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Fact sheet

Management of Iron in Irrigation Water

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Introduction

Iron is a common trace element in soils and groundwater. Iron is the fourth most abundant mineral in the earth's crust. The bulk iron content of soils is typically in the range of 0.5% to 5% (by volume), and is dependent upon the source rocks from which the soil was derived, transport mechanisms, and overall geochemical history. Iron occurs naturally in water in **soluble form** as **ferrous iron** (bivalent iron: **Fe^{+2}**) or **non-soluble form** as **ferric iron** (trivalent iron: **Fe^{+3}**).

During colonial times, bog iron was mined from bogs, streams, and waterways in the New Jersey Pine Barrens (Pinelands). The Pine Barrens include portions of seven counties: Ocean County, Burlington, Gloucester, Atlantic, Cumberland, Cape May, and Camden County. The soils are generally sandy and acidic. Drained water laden with organic acids (mainly humic and fulvic acids) from decaying vegetation percolates down to iron-rich clays underlaying much of the Pine Barrens, and in the process leaching out the soluble iron. Iron in groundwater quickly oxidizes to a reddish-brown product (hydrated iron oxide) when exposed to air. **Iron** is a common water contaminant that is not considered a health hazard; however, its presence **at elevated levels can cause aesthetic problems on ornamental plants, buildings and structures, and its accumulation on irrigation equipment can lead to clogged emitters (Figure 1 and 2)**.



Fig. 1. Iron rust deposits on plant leaves.



Fig. 2. Rusting of metal structures.

Iron Forms and Levels

Iron can be present in a water supply in many different forms (soluble, chelated, organic, and precipitated) and may or may not be apparent to the eye. These forms include ferrous (Fe^{+2}) or dissolved iron, which is invisible, while the ferric (Fe^{+3}) or oxidized (rusted) iron becomes apparent through precipitation, and usually appears as brownish red colored particles suspended in the water.

Irrigation water with iron levels above 0.1 ppm may cause clogging of drip irrigation emitters and above 0.3 ppm may lead to iron rust stains, and discoloration on foliage plants in overhead irrigation applications. These levels are generally below the levels that cause toxicities in plant tissue except when iron levels exceed 4 ppm or when the root medium pH is below 5.5.

Iron fixing bacteria, mainly from the **filamentous** genera such as *Gallionella* spp., *Leptothrix* and *Sphaerotilus* and less from the **rod type**, such as *Psendomonas* and

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Enterobacter, react with soluble iron, Fe^{+2}, through an oxidation process that changes the iron to an insoluble form, Fe^{+3}. When the ferric iron is surrounded by bacteria colonies, a bluish bronze sheen, sticky iron slime gel is created. The bacteria keeps the iron in the water from settling out and so when irrigation is applied to plants so is the bluish iron deposit. In propagation houses one often observes a slimy yellowish mass that is responsible for clogging the irrigation drippers or nozzles.

Management Methods to Control Iron in Irrigation Water

Depth of irrigation intake. Nursery growers can reduce the problem of iron deposits by making sure that their irrigation intakes are located 18 to 30 inches below the surface of the water. Intakes too close to the bottom pull settled iron sediment off the bottom of the pond. Those too close to the surface pull more of the oxidized form and other organisms that flourish on iron such as iron fixing bacteria. Treatment of iron depends on the form (soluble, chelated, organic, or precipitated) in which it occurs in the untreated water. Therefore, water testing is needed before considering or selecting the appropriate treatment equipment for effective removal of iron.

Sampling and testing irrigation water. Before implementing any control method, growers should draw water samples from the well or the pump into clean polyethylene bottles, filled completely, and closed tightly to avoid air oxidation and send them to a water analysis laboratory for complete analysis. For a list of certified water testing laboratories in each county in New Jersey visit the following website address: *www.rcre.rutgers.edu/pubs/publication.asp?pid=FS343*. A complete analysis of a water sample (including total hardness, pH, iron, turbidity, color, and taste or odor) and initial observation when the sample was drawn, will give an indication of the presence of iron. If a freshly drawn water sample appears clear, this does not mean that the water contains no iron. This is because **iron that is not visible to the eye can be present as ferrous bicarbonate ($Fe(HCO_3)_2$)**. However, during sampling and by the time the water sample reaches the laboratory, oxidation of some or all iron can occur and turbidity may show up in the results. Ferrous bicarbonate, when oxidized, changes into ferric hydroxide [$Fe(OH)_3$] producing carbon dioxide and lowering the pH.

