Water Chemistry
Field Study Manual

Portions adapted from: *U.S. Geological Survey, variously dated, National field manual for the collection of water-quality data: U.S. Geological Survey Techniques of Water-Resources Investigations, book 9, chaps. A1-A9, available online at http://pubs.water.usgs.gov/twri9A.*

Water Chemistry Parameters

Carbonate Hardness–a measurement of mineral content in water, specifically calcium and magnesium. Calcium and magnesium occur in natural waters as calcium carbonate ($CaCO_3$) and magnesium carbonate ($MgCO_3$). Carbonate hardness specifically measures the level of carbonate (CO_3^{2-}) and bicarbonate (HCO_3^-) ions. Sometimes, alkalinity is measured rather than total hardness; alkalinity is only a measure of calcium carbonate. Rocks and minerals are natural sources of hardness. Sewage outflow can also contribute carbonate minerals. Carbonate minerals play a role in regulating pH and can help remove dissolved toxic metals from the water. Although hardness is not directly linked to effects on aquatic organisms, carbonate minerals play a role in regulating pH, which has direct effects on the health of organisms.

pH–a measure of the concentration of hydrogen ions (H^+) in water. pH is measured on a logarithmic scale, where $pH = -\log [H^+]$. The pH scale ranges from 0 to 14, with pH values of <7 being acidic, =7 being neutral, and >7 being basic. pH is affected by the amount of carbon dioxide dissolved in the water, the substrates in the body of water, the geology and soils of the watershed and drainage from the watershed, and air pollution. Air pollution can cause acid rain, which can decrease the pH of the body of water into which the rain falls. Most aquatic organisms can survive a range of pH 6-9. pH values outside of this range are harmful, especially to juvenile organisms.

Phosphate (PO_4^{3-})–a nutrient required by all organisms. Phosphate occurs naturally in rocks, but can be introduced to the environment through runoff of cleaning products, fertilizers, or animal waste. Phosphate mining, industrial discharge, drinking water treatment, and land-use change can all impact the amount of phosphate that enters bodies of water. Excess phosphate can lead to a process called eutrophication, in which nutrient inputs cause small planktonic photosynthesizers to grow and reproduce very rapidly in a process called "blooming." Blooming is not sustainable for extended periods of time. The blooming of these small organisms leads to decreases in oxygen during the night and other periods of low light. These blooms also lead to increases in organisms that consume the bloomers and subsequent increases in decomposers, which eat the consumers and the bloomers as they both die off). This cycle of rapid growth, consumption, death, and decomposition often results in periods with DO concentrations below levels needed to sustain resident organisms. Large amounts of biotoxins are also often produced by blooming plankton and decomposers.

Turbidity–a measure of how clear or cloudy the water is. Turbidity can be caused by biogenic and abiogenic particles in the water, such as plankton, microscopic organisms, organic debris, sand, or silt. Biological factors, such as decay of plants and animals, benthic organisms that stir up sand and sediment, and algal blooms can increase turbidity. Water movement, such as flow, waves, runoff, or flooding, can stir up or carry particles. Wastewater carries many types of suspended solids and can increase turbidity. Particles associated with high turbidity can block light from reaching plants, slowing down photosynthesis, which can impact dissolved oxygen levels. If the water is very cloudy, aquatic organisms may be less able to see and catch food. Particles can clog fish gills and smother other organisms. Suspended particles also provide a surface for bacteria and pollutants to attach to.

Carbon Species

Carbonate (CO_3^{2-})–relates to water hardness, pH, and carbon dioxide levels. Carbonate occurs naturally in bodies of water and generally comes from dissolved rocks and minerals. Sewage outflow is a non-natural source of carbonates. Some organisms need carbonate to build calcium carbonate shells or skeletal structures. Depending on the method used, carbonate and pH can be used to determine carbon dioxide levels, rather than measuring carbon dioxide directly.

