Understanding Water Quality Parameters for Citrus Irrigation and Drainage Systems[1]

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Introduction

A basic knowledge of water quality is very useful for microirrigation system management and is an important consideration in the design and operation of the system. Therefore, the ability to read and understand a water quality analysis is important to the irrigation system manager. A careful analysis of the source water is prudent as a preliminary step to designing a microirrigation system. A microirrigation system requires good quality water free of all but the finest suspended solids, and relatively free of dissolved solids such as iron, which may precipitate out and cause emitter plugging problems. Neglecting to analyze the quality of source water and provide adequate treatment is one of the most common reasons for the failure of microirrigation systems to function properly.

Obtaining a Water Sample

It is important that a representative water sample be taken. If the source is a well, the sample should be collected after the pump has run for approximately a half hour. When collecting samples from a surface water source such as a ditch, river, or reservoir, the samples should be taken near the center and below the water surface. Where surface water sources are subject to seasonal variations in quality, these sources should be sampled and analyzed at various times throughout the irrigation season.

Glass or plastic containers are preferable for sample collection. For most analyses, the samples should be at least a pint. The containers should be thoroughly cleaned and rinsed before use to avoid contamination of the water sample. Sample bottles should be filled completely to the top (with all air removed), carefully labeled, and tightly sealed. Samples should be sent immediately to a water testing laboratory. For a microirrigation suitability analysis, the following tests should be requested from the laboratory: Electrical conductivity (EC) or total dissolved solids (TDS), pH, calcium, iron, alkalinity, chloride, and if filtration information is needed, the quantity and size of suspended solids. Other parameters that are sometimes also needed to properly assess the suitability of the irrigation water are sodium, boron, potassium, manganese, nitrate, and sulfides (hydrogen sulfide must be measured at the well head).

Units Used in Water Analysis

One obstacle in interpreting a water analysis is that water testing laboratories report results in various units. For example, the concentration of chemical constituents may be reported as parts per million (ppm), milligrams per liter (mg/L), or milliequivalents per liter (meq/L).

A concentration reported in weight-per-weight is a dimensionless ratio and is independent of the system of weights and measures used in determining it. For many years,

water analyses made in the U.S. were reported in parts per million (ppm). One ppm is equivalent to 1 mg of solute per kg of solution. Therefore, 1% is equal to 10,000 ppm. Parts per thousand (ppt, which is equivalent to 1 g solute per kg of solution) is frequently used to report salinity levels in sea water. In recent years, the development of more sensitive analytical equipment has led to terms such as parts-per-billion to report trace elements and pesticide concentrations.

Common laboratory analyses are made by initially measuring a volume of solution and then performing the appropriate procedure. The results are normally reported in units of weight per unit of volume (typically mg/L). To convert these values to ppm, it is generally assumed that 1 liter of water weighs exactly 1 kg. For most agricultural water, this assumption is close enough. Therefore, for agricultural water samples, ppm and mg/L can be considered the same. In some cases, water analyses (typically hardness) are presented in units of grains per gallon (1 grain/gallon equals 17.12 mg/L). Conversion factors for several water quality parameters are given in Table 1.

Sometimes analytical labs report results in terms of milliequivalents per liter (meq/L). This term recognizes that ions of different species have different weights and electrical charges. Conversions are made by dividing ppm (or mg/L) by the equivalent weight of the constituent ion (Table 2) as follows:

meq/L = (mg/L) / EW   Eq. 1

mg/L = meq/L x EW   Eq. 2

where,

meq/L = milliequivalent per liter

mg/L = milligram per liter

EW = equivalent weight

Example:

How many meq/L is 120 ppm Ca?

From Table 2, EW for Ca = 20

meq/L = (mg/L)/EW = 120 / 20 = 6 meq/L

How many mg/L is 5 meq/L of CO_3

From Table 2, EW for CO_3 = 30

mg/L = 5 meq/L x 30 mg/L per meq/L = 150 mg/L

Example

If a water analysis reports 7 grains per gallon of Ca, how many mg/L of Ca are there?

From Table 1, conversion factor is 17.12 mg/L per grain

7 grains x 17.12 mg/L per grain = 120 mg/L

Water Quality Parameters

Acid

An acid can be defined as a compound that releases hydrogen ions (H^+) in a solution. All acids contain hydrogen. In general, acids have more or less a sour taste, they change litmus paper red, and they react with bases to form salts and water. An example is acetic acid (vinegar), which is considered a weak acid because it releases only small amounts of free hydrogen ions into solution. Sulfuric acid (H_2SO_4) is considered a strong acid because it releases more hydrogen ions into solution.

Bases

Bases are substances which can release hydroxyl (OH^-) ions. Bases change litmus paper blue. As is the case with acids, bases demonstrate varying degrees of ionization (H^- release). Those that ionize to a large extent are called strong bases, and those that ionize only slightly are known as weak bases.

Salts

A normal salt is a compound that is formed by the union of the cations (ions that carry a positive charge) of any base and anions (ions that carry a negative charge) of any acid. Other salts can be formed from acids or bases.

