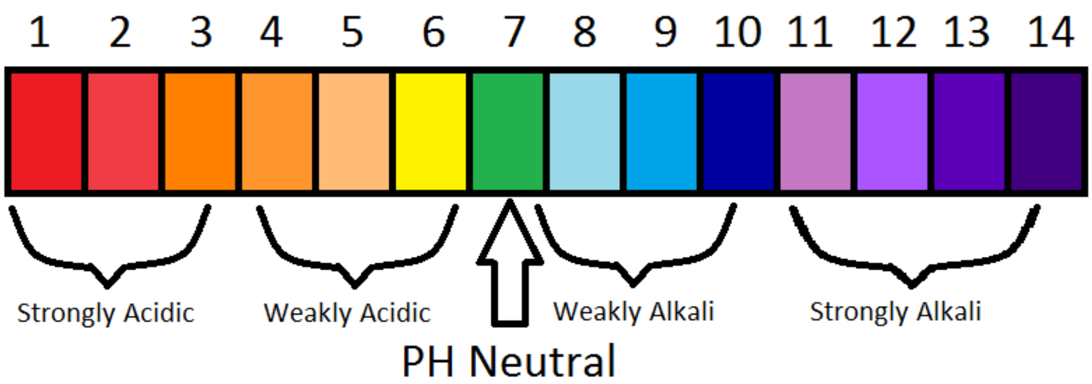


pH (The Power of Hydrogen!)

pH is a measure of acidity or alkalinity. The pH is measured on a scale of 1 to 14; 1 being the most acidic, 14 being the most alkaline, or basic, and 7 being neutral. pH is technically a way of measuring the concentration of hydrogen ions (H+) in water - at a higher pH there are fewer free H+ ions (more alkaline) and lower pH levels have more free H+ ions (more acidic).

A change in one pH unit equals a tenfold change in the concentration of H+. For example, there are ten times fewer H+ ions at a pH of 8 than at a pH of 7 (one pH unit), where a pH value of 9 has one hundred times fewer H+ ions as a value of 7 (two pH units).

A pH below 4 or above 10 will kill most fish, and very few animals can tolerate waters with extreme pH levels. Most aquatic animals and plants have adapted to life in water with a pH range of 6.5 to 8.5. Acidic waters can reduce the hatching success of fish eggs, irritate fish and aquatic insect gills, and damage membranes.

Heavy metals, such as copper, lead and zinc can leach from soil particles under acidic conditions, and may be ingested by aquatic organisms, with varying effects. Bio-accumulation seems to occur only with mercury. Phosphorus can also be more readily released, contributing to overgrowth of aquatic plants and low oxygen levels.

Natural factors can cause pH values to fluctuate. Both unpolluted rain and photosynthesis contribute carbon dioxide to water to form carbonic acid ($CO_2 + H_2O \rightleftharpoons H_2CO_3$), a very weak acid. Respiration of plants at night removes the carbon dioxide from the water; photosynthesis by algae and plants (which uses hydrogen) resumes with sunlight. These processes, as well as geology, tend to naturally balance changing pH and is called *buffering capacity*.

When radical changes in pH occur, we start looking for human influence.

Dissolved Oxygen (DO)
(Acceptable Levels: 5 ppm – 10 ppm)

