SECTION C

INVESTIGATING THE CAUSE OF THE FISH KILL

The challenge facing investigators of the Riverwood fish kill is to decide what in Snake River water was responsible for the crisis. In this section, you will learn about the process of dissolving, how solutions behave and are described, and what types of substances dissolve in water. What you learn will help to ensure that you have the knowledge and skills needed to evaluate the Riverwood data and to determine the cause of the fish kill.

C.1 SOLUBILITY OF SOLIDS

Could something dissolved in the Snake River have caused the fish kill? As you already know, a variety of substances can dissolve in natural waters. To determine whether any of these substances could be harmful to fish, you first need to know how solutions are formed and described. For example, how much of a certain solid substance will dissolve in a given amount of water?

Imagine preparing a water solution of potassium nitrate, KNO_3. What happens as you add a scoopful of solid, white potassium nitrate crystals to water in a beaker? As you stir the water, the solid crystals dissolve and disappear. The resulting solution remains colorless and clear. In this solution, water is the solvent, and potassium nitrate is the solute.

Remember: In a solution, the solvent is the dissolving agent and the solute is the dissolved substance.

What will happen if you add a second scoopful of potassium nitrate crystals to the beaker and stir? These crystals also may dissolve. However, if you continue adding potassium nitrate without adding more water, eventually some potassium nitrate crystals will remain undissolved on the bottom of the beaker, no matter how long you stir. The maximum quantity of a substance that will dissolve in a certain quantity of water (for example, 100 g) at a specified temperature is the **solubility** of that substance in water. In this example, the solubility of potassium nitrate might be expressed as "grams potassium nitrate per 100 g water" at a specified temperature.

Questions & Answers

From everyday experiences, you probably know that both the size of the solute crystals and the vigor and duration of stirring affect how long it takes for a sample of solute to dissolve at a given temperature. But do these factors affect how much substance will eventually dissolve? With enough time and stirring, will even more potassium nitrate dissolve in water? It turns out that the solubility of a substance in water is a characteristic of the substance and cannot be changed by any amount of stirring or time.

CD-ROM WWW. Dissolving Ionic Compounds

So what does affect the actual quantity of solute that dissolves in a given amount of solvent? As you can see from Figure 20 on page 46, the mass of

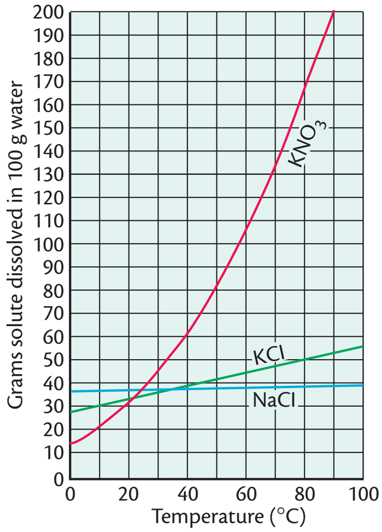


Figure 20 *Relationship between solute solubility in water and temperature.*

solid solute that will dissolve in 100 g water varies as the temperature of the water changes from 0 °C to 100 °C. The graphical representation of this relationship is called the solute's solubility curve.

Each point on the solubility curve indicates a solution in which the solvent contains as much dissolved solute as it normally can at that temperature. Such a solution is called a **saturated solution.** Thus each point on the solubility curve indicates a saturated solution. Look at the curve for potassium nitrate (KNO_3) in Figure 20. At 50 °C, how much potassium nitrate will dissolve in 100 g water to form a saturated solution? This value—80 g KNO_3 per 100 g water—is the solubility of potassium nitrate in 50 °C water. In contrast, the solubility of potassium nitrate in 20 °C water is only about 30 g KNO_3 per 100 g water. (Make sure that you are able to "read" this value on the graph.)

Note that the solubility curve for sodium chloride (NaCl) is nearly a horizontal line. What do you think this means about the solubility of sodium chloride as temperature changes? Compare the curve for sodium chloride with the curve for potassium nitrate (KNO_3), which rises steeply as temperature increases. You should be able to conclude that for some solutes, such as potassium nitrate (KNO_3), solubility in water is greatly affected by temperature, whereas for others, such as sodium chloride (NaCl), the change is only slight.

Now consider a solution containing 80 g potassium nitrate in 100 g water at 60 °C. Locate this point on the graph. Where does it fall with respect to the solubility curve? What does this tell you about the level of saturation of the solution? Because each point on the solubility curve represents a saturated solution, any point on a graph below a solubility curve must represent an unsaturated solution. An **unsaturated solution** is a solution that contains less dissolved solute than the solvent can normally hold at that temperature.

What would happen if you cooled this solution to 40 °C? (Follow the line representing 80 g to the left on the graph.) You might expect that some solid KNO_3 crystals would form and fall to the bottom of the beaker. In fact, this event is likely to occur. Sometimes, however, you can cool a saturated solution without forming any solid crystals, producing a solution that contains more solute than could usually be dissolved at that temperature. This type of solution is called a **supersaturated solution.** (Note that this new point lies above the solubility curve for potassium nitrate.) Agitating a supersaturated solution or adding a "seed" crystal to the solution often causes the "extra" solute to appear as solid crystals and settle to the bottom of the beaker, or precipitate. The remaining liquid then contains the amount of solute that represents a stable, saturated solution at that temperature.

One example of crystallization in a supersaturated solution may be familiar to you—the production of rock candy. A water solution is supersaturated with sugar. When seed crystals are added, they cause excess dissolved sugar to crystallize from the solution onto a string. Mineral deposits around a hot spring are another example of crystallization from a supersaturated solution. Water emerging from a hot spring is saturated with dissolved minerals. As the solution cools, it becomes supersaturated. The rocks over which the solution flows act as seed crystals, causing the formation of more mineral deposits.

SOLUBILITY AND SOLUBILITY CURVES

Building Skills 5

What is the solubility of potassium nitrate at 40 °C? The answer is found by using the solubility curve for potassium nitrate given in Figure 20. Locate the intersection of the potassium nitrate curve with the vertical line representing 40 °C. Follow the horizontal line to the left and read the value. The solubility of potassium nitrate in water at 40 °C is 60 g per 100 g water.

At what temperature will the solubility of potassium chloride be 25 g per 100 g water? Think of the space between 20 g and 30 g on the *y* axis in Figure 20 as divided into two equal parts, then follow an imaginary horizontal line at "25 g/100 g" to its intersection with the curve. Follow a vertical line down to the *x* axis. Because the line falls halfway between 10 °C and 20 °C, the desired temperature must be about 15 °C.

As you have seen, the solubility curve is quite useful when you are working with 100 g water. But what happens when you are working with other quantities of water? The solubility curve indicated that 60 g potassium nitrate will dissolve in 100 g water at 40 °C. How much potassium nitrate will dissolve in 150 g water at this temperature? You can "reason" the answer in the following way.

The amount of solvent (water) has increased from 100 g to 150 g—1.5 times as much solvent. That means that 1.5 times as much solute can be dissolved. Thus: $1.5 \times 60\text{ g} = 90\text{ g } KNO_3$.