Softeners to remove ferrous bicarbonate in water. The simplest method to remove ferrous bicarbonate iron from the water is to pass it through an **air tight water softener** containing a resinous cation exchanger: an insoluble matrix normally in the form of small (1–2mm diameter) beads, fabricated from an organic polymer substrate with a surface that easily traps and releases ions in a process called **ion exchange**. The capacity for removing iron depends on the capacity of resin. By using a basic softener regenerated with sodium chloride, iron can be removed along with calcium and magnesium.

Oxidation followed by filtration to remove high iron water content. If well water has a high iron content causing problems, then consider using a **basin aeration pump**. This pump keeps the water volume moving (creating ripples), resulting in iron precipitation due to **oxidation**. Removal of the precipitated iron can be achieved by **filtration**. In aerated water, the redox potential of the water is such that it allows oxidation of the ferrous iron into ferric iron, which precipitates into iron hydroxide, $Fe(OH)_3$, thus allowing a natural removal of the dissolved iron.

$$Fe^{+2} \xrightarrow{\text{oxidation}} Fe^{+3} \xrightarrow{\text{precipitation}} Fe(OH)_3$$

The time required for ferrous iron to undergo oxidation to the ferric state is dependent on many factors, the dominant being: pH; temperature; dissolved oxygen level; and the presence of other soluble ions. The lower the pH and temperature the longer the time required for completion of the oxidation reaction. Increasing dissolved oxygen decreases the time required for oxidation. For example:

- At pH 7.0, 90% Fe^{+2} oxidation requires 1 hour at 21°C and 10 hours at 5°C.
- At pH 8.0, 90% Fe^{+2} oxidation occurs in 30 seconds.
- At pH 6.0 it requires 100 hours.
- The critical dissolved oxygen concentration is 2 ppm. Below that ferrous iron oxidation occurs very slowly.

Allowing complete aeration of the water, then passing the aerated water through a **neutralizing filter (calcite filter)**, permits filtering out the suspended iron and raising the pH before the water is allowed to pass through a water softener. For this process, the typical filtration requirements are between 20 and 50 microns. **Oxidation and filtration** is usually the most economical method for iron removal in terms of operating costs because air is available for free. However, because large retention tanks may be required, this type of treatment may have higher capital costs. Oxidation followed by filtration is a relatively simple process.

Oxidants and oxidizing filters. Other methods of oxidation include the use of **oxidants** such as chlorine, chlorine dioxide, ozone, and potassium permanganate. Chlorination is widely used for oxidation of soluble divalent iron. Chlorine feed rates and contact time requirements can be determined by simple jar tests.

For a complete precipitation of iron, it is recommended to add a base to raise the pH. Iron precipitates more readily as the pH is raised above neutral. If aeration is limited by a pressure system and the pH of the water is above 6.8, an **oxidizing filter** can be an option (e.g. using **manganese greensand**). Greensand is a processed material consisting of nodular grains of the zeolite mineral glauconite. The material is coated with manganese oxide. The greensand will remove iron better, as manganese zeolite supplements the natural aeration of the water, helping to precipitate the iron.