Carbon dioxide (CO_2)–when dissolved in water, reacts to form carbonic acid, which can decrease the pH of the water. Carbon dioxide dissolves in water through contact with the atmosphere and through respiration. While carbon dioxide is part of the natural carbon cycle, burning of fossil fuels and land use change can increase the concentration of carbon dioxide in the atmosphere, which in turn increases the concentration of carbon dioxide in the ocean. Changes in carbon dioxide and pH can affect organisms and other chemical process. In the ocean, changes in carbon dioxide are generally buffered, or kept relatively constant, by the presence of carbonate and bicarbonate ions. However, increased carbon dioxide concentration can lead to the decrease of availability of carbonate to organisms that build calcium carbonate structures.

Nitrogen Species

Ammonia/Ammonium–Ammonium (NH_4^+) is the ionized form of ammonia (NH_3). The total amount of ammonia in water is the sum of ammonia and ammonium. Ammonia is the least stable form of nitrogen in water and is easily converted to nitrate or nitrogen gas. The form of ammonia depends on the pH and temperature of the water. Sources of ammonia include sewage pollution and industrial or pesticide runoff. High concentrations of ammonia can affect the hatching and growth rates of fish, and damage gills, liver, and kidneys. Because ammonia can be converted to nitrate, excess ammonia results in many of the same effects on natural systems as excess nitrate.

Nitrate (NO_3^-)–a nutrient that is essential for the growth of photosynthetic organisms, including phytoplankton and algae. Nitrate occurs in water as part of the natural nitrogen cycle, but excess nitrate can be introduced into the environment by agricultural fertilizers, sewage discharge, or runoff from livestock areas. Nitrates can also be introduced into the air by fossil fuel burning, leading to acid rain, which can affect nitrate levels in bodies of water. Excessive nitrate concentration can cause problems for marine life. In fish, excess nitrate can cause “brown blood disease,” in which nitrate reacts with hemoglobin and prevents the blood from carrying oxygen. If nitrate levels are high, algae and phytoplankton can grow in large quantities, in events called blooms. During and after blooms, as plants die, bacteria decompose them, a process that uses up the oxygen in the water, called eutrophication. If oxygen levels drop too low, many aquatic organisms will not have enough oxygen to live, resulting in die offs and events called fish kills, where large numbers of fish perish.

Nitrite (NO_2^-)–a form of nitrogen that is present in the natural nitrogen cycle. It is generally very short-lived, because it is converted to nitrate by bacteria. Nitrite has similar sources and effects as nitrate.

Hawai‘i Water Quality Standards

Each state sets its own water quality standards for natural bodies of water. These are the water quality standards set by the State of Hawai‘i Department of Health. The dry season is May through October; the wet season is November through April.

	Coastal Wetlands	**Open Coastal Water**
	Average Limit	
Nitrate + Nitrite	70 ug/L wet season 30 ug/L dry season	5.00 ug/L wet season 3.50 ug/L dry season
Ammonia	---	3.50 ug/L wet season 2.00 ug/L dry season
Turbidity	5.0 NTU wet season 2.0 NTU dry season	0.50 NTU wet season 0.20 NTU dry season
pH	5.5<x<8.0	8.1 +/- 0.5

http://gen.doh.hawaii.gov/sites/har/AdmRules1/11-54.pdf

Useful websites

http://ga.water.usgs.gov/edu/waterproperties.html

http://www.h2ou.com/h2wtrqual.htm

http://bcn.boulder.co.us/basin/natural/wqterms.html

http://coastalscience.noaa.gov/news/coastal-pollution/nutrient-pollution-of-coastal-waters-too-much-of-a-good-thing/

http://water.epa.gov/learn/resources/measure.cfm

http://ww2.surfrider.org/waterquality.asp

http://www.beachapedia.org/State_of_the_Beach/State_Reports

http://coralreefecology.ucsd.edu/files/2010/09/Dailer_et-al_2010_MPB.pdf

DISSOLVED OXYGEN GUIDE

Dissolved Oxygen (abbreviated DO)