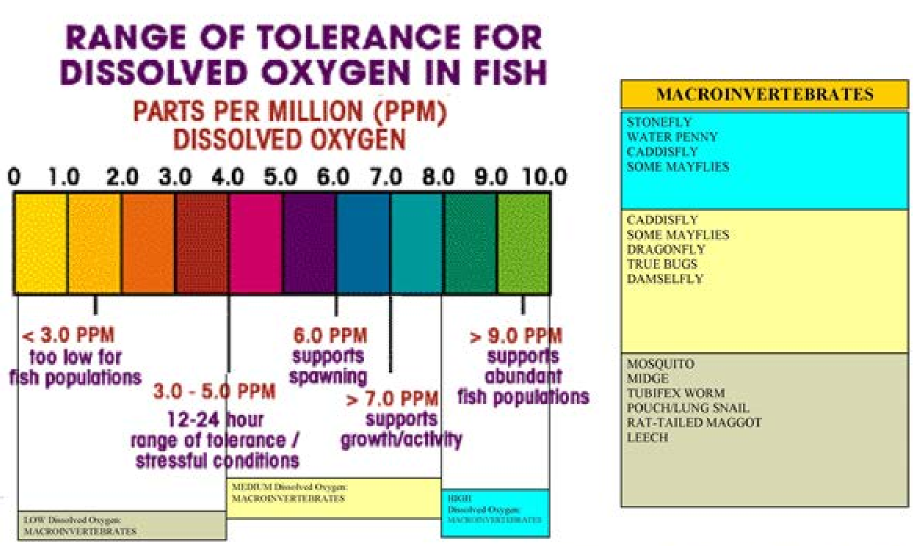
Because aquatic animals breathe through gills, they need dissolved oxygen to live. Aquatic plants, algae and bacteria also require oxygen for respiration. Oxygen is diffused into water from the atmosphere and when water bubbles over riffles in streams. Oxygen is also produced by aquatic plants, algae, and phytoplankton as a by-product of photosynthesis. Once dissolved in water, the oxygen is distributed by the actions of currents and turbulence within the waterbody. Cold water can hold more dissolved oxygen than warm water.

Stream systems produce oxygen, but they also consume it. Respiration, or use of oxygen, occurs by plants at night, by aquatic animals, and by bacteria decomposing organic material. Various chemical reactions within water will also use oxygen. The amount of oxygen used for these processes is called the biochemical oxygen demand, or BOD.

If more oxygen is consumed than is produced, dissolved oxygen levels will decline and organisms sensitive to changes in dissolved oxygen will be impacted to varying degrees. Some organisms can simply move away to a more oxygen rich habitat, but others may weaken or die.

The amount of oxygen required varies according to species of aquatic organisms and their stages of life. Dissolved oxygen levels of at least 6-7 ppm are usually required for spawning, growth and activity of fish. Levels between 3 and 5 ppm are stressful to most aquatic organisms, with levels below 3 often too low to support fish or less tolerant aquatic macroinvertebrates.

High levels of bacteria from sewage, decomposing organic matter, and high water temperatures can cause dissolved oxygen to decline, and will affect the ability of plants and animals to thrive.



http://www.water-research.net/Watershed/dissolvedoxygen.htm oec@wilkes.edu

Nitrate (NO_3^-)
Usual Range: < 1 ppm (5 ppm ranks fair, but is cause for concern)

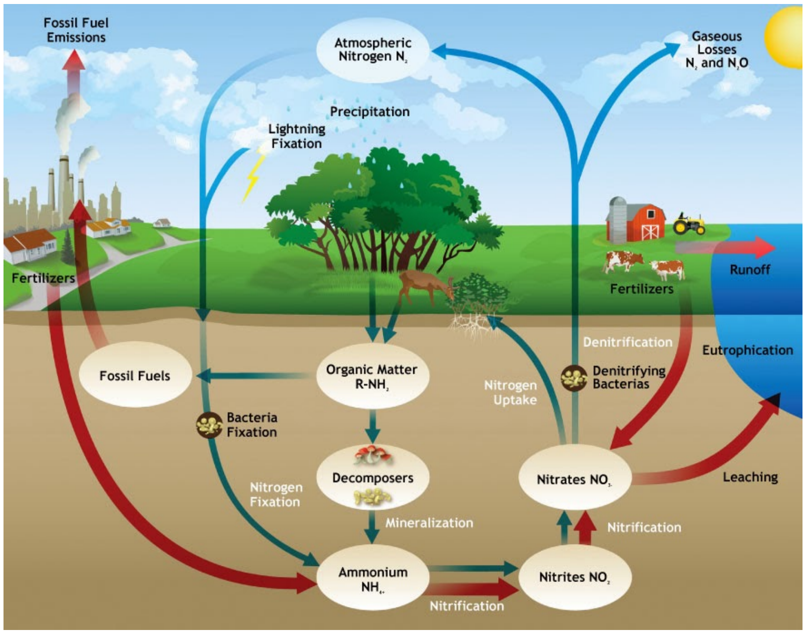
Nitrate is a nutrient needed by all aquatic plants and animals to build protein, and is excreted by living animals. It is released as dead plants and animals decompose. Nitrates in water may also indicate presence of other more serious residential or agricultural pollutants, such as bacteria or pesticides. When nitrates are present in quantities in excess of 10 mg/L, the water supply can pose a potentially fatal threat to infants under six months and to young and pregnant animals through a condition called methemoglobinemia, or Blue Baby Syndrome.