Chlorination as an oxidizing agent to remove organic-iron rich water. **Organic iron** is a compound formed from an organic acid and iron. This form is often found in water with more than 2 ppm of dissolved organic carbons (DOC). **Tannins** are natural organics produced by vegetation, which stain water a tea-color. Organic iron is difficult to remove and therefore, it is important to note its presence in water. Organic iron and tannins can occur in very shallow wells, or wells being affected by surface water. A complete analysis of a water sample and initial observation when the sample was drawn will give an indication of their presence. For example, a sample of **water that contains organic iron** may be clear when drawn at the pump and the iron may not precipitate, but it **appears in a colloidal form**. Organic iron and tannins can slow or prevent iron oxidation, so water softeners, aeration systems, and iron filters may not work well. **Chemical oxidation followed by filtration may be an option**. **Chlorination** can be considered as a treatment method, especially when iron exists in organic form. Chlorination breaks down the organic complexes, and the iron then may be oxidized and precipitated by aeration and pH adjustment. Chlorinating irrigation water will result in a much faster oxidation rate and it can be injected in gaseous or liquid form. Commercial irrigation contractors can install these systems. For additional information on locating irrigation contractors, contact your local Rutgers Cooperative Research & Extension Office. Gaseous chlorine, injected from cylinders, is more effective and economical over the long run than the liquid form, but it is extremely dangerous when cylinders have to be changed, particularly if the cylinders are housed in a building. Liquid chlorine (5–15% sodium hypochlorite), a safer alternative, is injected using a variable ratio injector. Liquid chlorine losses strength over time and hence, the injection rate must be increased. It is recommended to inject liquid sodium hypochlorite continuously at a rate of 1 ppm for each 1 ppm of iron in the irrigation water. Mixing liquid sodium hypochlorite in water results in the formation of hypochlorous acid (HOCL) and hydroxyl ions (OH), a reaction that raises the pH of the water. The amount of HOCl that will be present in solution, and thus active, will be larger at lower pH levels (more acidic conditions). At pH 8, only about 22% of the chlorine injected will be in the active HOCl form, at pH 7, about 73% will be in the HOCl form, and at pH 6, about 96% will be in the HOCl form. Hypochlorous acid reacts with iron in solution and oxidizes the ferrous iron into the ferric form. The ferric iron then becomes the insoluble ferric hydroxide as a precipitate. Chlorine should be injected before the filters so that these precipitates may be trapped in the filters. Chlorine may react with some metal and plastic components of irrigation systems. Therefore, always check with the manufacturer or supplier of system components to identify any potential problems before beginning a chlorine injection program. The water can be tested for free chlorine using an inexpensive D.P.D. (diethyl-phenylene-diamine) test kit. **A swimming pool kit** at the end of the irrigation line or riser can be used but it **should measure free chlorine**. Many pool test kits measure only total chlorine. Chlorination should be followed by filtration when organic complexes of iron are present in water. Caution should be exercised when chlorination is selected as a method for iron removal in irrigation water because some plants are sensitive to chlorine, naming a few: crapemyrtle, dogwood, hibiscus, hydrangea, juniper, rhododendron, rose, sugar maple, spruce, and viburnum.

Other products for sequestering iron in water. Examples known to the authors include: **Di-Solv®** (Flo-Tec Inc., Largo, Florida), a negatively charged compound can be added to irrigation water to eliminate mineral buildup, a nourishment source for bacterial growth. The use of Di-Solv® reduces plugged emitters due to biological contaminants. Another product by the same company is **Aqua-Solv®**, a chemical compound that sequesters iron, calcium, and manganese ions. Similar products may be available from other suppliers. Effective application depends upon equipment availability and concentration of iron or other problem mineral content.

Iron Testing Kits. Irrigation water can be tested for iron levels (0–10ppm) using a simple test kit. Examples known to these authors include an inexpensive, simple, easy, and safe CHEMets® colormetric iron test kit. The kit contains 30 individual tests, vaccum-sealed ampules, plus comparators. Similarly, a multi parameter test kit (hardness, pH, and iron) such as Hach® Model: HA-62A (100 tests) can

be used for testing iron in water in the range of 0–10ppm, whereas, Hach® Iron model IR-18 measures iron in water when iron concentration is below 5ppm. Other quick testing kits (48 individual tests) that measure ferrous iron concentration (test kit code 8052-310) or ferrous plus ferric forms (test kit code 8052-311) in the range of 0–10ppm can be found at the Analyticon Instruments Corporation® (Springfield, NJ), *www.analyticon.com/analyticon/products/laboratory/simplepack_testkit.htm.* La Motte® Model P-61 code 4447 is another test kit that can be used for testing iron (0.5–10.0ppm) in agricultural water supplies. Similar kits may be available from other suppliers.

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