Electrical Conductivity

Electrical conductivity (EC) is a measure of the ability of water to pass an electrical current. Conductivity in water is affected by the presence of inorganic dissolved solids such as chloride, nitrate, sulfate, and phosphate anions or sodium, magnesium, calcium, iron, and aluminum cations. Organic compounds like oil, phenol, alcohol, and sugar do not conduct electrical current very well and therefore have a low conductivity when in water. Conductivity is also affected by temperature: the warmer the water, the higher the conductivity. For this reason, conductivity measurements are reported as conductivity at 25° C.

EC measurements are taken with platinum electrodes and presented in units of conductance. The drop in voltage caused by the resistance of the water is used to calculate

the conductivity per centimeter. The metric (SI) unit of measurement is deci-Siemens per meter (dS/m), which is equal in magnitude to the commonly used conductance term of millimho/cm (mmho/cm). Both of these terms are generally in the range of 0.1 to 5.0 for waters used for irrigation. Conductivity is also reported in units 1,000 times smaller: micromhos per centimeter (μmhos/cm) or microsiemens per centimeter (μS/cm). Most waters used for irrigation in Florida will be in the order 100 to 5,000 μS/cm. The conversion from electrical conductance to total dissolved solids (TDS) depends on the particular salts present in the solution. The conversion factor of 700 x EC (in dS/m) is applicable for converting EC values to TDS for Florida soil extracts and irrigation waters (Table 1). In other areas, conversion values of 630-640 x EC (dS/m) are commonly used.

As with TDS, a relationship exists between electrical con-ductivity and chloride concentration. Chloride estimates, however, are less reliable. The Florida norm is 140 with 90% of all ratios measuring less than 250. The following equation is an approximate conversion from EC to chloride.

Cl (mg/L) = EC (dS/m) x 140   Eq. 3

Example:

A Floridan well was sampled and tested with a conductivity meter, with a reading reported as 2.55 dS/m. Estimate the TDS and Cl concentrations in ppm.

2.55 dS/m = 2550 uS/m = 2550 umho/cm

2.55 dS/m x 700 ppm/dS/m = 1785 ppm TDS

Cl = dS/m x 140 = 2.55 x 140 = 357 ppm Cl

pH

The term pH is used to indicate the alkalinity or acidity of a substance as ranked on a scale from 1.0 to 14.0. The pH of water affects many chemical and biological processes in water. Different organisms flourish within different ranges of pH (Fig. 1). The largest variety of aquatic animals prefer a range of 6.5 - 8.0. When the pH is outside this
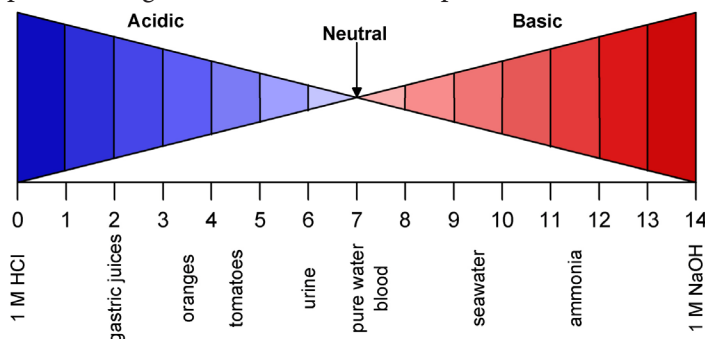

Figure 1. Range of pH for selected materials.

range, diversity within the water body may decrease due to physiological stresses and reduced reproduction. Low pH can also allow toxic elements and compounds to become more mobile and available for uptake by aquatic plants and animals. This can produce conditions that are toxic to aquatic life, particularly to sensitive species.

The pH scale measures the logarithmic concentration of hydrogen (H^+) and hydroxide (OH^-) ions, which make up water ($H^+ + OH^- = H_2O$). When both types of ions are in equal concentration, the pH is 7.0 or neutral. The pH value is the negative power to which 10 must be raised to equal the hydrogen ion concentration. Mathematically this is expressed as:

pH = -log [H^+]

Below 7.0, the water is acidic (there are more hydrogen ions than hydroxide ions). When the pH is above 7.0, the water is alkaline, or basic (there are more hydroxide ions than hydrogen ions). Since the scale is logarithmic, a drop in the pH by 1.0 unit is equivalent to a 10-fold increase in acidity. So, a water sample with a pH of 5.0 is 10 times as acidic as one with a pH of 6.0, and pH 4.0 is 100 times as acidic as pH 6.0. Generally, pH can be analyzed in the field or in the lab. If it is analyzed in the lab, pH must be measured within 2 hours of sample collection. The pH of a sample can change due to carbon dioxide from the air dissolving into the water.

A pH range of 6.0 to 9.0 appears to provide protection for the life of freshwater fish and bottom dwelling invertebrates in canals. Most resistant fish species can tolerate pH ranges from 4.0-10.0. At the extremes, fish eggs may hatch, but deformed young are often produced.