Nitrogen (N) is an essential plant nutrient in the form of NO_3, but in excess amounts can cause significant water quality problems similar to those associated with phosphorus. Together with phosphorus, N in excess amounts can accelerate eutrophication*, causing dramatic increases in aquatic plant growth and changes in the types of plants and animals living in a stream or lake. These changes subsequently affect dissolved oxygen levels, temperature, and other physical and chemical characteristics of streams and lakes.

Like phosphorus, nitrogen is classified as organic or inorganic, and there is a similar cycle that regulates the abundance of nitrogen in a stream system. Inorganic nitrogen is used as a nutrient by plants, converted to organic nitrogen which can then be used by animals. Sources of nitrogen include: atmospheric deposition via precipitation, wastewater treatment plants, runoff from fertilized lawns or croplands, failing septic systems, runoff containing animal wastes, power plants, and industrial discharges.

The natural level of nitrates in surface waters is typically low (less than 1 milligram per liter), but the effluent from wastewater or sewage treatment plants can have concentrations that can range up to 30 mg/L. Ammonia, another form of nitrogen, is not detectable with a nitrate test.

***Eutrophication** - *"The process by which a body of water acquires a high concentration of nutrients, especially phosphates and nitrates. These typically promote excessive growth of algae. As the algae die and decompose, high levels of organic matter and the decomposing organisms deplete the water of available oxygen, causing the death of other organisms, such as fish. Eutrophication is a natural, slow-aging process for a water body, but human activity greatly speeds up the process." - Art, 1993*