The calculation can also be written as a simple proportion, which will give the same answer:

$$\frac{60\text{ g }KNO_3}{100\text{ g }H_2O} = \frac{x\text{ g }KNO_3}{150\text{ g }H_2O}$$

$$x\text{ g }KNO_3 = \frac{(60\text{ g }KNO_3)(150\ \cancel{\text{g }H_2O})}{(100\ \cancel{\text{g }H_2O})} = 90\text{ g }KNO_3$$

Refer to Figure 20 to answer the following questions.

1. a. What mass (in grams) of potassium nitrate (KNO_3) will dissolve in 100 g water at 60 °C?
   b. What mass (in grams) of potassium chloride (KCl) will dissolve in 100 g water at this temperature?

2. a. You dissolve 25 g potassium nitrate in 100 g water at 30 °C, producing an unsaturated solution. How much more potassium nitrate (in grams) must be added to form a saturated solution at 30 °C?
   b. What is the minimum mass (in grams) of 30 °C water needed to dissolve 25 g potassium nitrate?

3. a. A supersaturated solution of potassium nitrate is formed by adding 150 g KNO_3 to 100 g water, heating until the solute completely dissolves and then cooling the solution to 55 °C. If the solution is agitated, how much potassium nitrate will precipitate?
   b. How much 55 °C water would have to be added (to the original 100 g water) to just dissolve all of the KNO_3?

C.2 SOLUTION CONCENTRATION

The general terms saturated and unsaturated are not always adequate for describing the properties of solutions that contain different amounts of solute. A more precise description of the amount of solute in a solution is needed—an exact, numerical measure of concentration.

Solution concentration refers to the quantity of solute dissolved in a specific quantity of solvent or solution. You have already worked with one type of solution concentration expression: The water-solubility curves in Figure 20 (page 46) reported solution concentrations as the mass of a substance dissolved in a given mass of water.

Another way to express concentration is with percents. For example, dissolving 5 g table salt in 95 g water produces 100 g solution with a 5% salt concentration (by mass).

$$\frac{5\text{ g salt}}{100\text{ g solution}} \times 100\% = 5\%\text{ salt solution}$$

"Percent" means parts per hundred parts. So a 5% salt solution could also be reported as five parts per hundred of salt (5 pph salt). However, percent is much more commonly used.

$$\frac{5}{100} = \frac{50\,000}{1\,000\,000}$$

5% (5 pph) = 50 000 ppm

For solutions containing much smaller quantities of solute (as are found in many environmental water samples, including those from the Snake River), concentration units of **parts per million (ppm)** are sometimes useful. What is the concentration of the 5% salt solution expressed in ppm? Because 5% of 1 million is 50 000, a 5% salt solution is 50 000 parts per million.

Although you may not have realized it, the notion of concentration is part of daily life. For example, preparing beverages from concentrates, adding antifreeze to an automobile, and mixing pesticide or fertilizer solutions all require the use of solution concentrations. The following activity will help you review the concept of solution concentration, as well as gain experience with the chemist's use of this idea.

DESCRIBING SOLUTION CONCENTRATIONS

Building Skills 6

A common intravenous (abbreviated as IV) saline solution used in medicine contains 4.55 g NaCl dissolved in 495.45 g sterilized distilled water. Because a solution is a homogeneous mixture, the NaCl is distributed uniformly throughout the solution. What is the concentration of this solution, expressed as grams NaCl per 100 g solution?

The answer can be calculated in the following way:

$$\frac{4.55\text{ g NaCl}}{4.55\text{ g NaCl} + 495.45\text{ g water}} = \frac{4.55\text{ g NaCl}}{500\text{ g solution}}$$

If 500 g solution contains 4.55 g NaCl, then you can determine the answer by calculating how much NaCl is contained in 100 g solution. So 100 g (or 1/5) of the solution will contain 1/5 of the total solute. One-fifth of the total solute is 0.91 g NaCl. Thus 100 g solution contains 0.91 g NaCl and 99.09 g water—1/5 as much as in the full 500-g solution:

4.55 g NaCl × 1/5 = 0.91 g NaCl

$$\frac{0.91 \text{ g NaCl}}{0.91 \text{ g NaCl} + 99.09 \text{ g water}} = \frac{0.91 \text{ g NaCl}}{100 \text{ g solution}} = 0.91\% \text{ NaCl}$$

The concentration of this solution can be expressed as 0.91 g NaCl per 100 g solution. The solution is 0.91% NaCl by mass.

Now consider this example: One teaspoon of sucrose, which has a mass of 10 g, is dissolved in 240 g water. What is the concentration of the solution, expressed as grams sucrose per 100 g solution? As percent sucrose by mass?

Because the solution contains 10 g sucrose and 240 g water, it has a total mass of 250 g. A 100-g solution would contain 2/5 as much solute, or 4 g sucrose. Thus 100 g solution contains 4 g sucrose and 96 g water, a concentration of 4 g sucrose per 100 g solution. To determine the percent sucrose by mass,

$$\frac{10 \text{ g sucrose}}{250 \text{ g solution}} \times 100\% = \frac{4 \text{ g sucrose}}{100 \text{ g solution}} \times 100\% = 4\% \text{ sucrose by mass}$$

1. One teaspoon of sucrose is dissolved in a cup of water. Identify
   a. the solute.
   b. the solvent.

2. What is the concentration of each of the following solutions expressed as percent sucrose by mass?
   a. 17 g sucrose is dissolved in 183 g water.
   b. 30 g sucrose is dissolved in 300 g water.

3. A saturated solution of potassium chloride is prepared by adding 45 g KCl to 100 g water at 60 °C.
   a. What is the concentration of this solution?
   b. What would be the new concentration if 155 g water were added?

4. How would you prepare a "saturated solution" of potassium nitrate (KNO_3)?

Sucrose, $C_{12}H_{22}O_{11}$, is ordinary table sugar.

C.3 CONSTRUCTING A SOLUBILITY CURVE

Laboratory Activity

Introduction

You have seen and used solubility curves earlier in this unit (pages 46–47). In this activity, you will collect experimental data to construct a solubility curve for succinic acid ($C_4H_6O_4$), a molecular compound. Before you proceed, think about how your knowledge of solubility can help you gather data to construct a solubility curve.

- How can the properties of a saturated solution be used?
- What temperatures can you investigate?
- How many times should you repeat the procedure to be sure of your results?

Discuss these questions with your partner or laboratory group. Your teacher will then discuss with the class how data will be gathered and will demonstrate safe use of the equipment that will be used.

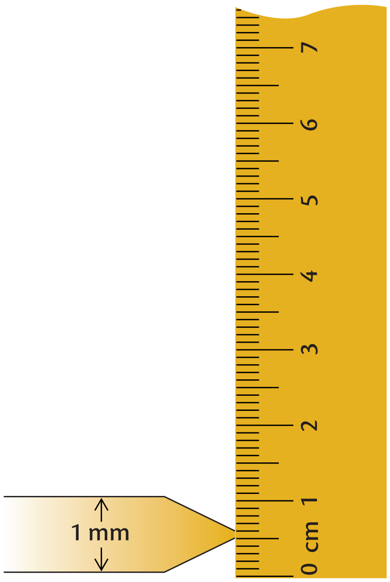
Safety

Keep the following precautions in mind while performing this laboratory procedure.

- The succinic acid that you will use is slightly toxic if ingested by mouth, so be sure to wash your hands thoroughly at the end of the laboratory.
- Never stir a liquid with a thermometer. Always use a stirring rod.
- Use insulated tongs or gloves to remove a hot beaker from a hot plate. Hot glass burns!
- Dispose of all wastes as directed by your teacher.