Alkalinity

Alkalinity is primarily determined by the presence of bicarbonates (HCO_3^-), carbonates (CO_3^-), and hydroxides (OH^-) in water. Alkalinity is a measure of the capacity of water to neutralize acids. Alkaline compounds in the water, such as bicarbonates (baking soda is one type), carbonates, and hydroxides, remove H^+ ions and lower the acidity of the water (which translates to increased pH). They usually do this by combining with the H^+ ions to make new com-pounds. Without this acid-neutralizing capacity, any acid added to a water source would cause an immediate change in the pH.

Total alkalinity is determined by measuring the amount of acid (e.g. muriatic acid) needed to bring the sample to a pH of 4.2. At this pH, all the alkaline compounds in the

sample are "used up" (see Fig. 2). The result is reported as milligrams per liter of calcium carbonate (mg/L of $CaCO_3$). At the pH of most irrigation water, alkalinity is primarily a measure of bicarbonate in the water. Alkalinity expressed as mg/L of $CaCO_3$ can be converted to an equivalent concentration of HCO_3^- by dividing by 0.82 (Table 1).
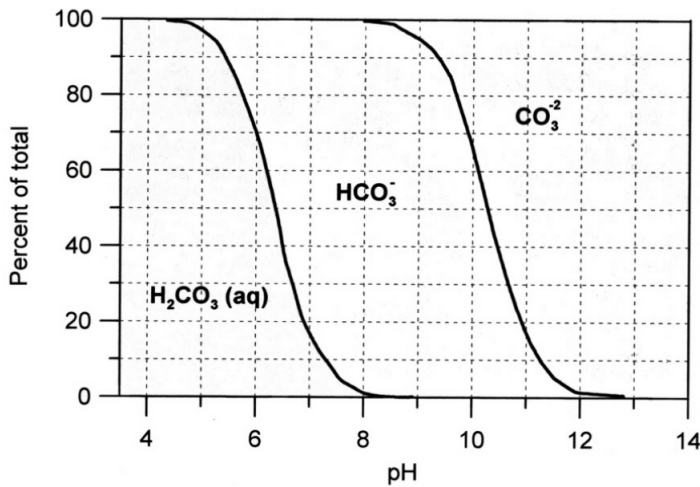

Figure 2. Percentages of total dissolved carbon species in solution for various pH ranges at 25°C and 1 atmosphere pressure.

Example:

Determine the equivalent expressed as concentration of HCO_3^- for an alkalinity of 80 mg/L $CaCO_3$.

80 mg/L x 0.82 = 65.6 mg/L as HCO_3^-

Hardness

The term "hardness" is one of the oldest terms used to describe characteristics of water. In fact, Hippocrates (450-354 B.C.) used the "hard" and "soft" terms in a discourse on water quality. He state "Consider the waters which the inhabitants use, whether they be marshy and soft, or hard and running from elevated and rocky situations ...". The hard term most likely referred to the condition of water which originated in the limestone formations of the upland regions. Over the years, hardness has come to be associated with the soap-consuming property of water or with the encrustations resulting from hard water when it is heated.

Hardness in water is caused primarily by calcium and magnesium, although iron and manganese also contribute to the actual hardness. Hardness may be divided into two types: carbonate and noncarbonate. Carbonate hardness is that portion of calcium and magnesium that can combine with bicarbonate to form calcium and magnesium carbonate. If the hardness exceeds the alkalinity (expressed as mg/L $CaCO_3$), the excess is termed noncarbonate hardness. The carbonate hardness is an indicator of the potential for calcium carbonate precipitation and scale formation. Total

hardness (carbonate and noncarbonate) is customarily expressed as equivalents of calcium carbonate ($CaCO_3$). Since the formula weight of $CaCO_3$ is near 100, hardness expressed in terms of mg/L of $CaCO_3$ can be converted to meq/L by dividing by 50. Hardness can be calculated from individual concentrations of calcium and magnesium using Equation 4.

$H = 2.5\ Ca + 4.1\ Mg$ ⟶ Eq. 4

where,

H = total hardness (mg/L as $CaCO_3$)

Ca = calcium concentration (mg/L of Ca)

Mg = magnesium concentration (mg/L of Mg)

Example:

Determine the hardness of a water sample that contains 120 mg/L of Ca and 45 mg/L of Mg.

H = 2.5 x 120 + 4.1 x 40 = 464 mg/L of hardness

464 mg/L / 17.12 = 27.1 grains/gal hardness

464 mg/L / 50 = 9.3 meq/L

The terms hard and soft as applied to water are inexact, and several attempts over the years have been made to add more descriptive qualifiers. Table 3 lists typical ranges for hardness designations.

Total Solids

Total solids are dissolved solids plus suspended and settleable solids in water. In stream water, dissolved solids consist of calcium, chloride, nitrate, phosphorus, iron, sulfur, other ions, and particles that will pass through a filter with pores of approximately 2 microns (0.002 cm) in size. Suspended solids include silt and clay particles, plankton, algae, fine organic debris, and other particulate matter. These are particles that will not pass through a 2-micron filter.

Suspended solids can serve as carriers for organic compounds. This is particularly of concern where the more water-insoluble pesticides are being used on irrigated crops. Pesticide concentrations associated with solids may increase well beyond those of the original application as water travels through ditches. High levels of solids can also clog drip and microsprinkler emitters and result in serious irrigation system maintenance problems.