- Refers to oxygen gas dissolved in water
- Dissolved oxygen is a free oxygen molecule (O2), rather than the oxygen atom bound to hydrogen in a water molecule (H2O).
- Dissolved oxygen can be compared to gas dissolved in soda pop. The difference is that soda contains carbon dioxide (rather than oxygen gas) and that the concentration of carbon dioxide in soda is supersaturated (which is why it bubbles out when the can is opened and pressure is decreased).
- Supersaturation of oxygen occurs when the water holds more oxygen molecules than usual for a given temperature. Sunny days with lots of photosynthesis or turbulent water conditions can lead to supersaturation. Water is "saturated" at 100% and "supersaturated" above 100%.

Concentration of DO depends on three factors:

1. **Oxygen Sources** – external & internal sources which deliver oxygen to the water
   - **Atmosphere** – External source
     - Water absorbs oxygen from air, which is a mixture of many different gasses
     - Surface area increases oxygen absorption
     - Agitating water increases exposed surface area, allowing it to absorb more oxygen. Examples are waterfalls, currents, and waves
   - **Photosynthesis** – Internal source
     - Water absorbs oxygen from algae and plants. Through photosynthesis, algae and plants release pure O2 directly into the water during daylight hours.
   - Of the two oxygen sources, photosynthesis has a greater effect on DO

2. **Oxygen Sinks** – water's inhabitants consume the oxygen
   - Oxygen is removed from the water through respiration (by algae, plants, animals, bacteria, and other microorganisms) as well as other chemical reactions, including decomposition of organic wastes entering the water.
   - The oxygen consumed by these processes is called the Biological Oxygen Demand or BOD.

3. **Oxygen Solubility** – the extent to which oxygen is able to dissolve in the water
   - **Pressure**. Gas solubility increases with pressure. Water at low elevations (more pressure) holds more O2 than at high elevations (less pressure).
   - **Salinity**. Gas solubility increases as salinity decreases. This means freshwater holds more O2 than saltwater. (Seawater holds about 20 percent less DO than fresh water at the same temperature.)
   - **Temperature**. Gas solubility increases as temperature decreases. This means cold water holds more O2 than warm water.

Importance of DO

DO is essential to aquatic life. The amount of DO indicates the health of a body of water.

- Oxygen has limited solubility in water, usually ranging from 6 to 14 mg L^{-1} [1].
- There is no one EPA standard for DO 5ppm for rivers). DO requirements depend on the type of water body and the type of life it is expected to support. For example, fish have much higher metabolism than mussels.

DO in the intertidal:

- The high intertidal zone is exposed to air for a large part of the day. Daily water submersion brings necessary food, nutrients, and dissolved oxygen to support a wider variety of organisms than in the splash zone. This area still provides little refuge from the hot sun and cold air.
- The mid-intertidal zone is has longer submersion with access to increased nutrients and dissolved oxygen, creating a more hospitable habitat for marine organisms. These areas are typically dominated by mussels, gastropods, and algae, which provide shelter for mobile organisms (Donovan, et al., 2005).
- During low tide, organisms are either exposed entirely or left in pools, which may experience considerable temperature increase, salinity increase (because of evaporation), and oxygen demand from respiration, thus becoming become anoxic. Organisms which occupy this zone tend to be highly resistant to oxygen deficiency.
- DO also tends to be very low in the early morning hours because no organisms can photosynthesize at night, yet they all need to respire.