http://imgarcade.com/1/decomposition-cycle/

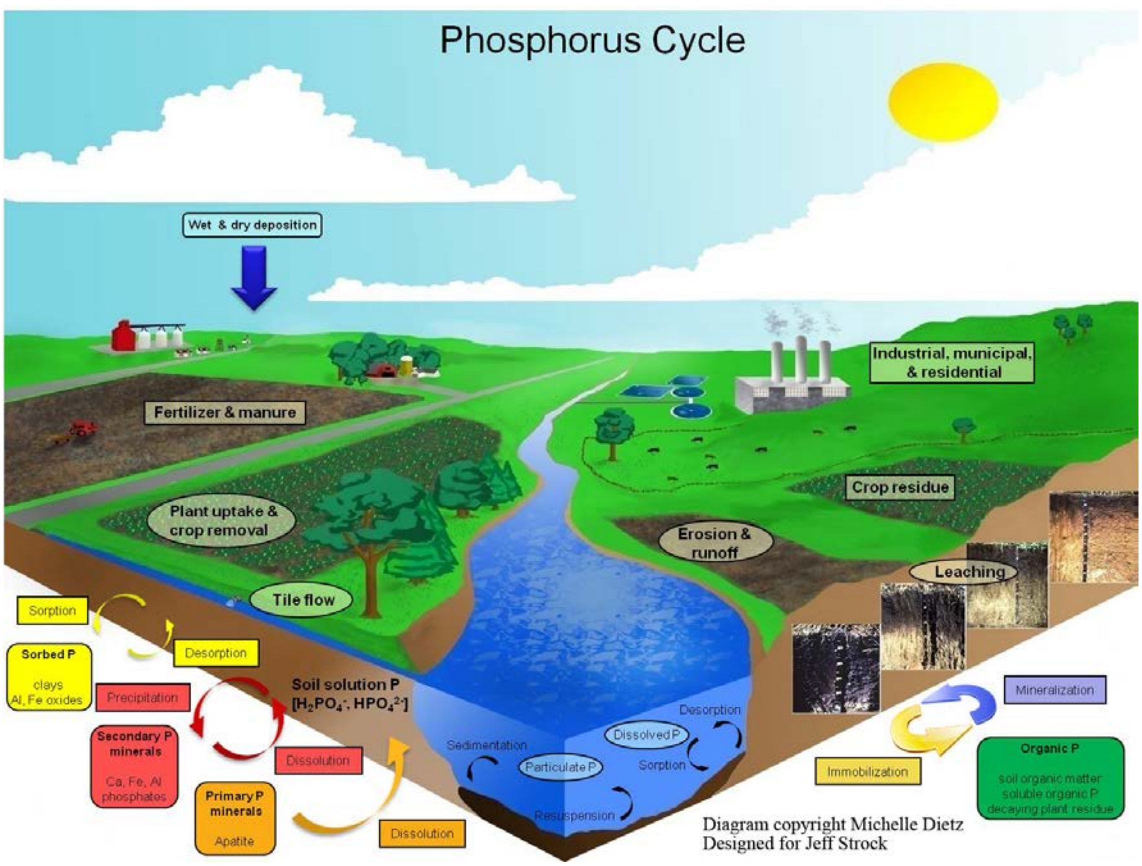
Phosphate (PO_4)
Good Range < 1-2 ppm Fair: 4 ppm

Like carbon, oxygen, hydrogen, and nitrogen, phosphorus is a limiting nutrient for all forms of life, which means that the potential for an organism's growth is limited by the availability of this vital nutrient. It forms part of the structure of DNA and RNA, is needed for energy transport in cells, provides structure to cellular membranes, and assists in giving bones and teeth their rigidity. In short, without phosphorus, we simply could not exist.
http://www.visionlearning.com/en/library/Earth-Science/6/The-Phosphorus-Cycle/197

Healthy aquatic ecosystems have naturally low concentrations of phosphorus. Too much of it in waterways leads to plant overgrowth, algae blooms (some toxic), and eutrophication. Unlike nitrates, which readily dissolve in water, phosphates tend to attach to soil particles.

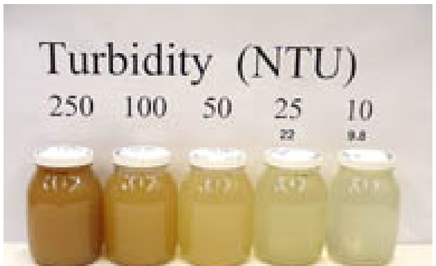
Phosphorus (P), like nitrogen and water, cycles through the environment, but does not have a gas phase. Its presence in water has many sources, both naturally occurring and human made - water percolating through phosphorus rich soil and rocks; discharge from wastewater treatment plants; runoff from fertilized lawns and croplands; failing septic systems; runoff containing particles of animal manure; disturbed land areas; drained wetlands; water treatment; and commercial cleaning preparations.

Significant reductions (25%-81%) in phosphates in NC waterways occurred after the 1988 phosphate detergent ban in NC, according to US Geological Service research.



http://swroc.cfans.umn.edu/ResearchandOutreach/SoilManagement/SoilResearch/PhosphorusCycle/index.htm