Total solids (especially suspended solids) also affect water clarity. High levels of solids decrease the passage of light through water, thereby slowing photosynthesis of

submersed aquatic plants. Water also heats up more rapidly and holds more heat as total dissolved solids increase. This, in turn, may adversely affect aquatic life that has adapted to a lower temperature regime. Sources of total solids include: algae, decaying plant material, soil particles, fertilizers, and surface water runoff.

The measurement of total solids can be useful as an indicator of the effects of runoff from urban and agricultural areas. Concentrations often increase sharply during rainfall. Regular monitoring of total solids can help detect trends that might indicate increasing erosion within watersheds. Total solids are related closely to canal flow and velocity. Any change in total solids over time should be measured at the same site, and at the same flow rate.

Total solids are measured by weighing the amount of solids present in a known volume of sample. This is done by weighing a beaker, filling it with a known volume, evaporating the water in an oven and completely drying the residue, and then weighing the beaker with the residue. The total solids concentration is equal to the difference between the weight of the beaker with the residue and the weight of the beaker without it.

Primary Constituent Ions

Molecules may dissociate in solution into two or more ions which are theoretically free to move about independently. An ion differs from an atom or a molecule in that it carries an electrical charge. There are two kinds of ions: the cation, which is positively charged, and the anion, which is negatively charged. In water, the sum of the positive charges equals the sum of the negative charges, so that the solution remains electrically neutral.

Bicarbonate

Bicarbonate (HCO_3) is common in natural waters. Sodium and potassium bicarbonates can exist as solid salts; baking soda (sodium bicarbonate) is an example. Calcium and magnesium bicarbonates exist only in solution.

Biological/Biochemical Oxygen Demand

Organic materials in the water are decomposed by microorganisms, which use oxygen in the process. The amount of oxygen consumed by these organisms in breaking down the waste is known as the biochemical (or biological) oxygen demand or BOD. BOD also measures the chemical oxidation of inorganic matter (i.e., the extraction of oxygen from water via chemical reaction). Stormwater runoff can contribute large amounts of BOD to surface water systems.

A BOD test measures the amount of oxygen consumed by organisms during a specified period of time (usually 5 days at 20° C). The rate of oxygen consumption in a stream is affected by a number of variables: temperature, pH, the presence of certain kinds of microorganisms, and the type of organic and inorganic material in the water. BOD directly affects the amount of dissolved oxygen in surface waters. The greater the BOD, the more rapidly oxygen is depleted, resulting in less oxygen available to higher forms of aquatic life. The consequences of high BOD are the same as those for low dissolved oxygen: aquatic organisms become stressed, suffocate, and die.

Sources of BOD include leaves and woody debris; dead plants and animals; animal manure; effluents from pulp and paper mills, wastewater treatment plants, feedlots, and food processing plants; failing septic systems; and urban stormwater runoff. BOD is affected by the same factors that affect dissolved oxygen (see above). Aeration of stream water by rapids and waterfalls, for example, will accelerate the decomposition of organic and inorganic material. Therefore, BOD levels at a sampling site with slower, deeper waters might be higher for a given volume of organic and inorganic material than the levels for a similar site in highly aerated waters. Chlorine can also affect BOD measurement by inhibiting or killing the microorganisms that decompose the organic and inorganic matter in a sample.

BOD measurement requires taking two samples at each site. One is tested immediately for dissolved oxygen, and the second is incubated in the dark at 68°F (20°C) for 5 days and then tested for the amount of dissolved oxygen remaining. The difference in oxygen levels between the first test and the second test, in milligrams per liter (mg/L), is the amount of BOD. This represents the amount of oxygen required by microorganisms to break down the organic matter present in the sample bottle during the incubation period. Because of the 5-day incubation, the tests should be conducted in a laboratory.

Boron

Boron (B) occurs in natural waters. A small amount of boron is essential for plant growth, but a concentration slightly above the optimum is toxic to plants. Some plants are more sensitive to a boron excess than others.

Calcium

Calcium (Ca) is found to some extent in all natural waters and is commonly a large component of the constituent ions in Florida water. Calcium combines with carbonate to form calcium carbonate scale, which can plug microirrigation

system emitters. Calcium concentration is sometimes expressed as calcium hardness. To determine calcium concentration (mg/L), multiply calcium hardness by 0.4.

Dissolved Oxygen

Surface water systems both produce and consume oxygen. Oxygen is gained from the atmosphere and from plants as a result of photosynthesis. Moving water, because of its churning, dissolves more oxygen than still water. Respiration by aquatic animals, decomposition of organic matter, and various chemical reactions consume oxygen.