Guidelines for Interpretation of Dissolved Oxygen Readings

For mg/L:

- 0-2 mg/L: not enough oxygen to support most fish life (but organisms like crabs and mussels do fine)
- 2-4 mg/L: only a few kinds of fish can survive
- 4-7 mg/L: acceptable for warm water fish
- 7-11 mg/L: very good for most stream fish including cold water fish

For percent saturation:

- Below 60%: poor; water too warm or bacteria using up DO
- 60-79%: acceptable for most aquatic organisms
- 80-125%: excellent for most aquatic organisms
- 112% or more: too high, may be dangerous to fish (can get air bubbles in blood)

Table 1. Oxygen Sensitivities of some Marine Invertebrates

CLASSIFICATION	INTERTIDAL LOCATION	RESPONSE TO LOW OXYGEN (mg O_2/L)	REFERENCE
Gastropoda: (periwinkle) *Littorina littorea*	varied, to 15 m depth	LD_{50} at 0.21 (10 degrees C): 15 days	Theede *et al.,*1969
Pelecypoda: (mussels) *Mytilus edulis*	attached, in upper littoral	LD_{50} at 0.21 (10 degrees C): 35 days	Davis, 1975 Theede *et al.,*1969
(Oyster) *Ostrea virginica*	attached, in upper littoral	-urvived less than 0.5 mg O_2/L for at least one week	Davis, 1975
Thoracia: (barnacles) *Balanus amphitrite*	polluted, oxygen deficient areas	respiratory regulation ceased at 3.5 mg O_2/L	Parasada Rao and Ganapati, 1968
Decapoda: (crabs) *Uca pugnax*	marsh	respiratory-independent to a partial pressure of 40-50 mm Hg (2.6 mg O_2/L)	Davis, 1975
Echinoidea: (urchins) *Strongylocentrotus drobachienis*	hard bottoms	respiratory-dependent below 60-70 mm Hg (4.7 mg O_2/L) oxygen partial pressure	Davis, 1975
Asteroidea: (sea stars) *Asterias rubens*	hard bottoms	LD_{50} at 0.21 (10 degrees C): 4.3 days	Theede *et al.,*1969

Measuring DO

- DO is influenced by agitation and temperature.
- For accuracy in comparison, samples must be retrieved from the same location and the same depth, at the same time of day, as previous test samples.
- Water temperature must be recorded.
- DO can be measured with a membrane probe, laboratory titration, or field test kits.
- **Units of Measurement**
  - **mg/L** –mg O2/L H2O. Represents milligrams gas per liter of H2O. For example, a DO measurement of 9 mg/L means “9 milligrams of O2 per liter of H2O”.
  - **PPM “parts per million” –** PPM is the number of O2 molecules per million molecules. For example, a DO measurement of 9 ppm means “for every million molecules, nine are O2”. For every 1,000,000 molecules, 9 are O2 and 999,991 are H2O (assuming ideal laboratory conditions).
  - When reporting DO data, **mg/L = PPM**
  - **Percent saturation** – a relative measure where DO is expressed as a percentage of maximum possible saturation given temperature and salinity.
  - A simple way is to get percent saturation is to use a “nomogram” (Fig 1.)
    - Draw a line between temperature and DO to get the percent saturation value; thenaddt 20% for “normal” sea water.
    - For example, at DO = 11.5, Temp = -6 Celsius, % Saturation 70% in fresh water and 90% in salt water