Procedure

1. To make a water bath, add approximately 300 mL water to a 400-mL beaker. Heat the beaker, with stirring, to either 45 °C, 55 °C, or 65 °C, as agreed to in your pre-lab class discussion. Ensure that the student team sharing your hot plate is investigating the same temperature. Carefully remove the beaker (using gloves or beaker tongs) when it reaches the desired temperature. NOTE: Do not allow the water-bath temperature to rise more than five degrees above the temperature that you have chosen. Return the beaker to the hot plate as needed to maintain the appropriate water-bath temperature.
2. Place between 4 g and 5 g succinic acid in each of two test tubes. **CAUTION:** *Be careful not to spill any of the succinic acid. If you do, clean up and dispose of the succinic acid as directed by your teacher.* Add 20.0 mL distilled water to each test tube.
3. Place each test tube in the water bath and take turns stirring the succinic acid solution with a glass stirring rod every 30 seconds for 7 minutes. Each minute, place the thermometer in the test tube and monitor the temperature of the succinic acid solution, ensuring that it is within 2 °C of the temperature that you have chosen.
4. At the end of 7 minutes, carefully decant the clear liquid from each test tube into a separate, empty test tube, as demonstrated by your teacher.
5. Carefully pour the hot water from the beaker into the sink and fill the beaker with water and ice.
6. Place the two test tubes containing the clear liquid in the ice bath for 2 minutes. Stir the liquid in each test tube gently once or twice. Remove the test tubes from the ice water. Allow the test tubes to sit at room temperature for 5 minutes. Observe each test tube carefully during that time. Record your observations.
7. Tap the side of each test tube and swirl the liquid once or twice to cause the crystals to settle evenly on the bottom of the test tubes.
8. Measure the height of crystals collected (in millimeters, mm). Have your partner(s) measure the crystal sample height and compare your results. Report the average crystal height for your two test tubes to your teacher.



A millimeter is 1/10th of a centimeter.

9. Rinse the succinic acid crystals from the test tubes into a collection beaker designated by your teacher. Make sure that your laboratory area is clean.
10. Wash your hands thoroughly before leaving the laboratory.

Data Analysis

1. Find the mean crystal height obtained by your class for each temperature reported.
2. Plot the mean crystal height in millimeters (y axis) versus the water temperature in degrees Celsius (x axis).

Questions

1. Why is it important to collect data from more than one trial at a particular temperature?
2. How did you make use of the properties of a saturated solution at different temperatures?
3. Did all the succinic acid that originally dissolved in the water crystallize out of the solution? Explain your answer.
4. Given the pooled class data, did you have enough points to make a reliable solubility curve for succinic acid? Would the curve be good enough to make useful predictions about succinic acid solubility at temperatures not investigated in this activity? Explain your answer.
5. What procedures in this activity could lead to errors? How would each error affect your data?
6. Using your knowledge of solubility, propose a different procedure for gathering data to construct a solubility curve.

C.4 DISSOLVING IONIC COMPOUNDS

You have just investigated the process of a compound dissolving in water. What you observed is called a macroscopic phenomenon. However, what chemistry is primarily concerned with is what happens at the submicroscopic level—atomic and molecular phenomena that are not easily observed. As you have seen, temperature, agitation, and time all contribute to dissolving a solid material. But how do the atoms and molecules of the solute and solvent interact to make this happen?

Experiments suggest that water molecules are electrically **polar.** Although the entire water molecule is electrically neutral, the electrons are not evenly distributed in its structure. A polar molecule has an uneven distribution of electrical charge, which means that each molecule has a positive region on one end and a negative region on the other end. Evidence also suggests that a water molecule has a bent or V-shape, as illustrated in Figure 21 on page 52, rather than a linear, sticklike shape as in H–O–H. The "oxygen end" is an electrically negative region that has a greater

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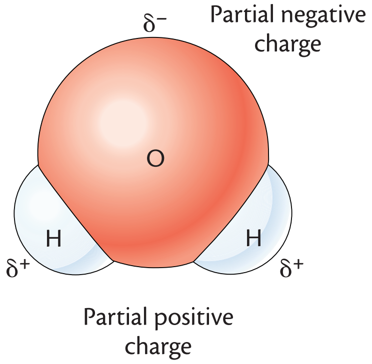


Figure 21 *Polarity of a water molecule. The δ+ and δ− indicate partial electrical charges.*

concentration of electrons (shown as δ−) compared with the two "hydrogen ends," which are electrically positive (shown as δ+). The Greek symbol δ (delta) means "partial"—thus partial plus and partial minus electrical charges are indicated. Because these charges balance, the molecule as a whole is electrically neutral.

Polar water molecules are attracted to other polar substances and to substances composed of electrically charged particles. These attractions make it possible for water to dissolve a great variety of substances.

One way to imagine the process of dissolving a substance in water is to liken it to a tug of war. Many solid substances, especially ionic compounds, are crystalline. In ionic crystals, positively charged cations are surrounded by negatively charged anions, with the anions likewise surrounded by cations. The crystal is held together by attractive forces between the cations and the anions. The substance will dissolve only if its ions are so strongly attracted to water molecules that the water "tugs" the ions from the crystal.

Water molecules are attracted to ions located on the surface of an ionic solid, as shown by the models in Figure 22a. The water molecule's negative (oxygen) end is attracted to the crystal's positive ions. The positive (hydrogen) ends of other water molecules are attracted to the negative ions of the crystal. When the attractive forces between the water molecules and the surface ions are strong enough, the bonds between the crystal and its surface ions become strained, and the ions may be pulled away from the crystal. Figure 22b uses models of water molecules and solute ions to illustrate the results of water "tugging" on solute ions. The detached ions become surrounded by water molecules, producing a water solution, as shown in Figure 22c.

The water solubility of some ionic compounds can be extremely low indeed. For example, at room temperature, lead(II) sulfide, PbS, has a solubility of only about 10^{-14} g (0.00000000000001 g) per liter of water solution.

Using the description and illustrations of this process, can you determine what influences whether an ionic solid will dissolve? Because dissolving entails competition among three types of attractions—those between solvent and solute particles, between solvent particles themselves, and between particles within the solute crystals—the properties of both solute and solvent affect whether two substances will form a solution. Water is highly polar, so it will be effective at dissolving charged or ionic substances. However, if positive–negative attractions between cations and anions in the crystal are sufficiently strong, a particular ionic compound may be only slightly soluble in water.

