Activity 7

Iron

Part 2

OBJECTIVES:

1. Using test strips, determine the amount of iron in tap water and in simulated mine runoff.
2. Demonstrate a process for removing iron contaminants from water.

VOCABULARY:

Precipitate – a solid reaction product that “falls out” of solution as it is formed

Flocculant – A chemical that causes tiny suspended particles to clump together

MATERIALS:

School Test Kit Materials:
Iron Test Strips
School Kit Color Chart and Instruction card
Plastic droppers
Data collections sheet

Equipment:
Clean Glass Jars with Lids
Heavy-Duty Paper Towels
Wide-Mouth Glass Container

Shopping list:
Iron nail, non-galvanized or non-coated, lightly sandpapered to remove rust
Distilled white vinegar (acetic acid, CH3COOH)
Lye (sodium hydroxide, NaOH - found in the drain cleaner section of the grocery store)
Alum (aluminum potassium sulfate, AlK(SO4)2 – found in the spice section of the grocery store)
Tap Water

SAMPLE PREPARATION:

__Iron solution__ – (simulated mine runoff) - Into a clean glass jar, add 100 ml distilled white vinegar and an iron nail. Leave overnight.

__Lye solution (NaOH)__ – WARNING Lye may cause severe burns - take the recommended precautions. A responsible adult should perform this step. Read the entire package label before opening and using. Mix 2 grams (approx. ½ teaspoon) of lye in 500 ml tap water. Alternatively, prepare a 10% NaOH solution in tap water.

__Alum solution__ – To create a saturated solution, add approx. 20 g alum to 500 ml of distilled water and mix thoroughly. Undissolved alum will remain in the bottle bottom of the container.

PROCEDURE:

1. In a clean glass jar, collect a tap water sample. Measure the pH of the water. Using the iron test strip, measure the level of iron in the water. Record pH and iron results in the data table.
2. Take approximately 50 ml of iron solution in a clean jar. Measure the pH of the solution. Using an iron test strip, measure the level of iron in the solution. Record pH and iron results in the data table. Describe the appearance of the solution.
3. Using the plastic dropper, carefully add drops of the sodium hydroxide solution into the iron solution and stir with a plastic spoon. Continue until a reddish brown precipitate forms or until the solution changes from clear to brownish. Record the appearance of the precipitate in the data table. Measure the pH of the solution containing the precipitate.
4. Add approximately 5 ml alum solution to the iron solution, and let sit for 5 to 10 minutes. Observe the gummy floc formation. Record the appearance in the data table.
5. Using a funnel and filter paper, or heavy-duty paper towel, filter the floc from the solution and collect the filtrate.
6. Using an iron test strip, measure the iron remaining in the filtrate. Record the results in the data table.

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Activity 7

Iron

Part 3

ANALYSIS AND APPLICATION

1. What happened to the iron in the solution between step 2 and step 6?
2. The treatment with sodium hydroxide and alum is similar to the process used by commercial water treatment systems to remove iron from water, but one more step is required. Observe the floc formation in step 4, and devise a procedure to remove the precipitated floc. What is in the floc? Suggest an environmentally friendly method of disposing of the waste from this experiment.

EXTENSION

1. Test for iron in clay. Compare iron measurements in reduced clay and oxidized clay samples.
2. Research natural treatment alternatives for the removal of heavy metals from mine drainage. Construct a working model.
3. If you have a mine drainage site in your community, develop a clean-up plan and adopt the site as a service project. Planning will require construction of a working model, budget planning, support of school administration, permission of the property owner, notification of any regulating authorities including the EPA, planning for multiyear support, and fund-raising.
4. Contact a local wastewater treatment plant. (Remember that many industries have private plants for treatment of water used during the manufacturing process.) Ask about their experiences in metals treatment.

Activity 8

Nitrates and Nitrites

Part 1

Nitrogen (N2) is a stable gas making up 79% of the air we breathe. All living things require nitrogen to build body tissues, especially proteins, enzymes, nucleic acids, hormones, and vitamins. Unfortunately, humans cannot use nitrogen (N2) in the pure state, but must get it by consuming plants and other animals. Most plants cannot use N2, but require nitrogen in the form of nitrates (NO3-) or ammonia (NH3). Natural mineral deposits contribute some nitrates to the soil. Bacteria make nitrates and ammonia available by breaking down nitrogen trapped in plant and animal products and tissues. Blue-green algae and nitrogen-fixing bacteria that grow in the roots of legumes offer the rare opportunity for nitrogen to leave the atmosphere and enter the biological system.