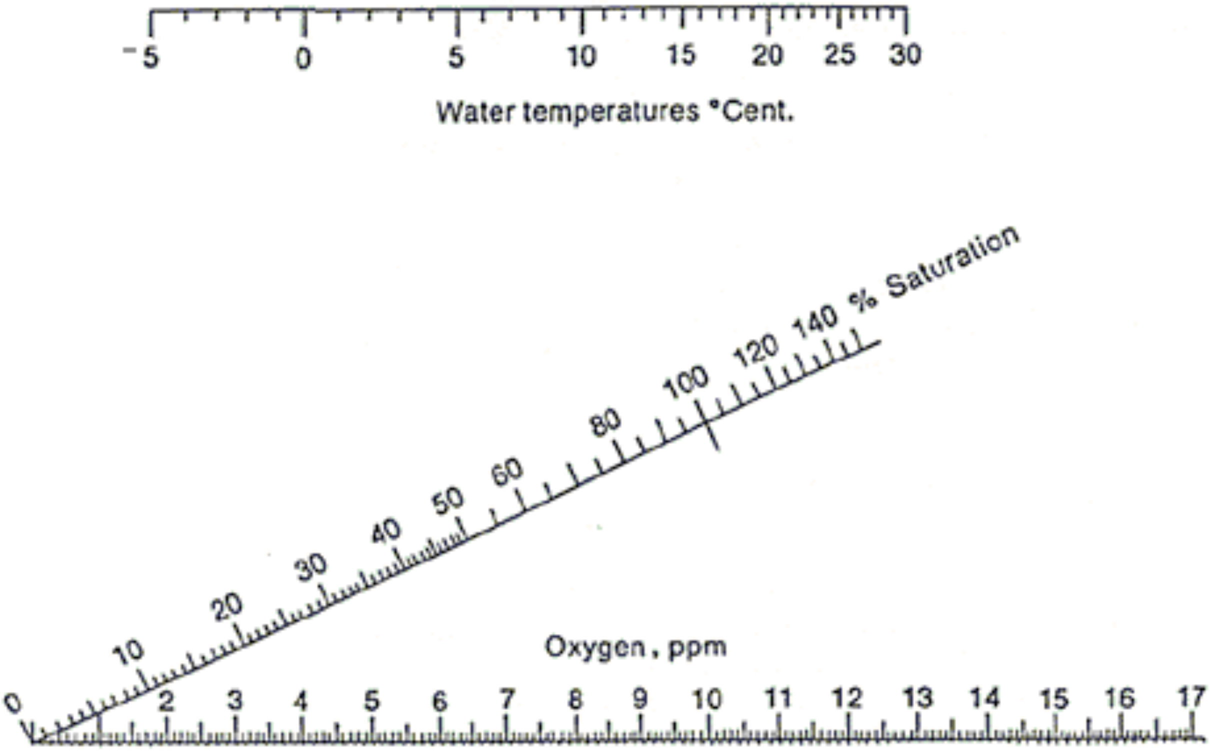

Fig. 1. Nomogram showing % DO saturation relative to temperature and oxygen concentration in fresh water.

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CHEMetrics Dissolved Oxygen Water Test Kit

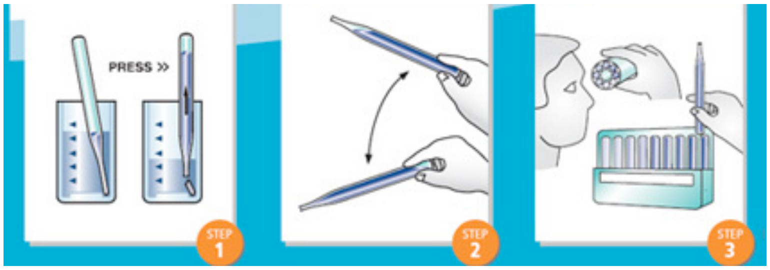
Catalog No.: C-7512
Range: 1-12 ppm
Increments: 1, 2, 3, 4, 5, 6, 8, 10, 12 ppm

Order: https://www.chemetrics.com

- **Audio and video instructions for test kit:** https://www.youtube.com/watch?v=GCQ9ckIbhbA
- **1 minute silent animation of test kit use:** https://www.youtube.com/watch?v=ttruI1RycRc
- **Video of test kit being used in Chesapeake Bay** https://www.youtube.com/watch?v=ek_qFEnAzec

Instructions: Vials contain an acidic solution and a yellow-green form of indigo carmine.

1. Fill the sample cup to 25mL with the sample to be tested.
2. Immerse the vial in the sample cup and snap the vial’s tip; the vial fills automatically.
3. Mix the vial by rotating it from end to end (there should be an air bubble that travels up and down).
4. Dry the vial and wait 2 minutes.
5. Oxygen in the water will oxidize the yellow-green indigo carmine into a colored blue form.
6. The resulting blue color is proportional to the dissolved oxygen concentration.
7. Test results are expressed as ppm (mg/L) O2.