Oxygen is measured in its dissolved form as dissolved oxygen (DO). If more oxygen is consumed than is produced, dissolved oxygen levels decline and some sensitive animals may move away, weaken, or die. DO levels fluctuate seasonally, daily, and with water temperature since cold water holds more oxygen than warm water). Aquatic animals are most vulnerable to lowered DO levels in the early morning on hot summer days when stream flows are low, water temperatures are high, and aquatic plants have not been producing oxygen since sunset.
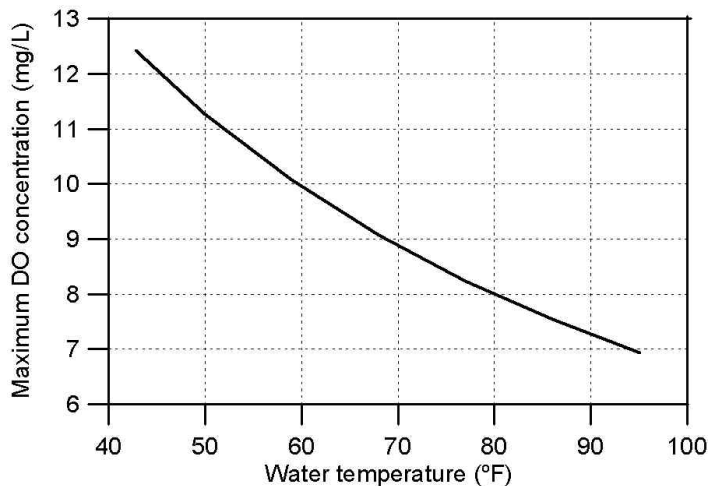

Figure 3. Relationship between water temperature and the maximum DO concentration possible in the solution.

In contrast to lakes, where DO levels are most likely to vary vertically in the water column, the DO in canals changes more horizontally along the course of the waterway. This is especially true in smaller, shallower ditches. The DO levels below spillways are typically higher than those in pools and slower-moving stretches. A time profile of DO levels at a sampling site is a valuable set of data because it shows the change in DO levels from the low point just before sunrise to the high point sometime at midday. Therefore it is important to note the time of DO sampling and to sample at similar times in order make meaningful comparisons.

DO is measured either in mg/L or percent saturation. Milligrams per liter is the amount of oxygen in a liter of water. Percent saturation is the amount of oxygen in a liter of water relative to the total amount of oxygen that the water can hold at that temperature. DO samples are collected using a special BOD bottle: a glass bottle with a turtleneck and a ground glass stopper. You can fill the bottle directly in the stream, or you can use a sampler that is dropped from a bridge or structure deep enough to submerge the sampler. Samplers can be made or purchased. The sample bottle should be submerged and allowed to fill without allowing air to mix with the sample (Fig. 4). The bottle should be completely filled and held submerged until the cap is firmly in place.
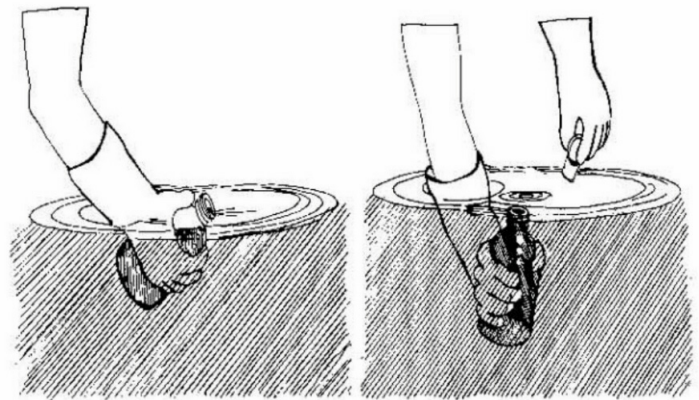

Figure 4. Proper sampling for DO involves submerging the sample bottle, allowing to gradually fill, and capping while under water.

Dissolved oxygen field kits using reagents and color wheels are relatively inexpensive. Field kits run between $35 and $200, and each kit comes with enough reagents to run 50 to 100 DO tests. Replacement reagents are inexpensive, and can be bought already measured out for each test in plastic pillows. If a high degree of accuracy and precision is needed for DO results, a digital titrator can be used. A kit that uses an eye-dropper-type or syringe-type titrator is suitable for most other purposes. The lower cost of this type of DO field kit might be attractive if you are relying on several teams of volunteers to sample multiple sites at the same time. Dissolved oxygen meters allow quick measurements, but they cost between $500 and $1,200, including a long cable to connect the probe to the meter. The advantage of a meter/probe is that you can measure DO and temperature quickly anywhere you can reach with the probe. You can also measure the DO levels at a certain point on a continuous basis. The results are read directly as milligrams per liter.

Adequate dissolved oxygen is necessary for good water quality. Oxygen is a necessary element to all forms of life. Natural stream purification processes require adequate

oxygen levels in order to provide for aerobic life forms. As dissolved oxygen levels in water drop below 5.0 mg/L, many forms of aquatic life are put under stress. The lower the concentration, the greater the stress. Oxygen levels that remain below 1mg/L for several hours can result in large fish kills.

Total dissolved gas concentrations in water should not exceed 110 percent. Concentrations above this level can be harmful to aquatic life. Fish in waters containing excessive dissolved gases may suffer from gas bubble disease. However, this is a very rare occurrence. The bubbles (emboli) block the flow of blood through blood vessels causing death. External bubbles (emphysema) can also occur and be seen on fins, on skin and on other tissue. Aquatic invertebrates are also affected by gas bubble disease but at levels higher than those lethal to fish.