At first glance, the description of the nitrogen cycle above would make it appear seems to say that nitrogen available for living things is a scarce commodity. However, certain human activities result in the addition of add large amounts of nitrogen to the aquatic system that may overwhelm the natural cycle. Runoff from fertilizers, septic fields, sewage systems, and animal lots add nitrogen-rich waste water. Shrinking wildlife habitats may concentrate large numbers of ducks, geese and seagulls into the remaining space, their droppings making a significant impact on nitrogen levels.

Nitrogen is good! Right? Nitrogen fertilizers provide nutrients that are necessary for plant growth; that is why we add them to our lawns and crops. But when high levels of nitrates accumulate in aquatic environments, they may begin a chain of events that become deadly to life.

The most common warning sign of high nitrate levels are algae blooms and excessive plant growth. In the presence of sunlight, plants perform photosynthesis, producing oxygen and carbohydrates. Plants also respire, consuming oxygen and carbohydrates and giving off carbon dioxide. When the sun sets, photosynthesis ends, but respiration continues. Algae blooms can actually consume more oxygen during the night than they produce during the day, creating dangerously low oxygen levels at night. Making matters worse, algae blooms usually occur during warm weather, when water holds less oxygen, and fish, being more active, demand more. To top off the recipe for disaster, more plants living also means more plants dying. When the extra algae dies plants die, the bacteria that break the plants down also "bloom," using up oxygen themselves (aerobic bacteria). The result may be sudden, overnight fish kills. Oxygen levels may eventually drop so low that even the bacteria die, leaving an environment where the only survivors are bacteria that can live in the absence of oxygen (anaerobic bacteria). Those bacteria generate the foul-smelling hydrogen sulfide (H2S) gas that creates the rotten egg-like stench associated with dead and dying ponds. In this scenario – nitrogen was NOT GOOD. [The process described above is the result of eutrophication: see vocabulary].

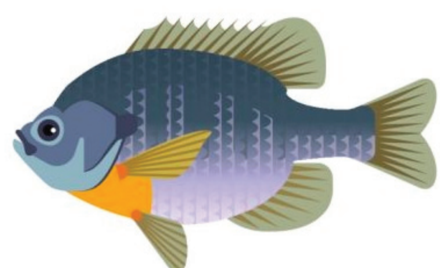
A short-lived but dangerous part of the nitrogen cycle is nitrite (NO2-). In the body, high levels of nitrites reduce the blood's ability to carry oxygen. This causes "brown blood disease" in fish and methemoglobinemia or "blue baby disease" in humans. Infants drinking formula made with water containing nitrites levels of 10 ppm or higher have died due to the inability of their blood to carry oxygen. Many of these poisoning cases have involved private wells where excess nitrogen from fertilizers or poorly functioning septic tanks has seeped into the ground water.

Many other conditions can interact with nitrate levels to affect the growth of plants, including light availability, temperature and pH. In addition, nutrient levels tend to have little effect in fast-moving water. As a result, high nutrient levels in a moving brook may not produce an algae bloom until the water reaches the stagnant pond.

In aquaria and ponds, nitrate levels may not be particularly toxic, but are a good measure of the general health of the ecosystem. They reflect a high level of biological activity that can lead to other problems: high ammonia levels, increasing acidity, decreasing buffering capacity (alkalinity), disease, and slow wound healing. Fish become stressed at levels above 10.0 ppm (as nitrogen).

Fish are many more times sensitive to nitrite. Although some fish are more sensitive than others, nitrite levels greater than 0.1ppm should be a concern. The best solution for elevated nitrate and nitrite levels in an aquarium is good maintenance and regular partial water changes.

For drinking water, the EPA recommends maximum levels of Total Nitrate = 10 ppm, Nitrite = 1 ppm, and Nitrate only = 10 ppm. Nitrate levels above 10 ppm are unsafe for drinking.



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