Caution: The most critical part of any dissolved oxygen test is sampling. It is difficult to obtain a sample that accurately reflects the oxygen content of the source. Exposure to the high oxygen content of “air” will cause a sample to approach saturation.

- Dipping and pouring of the sample cup should be performed with as little agitation as possible.
- Analysis should be performed immediately after sample collection.
- At the time of analysis, CHEMets® should be gently lowered into the color comparator to avoid introduction of atmospheric oxygen into the ampoule.

Safety: Safety Data Sheets (SDS) are available in the OPIHI PD section of exploringourfluidearth.org and at www.chemetrics.com. Read SDS before using these products.

- Breaking the tip of a vial in air rather than water may cause the glass vial to shatter.
- Wear safety glasses and protective gloves.

Note: Seawater may occasionally cause the reagent to precipitate.

TEMPERATURE GUIDE

Temperature

Temperature is a measure of the average kinetic energy of molecules, which is usually recorded on a recognized scale, such as Celsius. Water temperature impacts and is impacted by physical, chemical, and biological processes. The temperature in large bodies of water is generally controlled by the climate and currents in a region. In small pools or shallow lakes, temperature is affected by shading from rocks or vegetation, water flow rate, and water depth. Additionally, industrial discharge or sewage outflow can decrease or increase the temperature of the surrounding waters, an effect called thermal pollution. Changes in water temperature can change density and chemical speciation (the forms of chemical compounds present in the water).

Importance of Temperature

Because temperature influences a wide variety of physiological processes, it is often cited as the single most important abiotic factor impacting the ecology of marine organisms. Some organisms are able to withstand a wide range of temperatures, whereas others can be adversely affected by even a small change in temperature.

Temperature is likely more important in the tropics than other latitudes as organisms experience high temperature and irradiance levels and can be acclimated to only narrow thermal regimes. For example, an increase in only a few degrees above normal conditions in shallow marine waters results in coral bleaching.

Temperature in the Intertidal

Temperature in the intertidal depends on a number of factors including wave action, tidal height, air and water temperature, rainfall, substrate, and rugosity. The intertidal is very complex thermally, which a range of temperatures that can quickly change, but there are some general principles. There is generally a temperature gradient from higher to lower from the high to low intertidal. Intertidal areas with darker basalt substrate will have a higher temperature than lighter limestone substrate. Rugosity, or how uneven the substrate is, and slope can affect temperature by providing shade to organisms in cooler cracks and crevices. In intertidal areas with tidepools temperature is affected by the rugosity, volume, and depth of the pools. Invertebrates in intertidal areas exhibit cooling behaviors, like moving to shaded areas, and have morphologies to deal with thermal stress.

Measuring Temperature

Alcohol thermometers rely on the expansion of that liquid in a small, calibrated space. Temperature probes work on a similar physical characteristic. The probe tip has two dissimilar metals welded together to create a joint. As the temperature at that joint changes the metals expand or contract different amounts and therefore have different levels of electrical resistance. The smaller that joint or welding is, the more accurate the instrument will be at getting a temperature at a specific *point*. Therefore, thermistors, the larger probes, covered in a metal housing, are less point accurate than a thermocouple, the smaller probe with only wires.

Because temperature in the intertidal is so variable over temporal and spatial scales, consistency and repetition is key when examining it. Areas should be sampled at the same time (e.g., low tide) in multiple locations (e.g., multiple samples in the high, middle, and low intertidal). The time and location should be recorded as accurately as possible.

Temperature loggers record temperature at a set interval and can be left in place for long periods of time. Because temperature in the intertidal is so variable, they can give the most accurate description of a temperature at a certain point over time. However, they are expense, can become dislodged or stolen, and require permission from government authorities if they are left out for long time periods.