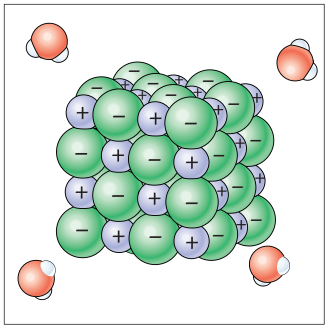


Figure 22a *Polar water molecules are attracted to the ions in an ionic crystal.*

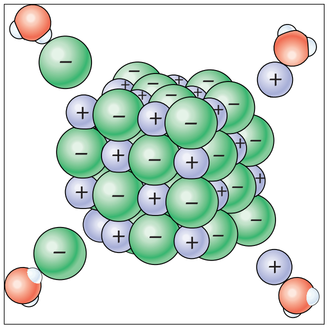


Figure 22b *Ions from the crystalline solid are pulled away by water molecules.*

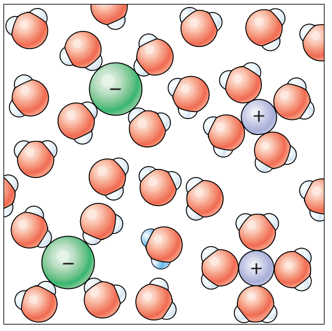


Figure 22c *A solution is formed when detached ions become surrounded by water molecules.*

MODELING MATTER

DISSOLVING IONIC COMPOUNDS

You have now learned about solubility, solubility curves, and the process of dissolving ionic compounds in water. As part of these discussions, visual models, such as those presented in Figure 22 (page 52) have been used to describe the process of dissolving. In this activity, you will combine these models with your knowledge of solubility curves to create new models of ions dissolved in water.

1. Suppose you dissolved 40 g potassium chloride (KCl) in 100 g water at 50 °C. You then let the solution cool to room temperature, about 25 °C.
   a. What changes would you see in the beaker as the solution cooled? See Figure 23.
   b. Draw models of what the contents in the beaker would look like at the molecular level at 50 °C, 40 °C, and 25 °C. Keep these points in mind:
   - You will need to consider whether the sample at each temperature is saturated, unsaturated, or supersaturated and draw the model accordingly. The solubility curve in Figure 23 will be helpful.
   - It is impossible to draw all the ions and molecules in this sample. The ions and molecules that you draw will represent what is happening on a much larger scale.
2. An unsaturated solution will become more concentrated if you add more solute. Decreasing the total volume of water in the solution (such as by evaporation) also causes an increase in the solution's concentration. Consider a solution made by dissolving 20 g KCl in 100 g water at 40 °C.
   a. Draw a model of this solution.
   b. Suppose that while the solution was kept at 40 °C, 25% of the water evaporated.
      i. Draw a model of this solution and describe how it differs from the model of the original solution.
      ii. How much water must evaporate at this temperature to cause the first potassium chloride crystals to form?
3. A solution may be diluted (made less concentrated) by adding water.
   a. Draw a model of a solution containing 10 g KCl in 100 g water at 25 °C.
   b. Suppose you diluted this solution by adding another 100 g water with stirring. Draw a model of this new solution.
   c. Compare your drawings in 3a and 3b. What key feature is different in the two models? Why?

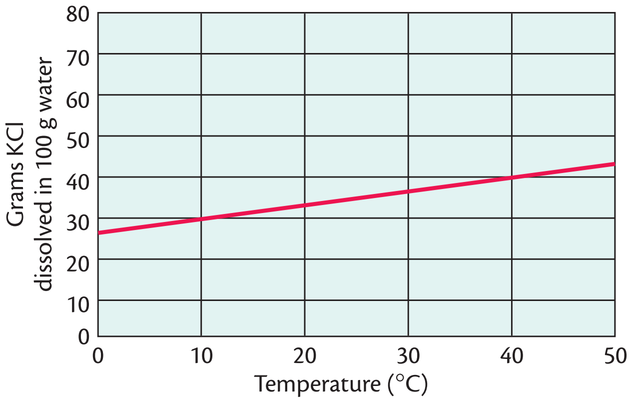


Figure 23 *Solubility curve for potassium chloride.*

Now that you know how solutions of ionic compounds are formed and described, it is time to consider some possible causes of the Riverwood fish kill. The information that you are about to read suggests some possible "culprits"—substances that can dissolve or be suspended in water and harm living things. These substances include heavy metals, acids and bases, molecular substances, and dissolved oxygen gas. Although all of these substances are normally found in natural water sources, the levels at which they are present can positively or negatively affect aquatic life.

C.5 INAPPROPRIATE HEAVY-METAL ION CONCENTRATIONS IN RIVER?

Many metal ions, such as iron(II) (Fe^{2+}), potassium (K^{+}), calcium (Ca^{2+}), and magnesium (Mg^{2+}), are essential to the health of humans and other organisms. For humans, these ions are obtained primarily from foods, but they may also be present in drinking water.

Not all metal ions that dissolve in water are beneficial, however. Some heavy-metal ions, called heavy metals because their atoms have greater mass than the masses of essential metallic elements, are harmful to humans and other organisms. Among the heavy-metal ions of greatest concern in water are cations of lead (Pb^{2+}) and mercury (Hg^{2+}). Lead and mercury are particularly likely to cause harm because they are widely used and dispersed in the environment. Heavy-metal ions are toxic because they bind to proteins in biological systems (such as your body), preventing the proteins from performing their intended tasks. As you might expect, because proteins play many important roles in body functioning, heavy-metal poisoning effects are severe. They include damage to the nervous system, brain, kidneys, and liver and even death.

The concentration of substances as they move through the food chain is known as bioaccumulation.

Unfortunately, heavy-metal ions are not removed as waste as they move up through the food chain. They become concentrated within the bodies of fish and shellfish, even when their abundance in the surrounding water is only a few parts per million. Such aquatic creatures then become hazardous for humans and other animals to consume.

In very low concentrations, heavy-metal ions are hard to detect in water and even more difficult and costly to remove. So how can heavy-metal poisoning be prevented? One of the easiest and most effective ways is to prevent the heavy-metal ions from entering water systems in the first place. This prevention can be accomplished by producing and using alternate materials that do not contain these ions and thus are not harmful to health or the environment. Such practices, which prevent pollution by eliminating the production and/or use of hazardous substances, are classified as examples of **green chemistry.** Such practices are applicable to heavy metals and to many other types of pollution.

Lead (Pb)

Locate lead on the Periodic Table at the back of your textbook. Compare its location with those of the essential metal ions that you just read about.

Lead is probably the heavy metal most familiar to you. Its symbol, Pb, is based on the element's original Latin name *plumbum,* also the source of the word "plumber."

Lead and lead compounds have been, and in some cases still are, used in pottery, automobile electrical storage batteries, solder, cooking vessels, pesticides, and paints. One compound of lead and oxygen, red lead (Pb_3O_4), is the primary ingredient in paint that protects bridges and other steel structures from corrosion.

Although lead water pipes were used in the United States in the early 1800s, they were replaced by iron pipes after it was discovered that water transported through lead pipes could cause lead poisoning. Romans constructed lead water pipes more than 2000 years ago; some of them are still in working condition. In modern homes, copper or plastic water pipes are used to prevent any contact between household water and lead.

Until the 1970s, the molecular compound tetraethyl lead, $Pb(C_2H_5)_4$, was added to gasoline to produce a better-burning automobile fuel. Unfortunately, the lead entered the atmosphere through automobile exhaust as lead oxide. Although the phaseout of leaded gasoline has reduced lead emissions, lead contamination remains in the soil surrounding heavily traveled roads. In some homes built before 1978 and not since repainted, the flaking of old leaded paint is another source of lead poisoning, particularly among children who may ingest the flaking paint.

Mercury (Hg)

Mercury is the only metallic element that is a liquid at room temperature. In fact, its symbol comes from the Latin *hydrargyrum,* meaning quick silver or liquid silver.

Locate mercury on the Periodic Table. Make a prediction about the locations of heavy metals on the Periodic Table.