Turbidity



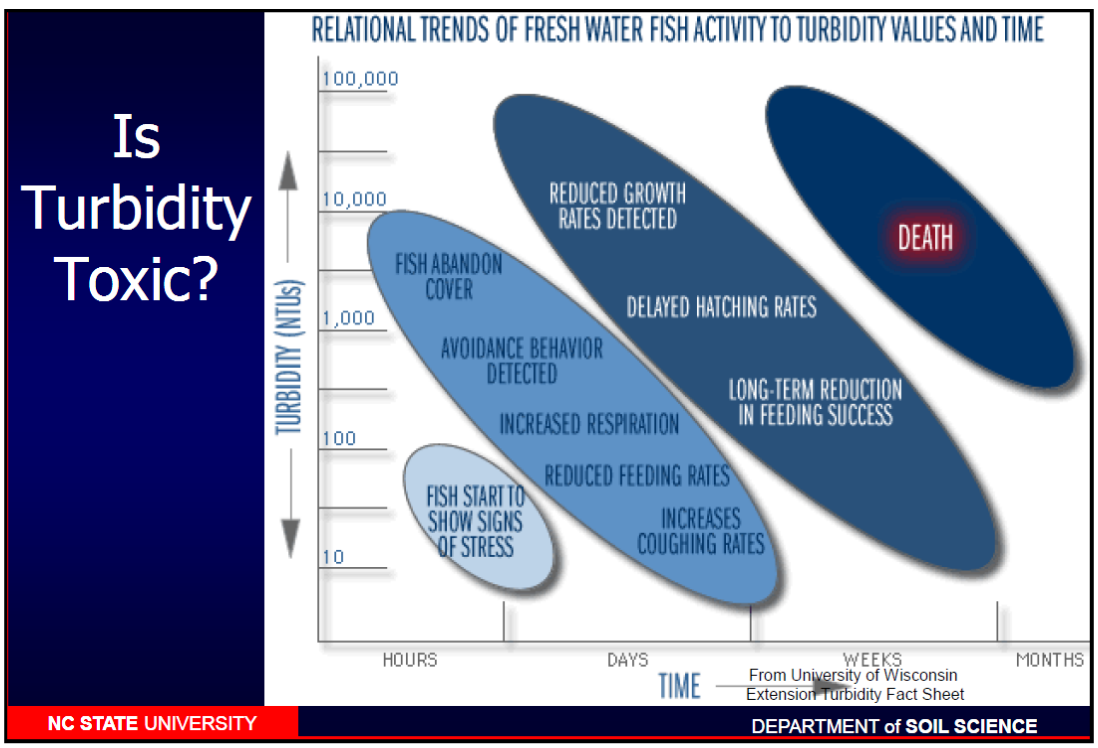
water.ncsu.edu/watershedss/info/turbid.html

Turbidity is a measure of the clarity of water. Turbidity should not be confused with color, since darkly colored water can still be clear and not turbid. Suspended materials such as clay, silt, organic and inorganic matter, and microscopic organisms, directly affect turbidity by decreasing the passage of light through the water. This reduces or prevents photosynthesis of plants and leads to decreases in the production of dissolved oxygen. High turbidity can also increase water temperatures as the suspended particles absorb more heat than the water could otherwise, which further decreases dissolved oxygen levels since warm water holds less dissolved oxygen than colder water.

Turbid waters can clog gills of aquatic organisms, reduce their resistance to disease, lower growth rates, and inhibit the ability to see and catch food. Lowered growth rates can affect the development of fish eggs and aquatic larvae. As solid particles settle, they can blanket the stream bottom; smothering fish eggs and benthic macro-invertebrates, which are essential for a healthy stream ecosystem. Harmful cyanobacteria (blue-green algae) are favored in turbid waters because they possess flotation mechanisms (McCabe et al., 1985), and can outcompete other plants and algae for light.

Common sources of turbidity include soil erosion, waste discharges, urban runoff, stream bank erosion, bottom sediment disturbance and excessive algal growth.

Turbidity is commonly measured by a meter, measured in Nephelometric Turbidity Units (NTU).



http://www.ncsu.edu/wrri/pdfs/pastevents/lp12012/McLaughlin.pdf

Heavy Metals (We do not test because of the expense, but information is good to know.)