Carbonate

Carbonate (CO_3) is found in some waters at high pH (>8.0). At normal pH levels alkalinity is primarily in the form of bicarbonate.

Chloride

Chloride (Cl) is found in most natural waters. In high concentrations it is toxic to some plants. All common chlorides are soluble and contribute to the total salt content (salinity) of soils. The chloride content should be determined to properly evaluate irrigation waters if TDS is greater than 1000 mg/L. Chlorides should be below 300 mg/L to avoid damage to citrus.

Iron

Iron is the fourth most abundant element, by weight, in the earth's crust. Iron in groundwater is normally present in the soluble ferrous (Fe^{++}) form. It is easily oxidized to the insoluble ferric (Fe^{+++}) state upon exposure to air.

Iron in water may be present in varying quantities depending upon the geological area and other chemical components of the water source. Ferrous (Fe^{++}) and ferric (Fe^{+++}) ions are the primary forms of concern in the aquatic environment. Other forms may be in either organic or inorganic wastewater streams. The ferrous form can persist in water void of dissolved oxygen and usually originates from groundwater wells. Black or brown swamp waters may contain iron concentrations of several mg/L in the presence or absence of dissolved oxygen, but this iron form has little effect on aquatic life. The current aquatic life standard is 1.0 mg/L based on toxic effects.

Iron in a soluble (ferrous) form may create emitter clogging problems at concentrations as low as 0.3 mg/L. Dissolved iron may precipitate out of the water due to changes in temperature or pressure, in response to a rise in pH, exposure to air, or through the action of bacteria. The presence of iron bacteria often results in the formation of an ochre sludge or slime mass capable of plugging the entire microirrigation system.

Magnesium

Magnesium (Mg) is also usually found in measurable amounts. Magnesium behaves much like calcium, but precipitates at higher pH levels and is not typically a problem in microirrigation systems. To determine magnesium concentration (mg/L), multiply Mg hardness by 0.24. Often laboratories will not separate calcium and magnesium but will report simply Ca + Mg in meq/L (as Hardness).

Manganese

Manganese (Mn) occurs in groundwater but is less common than iron, and it is generally in smaller amounts. Like iron, manganese in solution may precipitate out as a result of chemical or biological activity, forming a sediment which will clog emitters and other system components. The color of the deposits ranges from dark brown, if there is a mixture of iron, to black if the manganese oxide is pure.

Nitrogen

Nitrogen (N) is one of the most abundant elements. About 80 percent of the air we breathe is nitrogen. It is found in the cells of all living things and is a major component of proteins. Inorganic nitrogen may exist in the free state as a gas (N_2), or as nitrate (NO_3^-), nitrite (NO_2^-), or ammonia (NH_3^+). Organic nitrogen is found in proteins and is continually recycled by plants and animals.

Nitrogen-containing compounds act as nutrients in streams and rivers. Nitrate reactions in water can cause oxygen depletion. Thus, aquatic organisms depending on the supply of oxygen in the stream may be at risk. Bacteria in water quickly convert NO_2^- to NO_3^-. Together with phosphorus, nitrates in excess amounts can accelerate eutrophication, causing dramatic increases in aquatic plant growth and changes in the types of plants and animals that live in the stream. This, in turn, affects dissolved oxygen, temperature, and other indicators. Excess nitrates can cause hypoxia (low levels of dissolved oxygen) and can become toxic to warm-blooded animals at higher concentrations (10 mg/L or higher) under certain conditions.

Nitrites can produce a serious condition in fish called "brown-blood disease." Nitrites also react directly with hemoglobin in human blood and other warm-blooded animals to produce methemoglobin, which destroys the ability of red blood cells to transport oxygen. Nitrite-nitrogen levels below 90 mg/L and nitrate-nitrogen levels below 0.5 mg/L seem to have no effect on warm water fish.

Analysis of water that is polluted with nitrogen-rich organic matter might show low nitrate levels. Decomposition of the organic matter lowers the dissolved oxygen level, which in turn slows the rate at which ammonia is oxidized to NO_2^- and then to NO_3^-. Under such circumstances, it might be necessary to also monitor for nitrites or ammonia, which are considerably more toxic to aquatic life than nitrate.

Nitrate is not commonly found in large amounts in Florida waters. While beneficial as a plant nutrient, nitrate may indicate contamination from excessive use of fertilizers or from sewage. Nitrates have no effect on the physical properties of soil except to contribute slightly to its salinity, and nitrate is not harmful to irrigation systems.

Phosphorus

Phosphorus (P) is one of the key elements necessary for growth of plants and animals. Phosphorus reactions in water are complicated. In nature, phosphorus usually exists as part of a phosphate molecule (PO_4^{---}). Phosphates exist in three forms: orthophosphate (inorganic phosphorus), metaphosphate (or polyphosphate) and organically-bound phosphate. Each compound contains phosphorus in a different chemical arrangement. Phosphate that is not associated with organic material is considered inorganic. Inorganic phosphorus is the form required by plants. Animals can use either organic or inorganic phosphate.