Mercury has several important uses, some due specifically to its liquid state. It is an excellent electrical conductor, so it is used in "silent" light switches. It is also found in medical and weather thermometers, thermostats, mercury-vapor street lamps, fluorescent light bulbs, and some paints. Elemental mercury can be absorbed directly through the skin, and its vapor is quite hazardous to health. At room temperature, there will always be some mercury vapor present if liquid mercury is exposed to air, so any direct exposure to mercury is best avoided.

Because mercury compounds are toxic, they are useful in eliminating bacteria, fungi, and agricultural pests when used in antiseptics, fungicides, and pesticides. In the eighteenth and nineteenth centuries, mercury(II) nitrate, $Hg(NO_3)_2$, was used in making the felt hats popular at that time. After unintentionally absorbing this compound through their skin for several years, hat makers often suffered from mercury poisoning. Their symptoms included numbness, staggered walk, tunnel vision, and brain damage, thus giving rise to the expression "mad as a hatter."

The sudden release of a large amount of heavy-metal ions might cause a fish kill—depending on the particular metal ion, its concentration, the species of fish present, and other factors. Was such a release responsible for the Riverwood fish kill? As you read about other possible causes of the fish kill, keep in mind some questions that are relevant to all of the potential culprits. Is there a source of this substance along the Snake River near the site of the fish kill? What concentration of this solute would be toxic to various species of fish?

C.6 INAPPROPRIATE pH LEVELS IN RIVER?

You have likely heard the term pH used before, perhaps in connection with acid rain or hair shampoo. What is pH, and could it possibly help account for the fish kill in Riverwood? The **pH scale** is a convenient way to measure and report the acidic, basic, or chemically neutral character of a solution.

Nearly all pH values are in the range from 0 to 14, although some extremely acidic or basic solutions may be outside this range. At room temperature, any pH values less than 7 indicate an acidic condition; the lower the pH, the more acidic the solution. Solutions with pH values greater than 7 are basic; the higher the pH, the more basic the solution. Basic solutions are also called alkaline solutions. Quantitatively, a change of one pH unit indicates a tenfold difference in acidity or alkalinity. For example, lemon juice, with a

Some Common Acids and Bases

Name	Formula	Use
Acids		
Acetic acid	$HC_2H_3O_2$	In vinegar (typically a 5% solution of acetic acid)
Carbonic acid	H_2CO_3	In carbonated soft drinks
Hydrochloric acid	HCl	Used in removing scale buildup from boilers and for cleaning materials
Nitric acid	HNO_3	Used in the manufacture of fertilizers, dyes, and explosives
Phosphoric acid	H_3PO_4	Added to some soft drinks to give a tart flavor; also used in the manufacture of fertilizers and detergents
Sulfuric acid	H_2SO_4	Largest-volume substance produced by chemical industry; present in automobile battery fluid
Bases		
Calcium hydroxide	$Ca(OH)_2$	Present in mortar, plaster, and cement; used in paper pulping and dehairing animal hides
Magnesium hydroxide	$Mg(OH)_2$	Active ingredient in milk of magnesia
Potassium hydroxide	KOH	Used in the manufacture of some liquid soaps
Sodium hydroxide	$NaOH$	A major industrial product; active ingredient in some drain and oven cleaners; used to convert animal fats into soap

Figure 24 *The name, formula, and common use of some familiar acids and bases.*

pH of about 2, is nearly ten times as acidic as soft drinks, which have a pH of about 3.

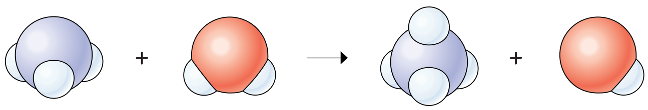
Acids and bases, some examples of which are listed in Figure 24, can also be identified by certain chemical properties. For example, the vegetable dye litmus turns blue in a basic solution and red in an acidic solution. Both acidic and basic solutions conduct electricity. Each type of solution has a distinctive taste and a distinctive feel on your skin. (**CAUTION:** *You should never test these sensory properties in the laboratory.*) In addition, concentrated acids and bases are able to react chemically with many other substances. You are probably familiar with the ability of acids and bases to corrode, or wear away, other materials. Corrosion is a type of chemical reaction.

Vinegar is an acid that you have tasted; that common kitchen ingredient is considered a dilute solution of acetic acid.

Most acid molecules have one or more hydrogen atoms that can be released rather easily in water solution. These "acidic" hydrogen atoms are usually written first in the formula for an acid. See Figure 24.

Many bases are ionic substances that include hydroxide ions (OH^-) in their structures. Sodium hydroxide, $NaOH$, and barium hydroxide, $Ba(OH)_2$, are two examples. Some bases, such as ammonia (NH_3) and baking soda (sodium bicarbonate, $NaHCO_3$), contain no OH^- ions but still produce basic solutions because they react with water to generate OH^- ions, as illustrated by the following equation.

$$NH_3 \quad + \quad H_2O \quad \rightarrow \quad NH_4^+ \quad + \quad OH^-$$



What about substances that display neither acidic nor basic characteristics? Chemists classify these substances as chemically neutral. Water, sodium chloride ($NaCl$), and table sugar (sucrose, $C_{12}H_{22}O_{11}$) are all examples of chemically neutral compounds.

At 25 °C, a pH of 7 indicates a chemically neutral solution. The pH values of some common materials are shown in Figure 25.

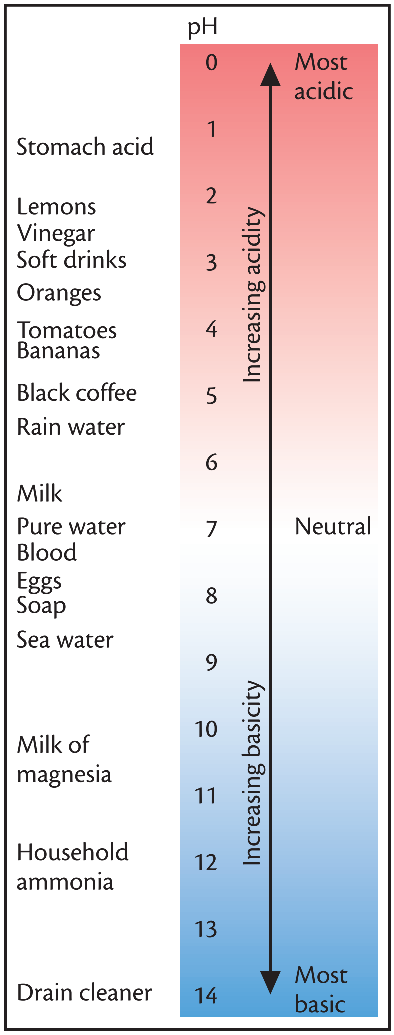


Figure 25 *The pH values of some common materials.*

As you can see in Figure 25, rainwater is naturally slightly acidic. This is because the atmosphere contains certain substances—carbon dioxide (CO_2) for one—that produce acidic solutions when dissolved in water. Both acidic and basic solutions have effects on living organisms—effects that depend on the pH of the water. When the pH of water is too low (meaning high acidity), fish-egg development is impaired, thus hampering the ability of fish to reproduce. Water solutions with low (acidic) pH values also tend to increase the concentrations of metal ions in natural waters by leaching the metals from surrounding soil. These metal ions can include aluminum ions (Al^{3+}), which are toxic to fish when present in sufficiently high concentration. High pH (basic contamination) is a problem for living organisms primarily because alkaline solutions are able to dissolve organic materials, including skin and scales.