Water Temperature of Hawaiian Coast

https://www.nodc.noaa.gov/dsdt/cwtg/hawaii.html

SALINITY GUIDE

Salinity

Salinity is a measure of the dissolved salt content of water. Salinity in ocean water is due to minerals carried by rivers from continental materials and geologic processes that occurred when the oceans formed. Salinity can be affected by evaporation and freshwater input from rain, rivers, and glacial melting. Most marine and freshwater organisms have a narrow range of salt tolerance, however, some organisms that live in environments like the intertidal and estuaries have a wide range of salt tolerance.

Salinity in the Intertidal

Salinity in the intertidal is affected by rain and other freshwater input, all of the factors that effect evaporation rates (temperature, wind, wave action,) and location in the intertidal (high vs. low). In intertidal areas with tidpools in the high intertidal can experience high evaporation rates and become very salty during low tide. Pools in the splash zones that are never underwater can develop salt crystals.

Most organisms in the intertidal have marine ancestors. Marine organisms are generally in osmotic balance with seawater. When they are immersed in water that is higher or lower in salinity than normal they face osmotic imbalance and gain or lose water. Lower salinity water (e.g., due to big rain events) is generally considered more stressful than higher salinity water (e.g., due to evaporation) as tides generally flush out hypersaline water regularly, so organisms are more adapted to this stress.

Salinity—Parts Per Thousand

Salinity is the measure of the number of grams of salts per kilogram of seawater, which is expressed in parts per thousand. **Parts per thousand** can be defined as how many parts, or grams, of salt there are per thousand parts, or kilogram (1,000 g), of seawater. The symbol for parts per thousand is ‰. Parts per thousand is commonly abbreviated as ppt.

$$\text{Salinity in ppt} = \frac{\text{grams of dissolved salts}}{\text{1,000 grams of seawater}}$$

The average salinity of seawater is about 35 grams per kilogram (g/kg) of seawater, or 35 ppt.

Measuring Salinity

Hydrometer

When salt is dissolved in fresh water, the density of the water increases because the mass of the water increases. In a sample of water with a known temperature, density can be measured with a hydrometer. A hydrometer determines the density of a liquid by measuring the level at which the hydrometer (or part of a hydrometer) floats in it. The more dense the liquid, the greater the buoyant force it exerts, and the higher the hydrometer will float in the liquid. How “high” the hydrometer (or part of the hydrometer) floats in a liquid corresponds to the density of the liquid.

Hydrometers measure specific gravity. **Specific gravity** is the ratio of the density of a substance to the density of a reference substance, generally water, at a certain temperature and pressure. Substances with a specific gravity of one are neutrally buoyant in water, substances with a specific gravity of less than one are less dense than water, and substances with a specific gravity greater than one are more dense than water.

The average specific gravity of seawater is about 1.026.

Hydrometers are standardized for temperature as the density of water can be affected by temperature. When the same amount of water is heated or cooled, its density changes. When the water is heated, it expands, increasing in volume. The warmer the water, the more space it takes up, and the lower its density.

Using a Hydrometer: Instant Ocean

The instant ocean hydrometer operates on the principle that the pointer will read the salinity of a sample when it is neutrally buoyant and the buoyant force (weight of water displaced) is equal to the gravitational force of the pointer.

Instructions:

Use in water temperatures of 68–85 F (20–30 C);
Before use, rinse with fresh water.

Step 1 - Slowly fill the Instant Ocean Hydrometer by dipping bottom corner fill port below water surface until water flows up and over inner weir. Slow fillings means you will have less bubble formation.
Step 2 – Dump this fill as the rinse with fresh water may lower salinity readings.
Step 3 – Repeat step 1.
Step 4 - Dislodge air bubbles by gently tapping hydrometer or pointer. Air bubbles on the pointer may result in inaccurate results.
Step 5 - Place the hydrometer on a level surface and read specific gravity (inside scale) and salinity (outside scale).
Step 6 – Take multiple readings and average results.
Step 7 – Rinse hydrometer with fresh water before storing.

Calibration:

Periodically calibrate your hydrometer by mixing up known salinity solutions and measuring them.

