range ammonia readings [Ammonia (L)] and one for high-range ammonia readings [Ammonia (H)].

2. Insert the ammonia ezSample snap vial calibration ampoule into the ampoule holder and cover it with the black cap. Be sure to cover the ampoule completely with the cap.

   Note: Handle the ampoule by the tip and wipe the outside with a clean, non-abrasive cleaning tissue.

   Important: To avoid errors when calibrating or measuring, do not use the Water Quality Colorimeter around bright lights.

3. Press the green Calibration button on the Water Quality Colorimeter.

   Result: The green light illuminates to indicate calibration is in progress.

4. Wait for the green light to turn off and then remove the calibration ampoule.

5. To check your calibration, reinsert the calibration ampoule and cover it with the black cap. The value should be approximately 0. If it is not, repeat the calibration with special care.

   Note: Because of the calculations involved in the ammonia calibration curve, the calibrated value may not be exactly 0, but may be approximately 0. However, the PASPORT Water Quality Colorimeter will nevertheless produce values within the accuracy range listed in the instruction sheet for the Ammonia ezSample Snap Vial.

   Note: All calibration readings are stored in flash memory inside the PASPORT Water Quality Colorimeter. When you unplug the Water Quality Colorimeter and reconnect it, the Water Quality Colorimeter retains the last calibration reading. However, since it is possible the calibration button may be inadvertently pushed, it is important for the most reliable results to do the calibration procedure immediately before taking the measurement.

Note: A blinking red light on the Calibration button means either (a) the dark count is too high—stray light is entering the ampoule holder or (b) the sensor measurement is out of range. The blinking red light turns off when the reading is within the normal range.

3.2 Calibrating the Turbidity Sensor

1. Place a cuvette filled with distilled water in the turbidity sensor and close the lid.
2. Press the **Calibration** button on the sensor.

   Result: The green light illuminates to indicate calibration is in progress.
3. When the light starts blinking, replace the cuvette with the standard 100 NTU cuvette (included with the sensor) and close the lid.
4. Press the button again.

   Result: The green light illuminates to indicate calibration is in progress.
5. When the light turns off, the calibration process is complete.