The U.S. Environmental Protection Agency (EPA) requires drinking water to be within the pH range from 6.5 to 8.5. However, most fish can tolerate a slightly wider pH range, from about 5 to 9, in lake or river water. Serious freshwater anglers try to catch fish in water between pH 6.5 and 8.2.

On a normal day, the pH of the water in the Snake River in Riverwood ranges between 7 and 8, nearly optimal for freshwater fishing. Could the pH have changed abruptly, killing the fish? Was acidic or basic contamination responsible for the Riverwood crisis?

C.7 INAPPROPRIATE MOLECULAR-SUBSTANCE CONCENTRATIONS IN RIVER?

Solubility of Molecular Substances

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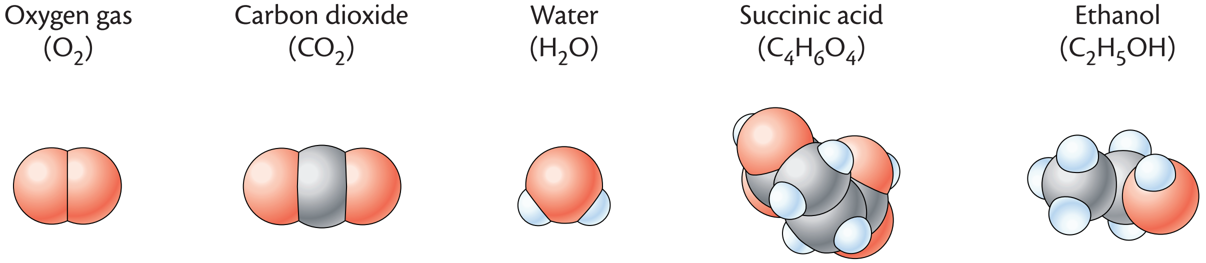
Until now, the types of substances considered suspects in the Riverwood mystery have been ionic substances, those that dissolve in water to release ions. Are there other types of substances that dissolve in water and possibly present a hazard to aquatic life? Some substances, such as sugar and ethanol, dissolve in water but not in the form of ions. These substances belong to a class of materials known as **molecular substances** because they are composed of molecules.

Unlike ionic substances, which are crystalline solids at normal conditions, molecular substances can be found as solids, liquids, or gases at room temperature. Some molecular substances such as oxygen (O_2) and carbon dioxide (CO_2) have little attraction between their molecules and are thus gases at normal conditions. Molecular substances such as ethanol (ethyl alcohol, C_2H_5OH) and water (H_2O) have larger between-molecule attractions, causing these "stickier" molecules to form liquids at normal conditions. Other molecular substances with even greater between-molecule attractions—succinic acid ($C_4H_6O_4$), for example—are solids at normal conditions. Their stronger attractive forces hold the molecules together more tightly, in effect determining in which state the substance will be found.

You investigated the solubility behavior of succinic acid earlier. See p. 49.



What determines how soluble a molecular substance will be in water? The attraction of a substance's molecules for each other compared with their attraction for water molecules plays a major part. But what causes these attractions? The distribution of electrical charge within molecules has a great deal to do with it.

Most molecular compounds contain atoms of nonmetallic elements. As you learned earlier (page 32), these atoms are linked together by the attraction of one atom's positively charged nucleus for another atom's negatively charged electrons. If differences in electron attraction between atoms are large enough, electrons move from one atom to another, forming ions. This is what often happens between a metallic atom and a nonmetallic atom when an ionic compound is formed. An atom's ability to attract shared electrons in its bonding within a substance is known as its **electronegativity.** In molecular substances, these differences in electron attraction, or electronegativities, are not large enough to cause ions to form, but they may cause the electrons to be unevenly distributed among the atoms.

You already know that the "oxygen end" of a water molecule is electrically negative compared with its positive "hydrogen end." That is, water molecules are polar and serve as the most common example of a polar solvent. Such charge separation (and resulting molecular polarity) is found in many molecules whose atoms have sufficiently different electronegativities.

Polar molecules tend to dissolve readily in polar solvents such as water. For example, water is a good solvent for sugar and ethanol, both composed of polar molecules. Similarly, nonpolar liquids are good solvents for other nonpolar molecules. Nonpolar cleaning fluids are used to "dry clean" clothes because they readily dissolve nonpolar body oils found in fabric. In contrast, nonpolar molecules (such as those of oil and gasoline) do not dissolve well in polar solvents (such as water or ethanol).

This pattern of solubility behavior—polar substances dissolving in polar solvents, nonpolar substances dissolving in nonpolar solvents—is often summarized in the generalization "Like dissolves like." This rule also explains why nonpolar liquids are usually ineffective in dissolving ionic and polar molecular substances.

Were dissolved molecular substances present in the Snake River water where the fish died? Most likely yes; at least in small amounts. Were they responsible for the fish kill? That depends on which molecular substances were present and at what concentrations. And that, in turn, depends on how each solute interacts with water's polar molecules.

In the following laboratory activity, you will investigate and compare the solubility behavior of some typical molecular and ionic substances.

Unfortunately, many nonpolar dry-cleaning solvents are damaging to both human health and the environment. However, the recent development of new technologies allows environmentally benign nonpolar solvents such as liquefied carbon dioxide to be used in the dry-cleaning industry.

Various molecular substances may normally be present at very low levels—so low that no harm to living things is observed.

C.8 SOLVENTS

Laboratory Activity

Introduction

The *Riverwood News* reported earlier that Dr. Brooke believes that a substance dissolved in the Snake River is one likely fish-kill cause. She based her judgment on her chemical knowledge and experiences with water and other substances. Dr. Brooke also has a general idea about which contaminating solutes she can initially rule out: those that cannot dissolve appreciably in water. Such background knowledge helps Dr. Brooke (and other chemists) reduce the number of water tests required in the laboratory.

What, then, do the terms "soluble" and "insoluble" mean? Is anything truly insoluble in water? It is likely that at least a few molecules or ions of any substance will dissolve in water. Thus the term "insoluble" actually refers to substances that are only very, very slightly soluble in water. Chalk, 1.53 mg is equal to only 0.00153 g.

for example, is considered insoluble in water, even though 1.53 mg calcium carbonate ($CaCO_3$, the main ingredient in chalk) can dissolve in 100 g water at 25 °C.

In this laboratory activity, you will first investigate the solubilities of various molecular and ionic solutes in water. These solubility data, along with toxicity data, will help you rule out some solutes as likely causes of the fish kill. You will then test other solvents and examine the solubility data for any general patterns.

Your teacher will tell you which particular solutes you will investigate. List them in your laboratory notebook.

Part I: Designing a Procedure for Investigating Solubility in Water

Your teacher will direct you to discuss with either the whole class or your laboratory partner a procedure for testing the room-temperature solubilities of the substances listed in your laboratory notebook. (If you have performed solubility tests before, it may be useful to recall how you did them.) With your partner, design a step-by-step investigation that will allow you to determine whether each solute is soluble (S), slightly soluble (SS), or insoluble (I) in room-temperature water.

In this activity "insoluble" (I) means that the solubility is so low that no solute was observed to dissolve.

The following questions will help you design your procedure.