Care:

Rinse thoroughly in freshwater after each use to prevent mineral and salt buildup which can affect accuracy. Salt and calcium deposits that accumulate on the pointer will result in incorrect results. If crystallization occurs inside the unit, soak in lukewarm water or vinegar for 30 minutes, rinse with freshwater and air dry.

Refractometer

When light enters a liquid it changes direction; this is called refraction. Refractometers measure the degree to which the light changes direction, called the angle of refraction. A refractometer takes the refraction angles and correlates them to refractive index (nD) values that have been established. Using these values, you can determine the concentrations of solutions.

The prism in the refractometer has a greater refractive index than the solution. Measurements are read at the point where the prism and solution meet. With a low concentration solution, the refractive index of the prism is much greater than that of the sample, creating a large refraction angle and a low reading. The reverse would happen with a high concentration solution. The refraction index is very temperature dependent.

http://www.coleparmer.com/TechLibraryArticle/633

Tips for Using a Refractometer

1. Make sure to rinse with deionized water and dry completely with a wipe that will not leave lint.
2. Each reading should be done by the same person and facing the same direction to the sun for consistency.

Conductivity Probe

Conductivity probes measure the conductance, the inverse of resistance, between two electrodes as ions flow between them. Liquids that have more salts are more conductive. Probes are calibrated to operate best within certain salinity ranges.

TURBIDITY GUIDE

A turbidity tube uses the correlation between visibility and turbidity to approximate a turbidity level. A marker is placed at the bottom of the turbidity tube until it can no longer be seen from above due to the “cloudiness” of the water. This height from which the marker can no longer be seen correlates to a known turbidity value. For more information see the Turbidity Tube guide from Myre and Shaw of Michigan Technological University.

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Water Sampling Methodology

Questions to Consider When Developing a Sampling Scheme

Purpose	• Why are you collecting this data?
Representativeness	• How will the samples represent the time and space in which they are collected?
Measurements	• What measurements will you make? • What instruments or tests will you use? • What types of samples will you need? • How will you prepare your equipment before your field study?
Accuracy and Precision	• How can you get most accurate and precise data? • How can you avoid error? • What quality control can you practice?
Contamination	• What are you measuring? • What might contaminate your sample relative to what you are measuring? • How can you avoid contamination?
Sampling Sites	• How will you choose sites that meet the purpose of your study, the types of data you need, and the equipment/methods you have available to you? • What physical characteristics in the area might impact the chemistry of the water? • What biological characteristics in the area might impact the chemistry of the water?

Sampling Sites (cont.)	• What human impact might have an effect on the chemistry of the water? • Is water flowing in one direction and should you sample a transect along that flow? • Might depth play a role in water chemistry? • Might there be a difference in water chemistry at different depths? • Have samples been collected in this area before? If so, where and what sampling scheme was used? • How can you cover the study area within time and other constraints?
Environment	• What conditions in the environment could affect your sampling? What might you need to change about your scheme based on changes in the environment?
Safety	• What safety measures do you need to take to protect yourself in and around the water? With chemicals? Near dangerous animals/plants? • What safety measures do you need to take to protect the environment?
Supplies	• What supplies do you need in the field? • How will you prepare these supplies? • What consumable supplies will you need extras of (deionized water, batteries, gloves, pencils, etc)?

Water Collection and Testing Procedures

Clean all sampling and testing containers before going into field	• Wash with soap and water • Rinse three times with tap water • Rinse three times with deionized water
Collecting samples	• Use clean hands/dirty hands technique • Wear gloves • Don't touch contaminating substances (equipment, food, etc.) • Do not touch sample with gloved or ungloved hands! • Rinse collection container once with water to be collected • Collect water sample, cover with lid • Leave sample covered as much as possible to avoid contamination
Reagent methods	• In between samples, rinse three times with tap water, three times with deionized water, then three times with small amount of sample
Probe methods	• If necessary, rinse probe with DI water in between samples
All testing methods	• Follow procedure specified for method