Heavy metals are elements that have atomic weights between 63.5 and 200.5 and have a specific gravity greater than four. Living organisms require trace amounts of some heavy metals to survive, and these metals are known as essential metals. They include: cobalt, copper, iron, manganese, molybdenum, vanadium, strontium, and zinc. Though these metals are essential, excessive amounts can be harmful to living organisms. There are also certain non-essential heavy metals that are of particular concern for water quality, including: cadmium, chromium, mercury, lead, arsenic, and antimony. All heavy metals exist in surface waters in particulate, dissolved, and colloidal (evenly distributed microscopic particles in a dispersing solution, i.e. water) phases, though dissolved concentrations tend to be low. Particulate and colloidal phases of these metals may be found bonded to other molecules (hydroxides, oxides, silicates, and sulfides) or as part of silt, clay, or varying types of organic matter. The solubility, or ease with which metals become part of the water column, is predominantly controlled by the pH concentration of water, redox potential (basically indicates how easily chemical reactions are taking place in the water), and the type and concentration of ligands (ions, molecules, or molecular group that binds to a chemical entity to form a larger complex) on which the metals could adsorb (bind to the surface of another particle).

The behavior of metals in natural waters is a function of the sediment composition, suspended sediment composition, and water chemistry. Different types of sediment, i.e. clay versus sand, have varying potential to bind to and subsequently store metals. The water chemistry of a system controls the rate of *adsorption* and *desorbtion* of metals. In the case of water quality, adsorption occurs when metals in the water column bind to the surface of sediment particles, which will generally sink to the bottom and become part of the substrate sediment. Desorbtion occurs when the metals are released from their bond to sediment and returned to the water column. The process of desorbtion allows for the recirculation of metals and potential bioassimilation (use by living organisms). Metals may be desorbed when salinity (salt content) of water increases, when redox potential changes, or pH decreases.

There are both natural and anthropogenic (human-made) sources of heavy metals in surface and ground water, though human-made sources of metals in natural waters greatly exceed natural sources. Natural sources include chemical and physical weathering of igneous and metamorphic rock and soil, decomposition of plant and animal detritus, precipitation or atmospheric deposition of airborne particles from volcanic activity, wind erosion, forest fire smoke, and oceanic spray. Human-made sources include: combustion of fossil fuels, urban stormwater runoff, domestic wastewater effluent, corrosion of water pipes, consumer products, and industrial effluents and waste sludge.

Excess levels of metals in natural waters pose a health risk to the entire biotic community (all living organisms) and the environment. Slightly elevated levels in natural waters can cause changes in tissue structure of aquatic organisms, changes in physiology and circulation of aquatic organisms, changes in biochemistry, changes in behavior, and changes in nearly all facets of reproduction. Many organisms can regulate the concentrations of metals in their tissues. Fish and crustaceans can excrete excess essential metals, but aquatic plants and bivalves (shellfish and mollusks) cannot regulate metal uptake, and subsequently suffer from metal accumulation and eventual metal toxicity.

Please plan your monitoring date and call the Stormwater office at 919-969-7246 to schedule pickup of your testing kits. Kits may be checked out for 10 days at a time.

We will pack instructions for conducting tests, and the following equipment and supplies for water quality testing:

Color comparator

Test tubes

Prepackaged reagents

Disposal bottle to hold rinse water & reacted test water

Clean, lint-free wipes to clean and dry the sample tubes

Recording form

**DO NOT DUMP REAGENTS, REACTED TEST SAMPLES, or any chemicals on the ground or in the water. These must be contained and may be safely flushed down a sanitary sewer drain with tap water.

Testing for Dissolved Oxygen

1. Record the temperature of the stream water from where you will fill the test tubes.

2. Fill the smallest test tube to overflowing with sample water. Carefully remove the tube from the water, keeping the tube full to the top (to prevent air from entering).

3. Add 2 DO TesTabs to the test tube. Water will overflow when tablets are added.

4. Screw the black cap on the tube. More water will overflow as the cap is tightened. Be sure no air bubbles are in the sample.