1. What particular observations will allow you to judge how well each solute dissolves in the polar solvent water? That is, how will you decide whether to classify a given solute as soluble, slightly soluble, or insoluble in water?
2. Which variables will need to be controlled? Why?
3. How should the solute and solvent be mixed—all at once or a little at a time? Why?

In designing your procedure, keep these concerns in mind.

- Avoid any direct contact of your skin with any solutes.
- Follow your teacher's directions for waste disposal.

When you and your partner have agreed on a written procedure, get it approved by your teacher. Construct a data table for your results and you are ready for Part II.

Part II: Investigating Solubility in Water

Use your approved procedure to investigate the solubility in water of the listed substances. Record the data in your data table.

Part III: Investigating Solubility in Ethanol and Lamp Oil

It is clear that the task of determining what may have caused the fish kill can be simplified somewhat by focusing efforts on substances that will dissolve appreciably in water. However, in dealing with other solubility-based problems, chemists sometimes find it helpful to use solvents other than water—ethanol and lamp oil serve that role in this activity.

You and your partner will investigate the solubility of some or all of the solutes from Part II in ethanol and lamp oil. You should also test the solubility of water in ethanol and in lamp oil. By gaining experiences with three liquid solvents—lamp oil, ethanol, and water—you will be prepared to recognize some general patterns regarding solubility behavior.

Can you use the same procedure that you designed for Part II? If not, what parts of that procedure should be revised? In considering your Part III procedure, again keep the questions listed on page 60 in mind.

Have your proposed procedure approved by your teacher. Before you start the laboratory work, test your interpretation of the results of Part II by predicting what you think you will observe regarding solubility in each case. Include these predictions in your data table for Part III. Then collect and record your data for both solvents.

Wash your hands thoroughly before leaving the laboratory.

Questions

Part II

1. According to your data, which tested solutes are least likely to be dissolved in the Snake River? Why?
2. Compare your data with those of the rest of the class. Are there any differences? If so, how can those differences be explained?

Part III

3. a. How does the behavior of ethanol as a solvent compare with that of water?
   b. How does ethanol's behavior compare with that of lamp oil?
4. a. Were any of your solubility observations unexpected?
   b. If so, explain what you expected, why it was expected, and how your expectations compare with what you actually observed.
5. Based on your data, what general pattern of solubility behavior can you summarize and describe?
6. Predict the solubility behavior of each solid solute in:
   a. hexane, a liquid that is essentially insoluble in water.
   b. ethylene glycol, a liquid that is very soluble in ethanol.
7. a. Given that water is a polar solvent and lamp oil is a nonpolar solvent, classify each molecular solute tested as polar or nonpolar.
   b. How did you decide?
8. How useful is the rule "like dissolves like" for predicting solubility? Explain your answer on the basis of your results.
9. In Part II, water was the "solvent," but in Part III, water was a "solute."
   a. How can it be both?
   b. How do you know whether a substance is a solute or a solvent?

The familiar saying "oil and water do not mix" has a chemical basis!

You now know about the solubility of some solids and liquids in water. As you read on to learn about the behavior of gases in natural waters, consider the possibility that a dissolved gas was responsible for the Snake River catastrophe.

C.9 INAPPROPRIATE DISSOLVED OXYGEN LEVELS IN RIVER?

You have noted that the solubility of ionic and molecular solids in water usually increases when the water temperature is raised. Do gases behave similarly to solids in solution? Look at Figure 26, which shows the solubility curve for oxygen gas, plotted as milligrams oxygen dissolved per 1000 g water.

What is the solubility of oxygen gas in 20 °C water? In 40 °C water? As you can see, increasing the water temperature causes the gas to be less soluble! Note also the magnitude of the values for oxygen solubility. Compare these values with those for solid solutes as shown in Figure 20. At 20 °C, about 30 g potassium nitrate will dissolve in 100 g water. In contrast, only about 9 mg (0.009 g) oxygen gas will dissolve in ten times as much water—1000 g water—at this temperature. It should be clear that most gases are far less soluble in water than are many ionic solids.

When you considered the solubility of molecular and ionic solids in water, you found that solubility depended on two factors—temperature and the nature of the solvent. The solubility of a gas depends on these two factors as well. But it also depends on gas pressure. Referring to Figure 27, note what happens to oxygen's solubility as the pressure of oxygen gas above it is increased. Does more or less gas dissolve in the same amount of water? What happens if the gas pressure is doubled? For example, look at the solubility of O_2 at one atmosphere and at two atmospheres of pressure. Also consider the shape of the graph line. What type of relationship does a linear graph line indicate?

The metric unit for pressure is the pascal (Pa): 1 atmosphere (atm) = 101 325 Pa

As you have by now deduced, gas solubility in water is directly proportional to the pressure of that gaseous substance on the liquid. You see one effect of this proportionality every time you open a can or bottle of carbonated soft drink. As the gas pressure on the liquid is reduced by opening the container to the air, some dissolved carbon dioxide gas (CO_2) escapes from the liquid in a rush of bubbles.

Because there is not very much CO_2(g) present in air, carbon dioxide gas must be forced into the carbonated beverage at high pressure just before

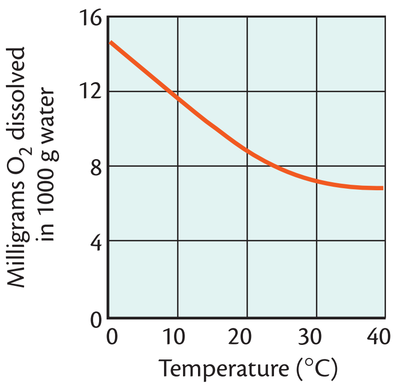


Figure 26 *Solubility curve for O_2 gas in water in contact with air.*

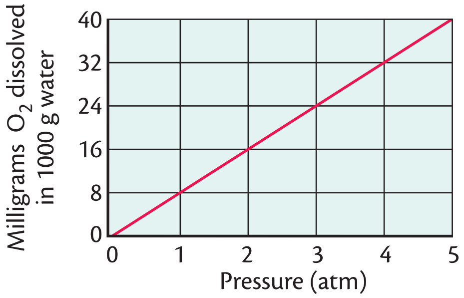


Figure 27 *Relationship between solubility and pressure of O_2 at 25 °C.*

the container is sealed. This increases the amount of carbon dioxide that can dissolve in the beverage. When the can or bottle is opened, CO_2 gas pressure on the liquid drops back to its usual low level. Dissolved carbon dioxide gas escapes from the liquid until it reaches its (lower) solubility at this lower pressure. When the fizzing stops, you might describe the beverage as having "gone flat." Actually, the excess carbon dioxide gas has simply escaped into the air; the resulting solution is still saturated with CO_2 at the new conditions.

C.10 TEMPERATURE, DISSOLVED OXYGEN, AND LIFE

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Temperature, Dissolved Oxygen & Life

On the basis of what you know about the effect of temperature on solubility, you may be wondering if the temperature of the Snake River had something to do with the fish kill. As you just learned, water temperature affects how much oxygen gas can dissolve in the water. Various forms of aquatic life, including the many species of fish, have different requirements for the concentration of dissolved oxygen needed to survive. Figure 28 contains this information. How, then, does a change in the temperature of the natural waters affect the fish internally?