Orthophosphates are produced by natural processes. Polyforms are used for treating boiler waters and in detergents. In water solutions, they change into the orthoform. Organic phosphates are important in nature. Organic phosphate consists of a phosphate molecule associated with carbon molecules, as in plant or animal tissue. Their occurrence may also result from the breakdown of organic pesticides which contain phosphates. They may exist in solution, as particles, loose fragments, or in the bodies of aquatic organisms. Both organic and inorganic phosphorus can either be dissolved in the water or suspended (attached to particles in the water column).

Rainfall can wash phosphates from grove soils into drainage ditches and canals. Since phosphorus is the nutrient in short supply in most fresh waters, even a modest increase in phosphorus can, under the right conditions, set off a whole chain of undesirable events including: accelerated plant growth, algae blooms, low dissolved oxygen levels, and the death of certain fish, invertebrates, and other aquatic animals. Respiration from excessive growth of algae and aquatic plant populations consume large amounts of oxygen, possibly making the water unsuitable for other forms of life.

Phosphorus cycles through the environment, changing form as it does so (Fig. 5). Aquatic plants take in dissolved inorganic phosphorus and convert it to organic phosphorus as it becomes part of their tissues. Animals obtain the organic phosphorus they need by eating either aquatic plants or animals.
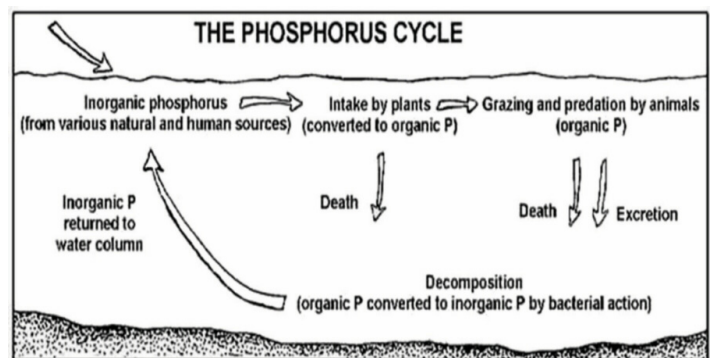

Figure 5. Phosphorus changes as it cycles through the aquatic environment.

As plants and animals excrete solid wastes or die, the organic phosphorus they contain sinks to the bottom, where bacterial decomposition converts it back to inorganic phosphorus, both dissolved and attached to particles. This inorganic phosphorus gets back into the water column when the bottom is stirred up by animals, human activity, chemical interactions, or water currents. Then it is taken up by plants and the cycle begins again.

In a watershed system, the phosphorus cycle tends to move phosphorus downstream as the current carries decomposing plant and animal tissue and dissolved phosphorus. Phosphorus becomes stationary only when it is taken up by plants or is bound to particles that settle to the bottom. In solution, the phosphate form depends on the pH (Fig. 6). The term "orthophosphate" is a chemistry-based term that refers to the phosphate molecule all by itself. "Reactive phosphorus" is a corresponding method-based term that describes what you are actually measuring when you perform the test for orthophosphate. Because the lab procedure isn't quite perfect, you get mostly orthophosphate but you also get a small fraction of some other forms. More complex inorganic phosphate compounds are referred to

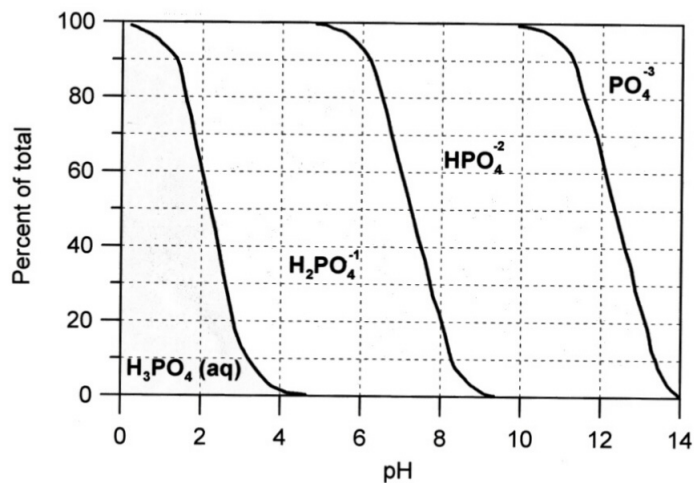as condensed phosphates or polyphosphates. The method-based term for these forms is acid hydrolyzable.

Figure 6. Percentages of total dissolved phosphate species in solution as a function of pH at 25°C and one atmosphere pressure.

Monitoring phosphorus is challenging because it involves measuring very low concentrations down to 0.01 mg/L or even lower. Even such very low concentrations of phosphorus can have a dramatic impact on streams. Less sensitive methods should be used only to identify serious problem areas. While there are many tests for phosphorus, only the following are likely to be performed by growers in the field.

- The total orthophosphate test is largely a measure of orthophosphate. Because the sample is not filtered, the procedure measures both dissolved and suspended orthophosphate. The EPA-approved method for measuring total orthophosphate is known as the ascorbic acid method. Briefly, a reagent (either liquid or powder) containing ascorbic acid and ammonium molybdate reacts with orthophosphate in the sample to form a blue compound. The intensity of the blue color is directly proportional to the amount of orthophosphate in the water.