5. Gently invert the test tube over and over until the tablets have disintegrated: about 4 minutes.

6. Wait 5 more minutes for the color to develop.

7. Compare the color of the sample to the Dissolved Oxygen Color Chart.

8. Record the result as ppm DO and % Saturation

PERCENT SATURATION

TEMP°C	DISSOLVED OXYGEN 0 ppm	4 ppm	8 ppm
2	0	29	58
4	0	31	61
6	0	32	64
8	0	34	68
10	0	35	71
12	0	37	74
14	0	39	78
16	0	41	81
18	0	42	84
20	0	44	88
22	0	46	92
24	0	48	95
26	0	49	99
28	0	51	102
30	0	53	106

*Calculations based on solubility of oxygen in water at sea level, from Standard Methods for the Examination of Water & Wastewater, 18th edition.

Locate the temperature of the water sample on the Percent Saturation chart. Locate the Dissolved Oxygen result of the water sample at the top of the chart. The Percent Saturation of the water sample is where the temperature row and the Dissolved Oxygen column intersect.

FOR EXAMPLE: if the water sample temperature is 16°C and the Dissolved Oxygen result is 4 ppm, then the Percent Saturation is 41.

DISSOLVED OXYGEN (% SATURATION)	SCORE
91-110	4 (excellent)
71-90	3 (good)
51-70	2 (fair)
<50	1 (poor)

Record the score on the Data Sheet.

Testing for pH

1. Fill the test tube to the 10ML line with your water sample.

2. Add one pH Wide Range (WR) TesTab

3. Cap and mix by inverting until the tablet has disintegrated. Bits of material may remain in the sample.

4. Compare the color of the sample to the pH color chart. Record the result as pH.

Testing for Phosphate

1. Fill the test tube to the 10 mL line with water sample.

2. Add one Phosphate WR TesTab.

3. Cap and mix by inverting until the tablet has disintegrated. Bits of material may remain in the sample.

4. Wait 5 minutes for the blue color to develop.
   (NOTE: If the sample remains colorless, record the result as 0 ppm)

5. Compare the color of the sample to the Phosphate color chart.
   Record the result as ppm Phosphate.

Testing for Nitrate

1. Fill the test tube to the 5 mL line with water sample.

2. Add one Nitrate WR TesTab. Immediately slide the test tube into the protective sleeve.

3. Cap and mix by inverting for 2 minutes to disintegrate the tablet. Bits of material may remain in the sample.

4. Wait 5 minutes for the red color to develop. Remove the tube from the sleeve.

5. Compare the color of the sample to the Nitrate color chart. Record the result as ppm Nitrate.

Basic Instructions for the Turbidity Tube

1. Close the drain tube by squeezing the crimp.
2. Fill the transparency tube with the water sample.
3. While looking down through the opening of the tube, partially open the drain crimp and slowly draw off the sample. (Control the flow by squeezing the crimp.)
4. When the black and white pattern, at the base of the transparency tube, faintly begins to appear — immediately tighten the crimp and record the level of water remaining via the centimeter rule on the side of the tube.

5. Record in Cm.

Distance from bottom of tube (cm)	NTU's
< 6.25	> 240
6.25 to 7	240
7 to 8	185
8 to 9.5	150
9.5 to 10.5	120
10.5 to 12	100
12 to 13.75	90
13.75 to 16.25	65
16.25 to 18.75	50
18.75 to 21.25	40
21.25 to 23.75	35
23.75 to 26.25	30
26.25 to 28.75	27
28.75 to 31.25	24
31.25 to 33.75	21
33.75 to 36.25	19
36.25 to 38.75	17
38.75 to 41.25	15
41.25 to 43.75	14
43.75 to 46.25	13
46.25 to 48.75	12
48.75 to 51.25	11
51.25 to 53.75	10
53.75 to 57.5	9
57.5 to 60	8
Over 60	< 8