- The total phosphorus test measures all the forms of phosphorus in the sample (orthophosphate, condensed phosphate, and organic phosphate). This is accomplished by first digesting (heating and acidifying) the sample to convert all the other forms to orthophosphate. Orthophosphate is then measured by the ascorbic acid method. Because the sample is not filtered, the procedure measures both dissolved and suspended orthophosphate.

- The dissolved phosphorus test measures that fraction of the total phosphorus which is in solution in the water (as opposed to being attached to suspended particles). It is determined by first filtering the sample, then analyzing the filtered sample for total phosphorus. Insoluble

phosphorus is calculated by subtracting the dissolved phosphorus result from the total phosphorus result.

All these tests have one thing in common, they all depend on measuring orthophosphate. The total orthophosphate test measures the orthophosphate that is already present in the sample. The others measure that which is already present, and that which is formed when the other forms of phosphorus are converted to orthophosphate by digestion.

Sample containers made of either plastic or glass are acceptable, depending on the parameters of concern. Because phosphorus molecules have a tendency to adsorb (attach) to the inside surface of sample containers, if containers are to be reused they must be acid-washed to remove adsorbed phosphorus. Therefore, the container must be able to withstand repeated contact with hydrochloric acid. Plastic containers, either high-density polyethylene or polypropylene, might be preferable to glass from a practical standpoint because they will better withstand breakage.

The only form of phosphorus recommended for field analysis is total orthophosphate, which uses the ascorbic acid method on an untreated sample. Analysis of any of the other forms involves adding potentially hazardous reagents and heating the sample to boiling, which requires too much time and equipment to be practical. In addition, analysis for other forms of phosphorus is prone to errors and inaccuracies in a field situation. Pretreatment and analysis for these other forms should be handled in a laboratory.

Potassium

Potassium (K) is usually found in lesser amounts in natural waters. It behaves much like sodium in the water, but it is a major plant nutrient.

Sodium

Sodium (Na) salts are all very soluble and as a result are found in most natural waters. High sodium in the irrigation water can impact both the soil and the plant. A soil with a large amount of sodium associated with a clay fraction has poor physical properties for plant growth and water infiltration. At high concentrations, Na can also be toxic to many plants. Sodium does not cause problems with the irrigation system. In order to prevent scale formation, domestic water supplies are often softened by replacing calcium with sodium.

Sulfate/Sulfide

Sulfate (SO_4) is abundant in nature. Sodium, magnesium, and potassium sulfates are readily soluble in water. Sulfate

has no characteristic action on the soil except to contribute to the total salt content. The presence of soluble calcium will limit sulfate solubility.

Groundwater that contains dissolved hydrogen sulfide gas (H_2S) is easily recognized by its "rotten egg" odor. As little as 0.5 ppm is noticeable. If the irrigation water contains more than 0.1 ppm of total sulfides, sulfur bacteria may grow within the irrigation system, forming masses of slime which may clog filters and emitters.

References

Hem, J.D. 1970. Study and interpretation of the characteristics of natural waters. USGS Water Supply Paper 1473. US GPO, Washington D.C.

Table 1. Conversion factors for water quality data.

To convert-	To-	Multiply by-
grains/gallon	mg/L	17.12
mg/L	grains/gallon	0.0584
dS/m	uS/cm	1000
dS/m	ppm	700
uS/cm	ppm	0.7
ppm	uS/cm	1.429
ppm	dS/m	0.001429
Ca^{++}	$CaCO_3$	2.5
$CaCl_2$	$CaCO_3$	0.9
HCO_3^-	$CaCO_3$	0.82
HCO_3^-	CO_3	0.49
Mg^{++}	$CaCO_3$	4.12
Na_2CO_3	$CaCO_3$	0.94
NO_3^-	N	0.23
N	NO_3^-	4.43

Table 2. Equivalent weights of common ions in irrigation water.

Cation (+ charge)	Equivalent weight	Anions (-charge)	Equivalent weight
Ammonium (NH_4^+)	18	Carbonate ($CO3^{--}$)	30
Calcium (Ca^{++})	20	Chloride (Cl^-)	35.5
Copper (Cu^{++})	31.8	Hydroxide (OH^-)	17
Magnesium (Mg^{++})	12.2	Nitrate (NO_3^-)	62
Manganese (Mn^{++})	27.5	Nitrate (NO_2^-)	46
Sodium (Na^+)	23	Sulfate (SO_4^-)	48
Iron (Fe^{++})	27.9	Bicarbonate (HCO_3^{--})	61
Iron (Fe^{+++})	18.6	Phosphate ($PO_4{---}$)	31.7
Potassium (K^+)	39.1	Phosphate (HPO_3^{--})	48
Sodium (Na^+)	23	Phosphate ($H2PO_4^-$)	97
Zinc (Zn^{++})	32.7	Sulfide (S^{--})	16

Table 3. Typical hardness ranges.

Hardness range		Descriptor
grains $CaCO_3$	mg/L of $CaCO_3$	
0-3.5	0-60	Soft
3.5-7.0	61-120	Moderately hard
7.1-10.5	121-180	Hard
>10.5	> 180	Very hard