Fish are "cold-blooded" animals; their body temperatures rise and fall with the surrounding water temperature. If the water temperature rises, the body temperatures of fish also rise. This increase in body temperature in turn increases fish metabolism, a complex series of interrelated chemical reactions that keep fish alive. As these internal reaction rates speed up, the fish eat more, swim more, and require more dissolved oxygen. The rate of metabolism also increases for other aquatic organisms, such as aerobic bacteria, that compete with fish for dissolved oxygen.

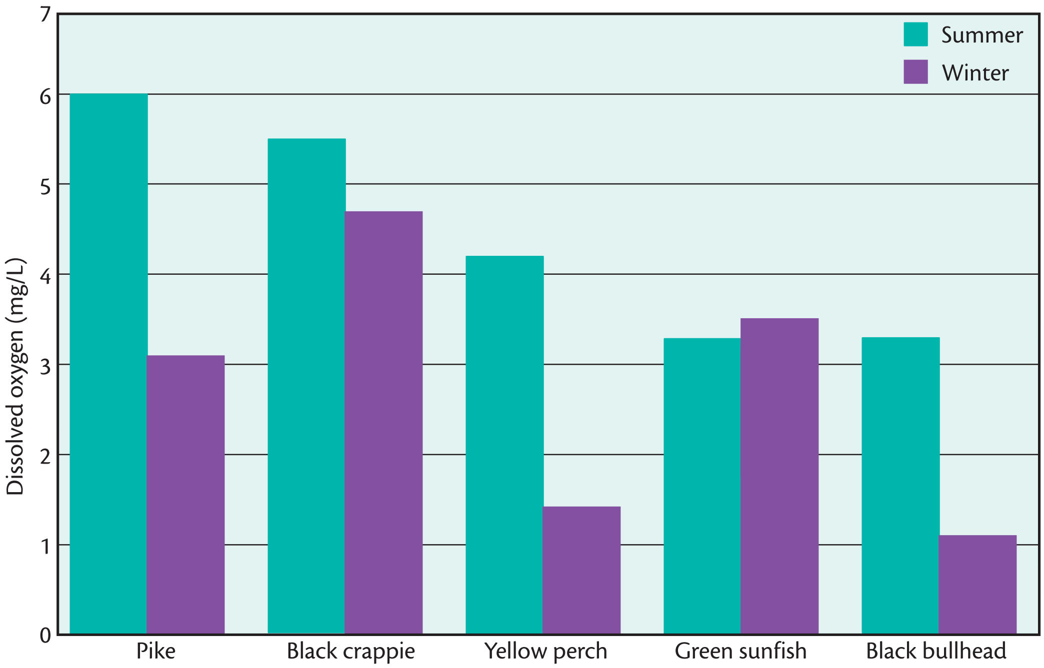

Figure 28 *Dissolved-oxygen requirements of various fish.*

Figure 29 *Maximum water temperature tolerance in fish.*

Maximum Water Temperature Tolerance in Fish (24-hour Exposure)

	Maximum Temperature	
Fish	**°C**	**°F**
Trout (brook, brown, rainbow)	24	75
Channel catfish	35	95
Lake herring (cisco)	25	77
Largemouth bass	34	93
Northern pike	30	86

As you can now see, an increase in water temperature affects fish both externally and internally. A long stretch of hot summer days sometimes results in large fish kills, in which hundreds of fish literally suffocate. Figure 29 summarizes the maximum water temperatures at which selected fish species can survive.

Sometimes hot summer days are not to blame. Often, high lake or river water temperatures can be traced to human activity. Many industries, such as electrical power generation, depend on natural bodies of water to cool heat-producing processes. Cool water is drawn from lakes or rivers into an industrial or power-generating plant, and devices called heat exchangers transfer thermal energy (heat energy) from the processing area to the cooling water. The heated water is then released back into the lakes or rivers, either immediately or after the water has partly cooled. If the water is too warm, it can upset the balance of life in lakes and rivers.

At this point, it should be clear that there is a lower limit on the amount of dissolved oxygen needed for fish to survive. Is there an upper limit? Can too much dissolved oxygen be a problem for fish? In nature, both oxygen and nitrogen gas are present in the air at all times. See Figure 30. When oxygen gas dissolves, so does nitrogen gas. This fact turns out to be significant when considering the upper limit for dissolved gases. When the total amount of dissolved gases—primarily oxygen and nitrogen—reaches

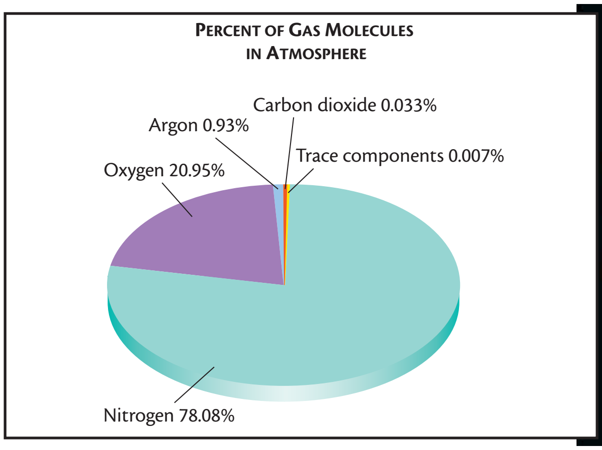

Figure 30 *The percent composition of the atmosphere—the envelope of gases that surrounds planet Earth.*

between 110% and 124% of saturation (a state of supersaturation), a condition called gas-bubble trauma may develop in fish.

This situation is dangerous because the supersaturated solution causes gas bubbles to form in the blood and tissues of fish. Oxygen-gas bubbles can be partially utilized by fish during metabolism, but nitrogen-gas bubbles block the flow of blood within the fish. This blockage results in the death of the fish within hours or days. Gas-bubble trauma can be diagnosed by noting gas bubbles in the gills of dead fish if they are dissected promptly after death. Supersaturation of water with nitrogen and oxygen gases can occur at the base of a dam or a hydroelectric project as released water forms "froth," trapping large quantities of air.

So, back to the original question: What caused the Riverwood fish kill? You have considered several possible causes. How will you determine which one was the actual cause? You will start by examining water-related data collected by scientists and engineers on the Snake River. From these data and what you have learned, you will make your own decision.

C.11 DETERMINING THE CAUSE OF THE FISH KILL

Making Decisions

Snake River watershed data have been collected and monitored since the early 1900s. Although some measurements and methods have changed during that time, excellent data have been gathered, particularly in recent years. Data are available regarding the following factors.

- water temperature and dissolved oxygen
- rainfall
- water flow
- dissolved molecular substances
- heavy metals
- pH
- nitrate and phosphate levels
- organic carbons

Your teacher will assign you to a group to study some of the data just listed. Each group will complete data-analysis procedures for its assigned data. After groups have finished their work, the class will share their analyses and draw conclusions about factors possibly related to the fish kill.

The following background information will help you complete the analysis of the data assigned to you by your teacher.

DATA ANALYSIS

Introduction

Interpreting graphs of environmental data requires a slightly different approach from the one you took in interpreting a solubility curve (page 47). Rather than seeking a predictable relationship, you will be looking for regularities or patterns among the values. Any major irregularity in the data may suggest a problem related to the water factor being evaluated. The following suggestions will help you prepare and interpret such graphs.

- Choose your scale so that the graph becomes large enough to fill most of the available space on the graph paper.
- Assign each regularly spaced division on the graph paper some convenient, constant value. The graph-paper line interval should have a value easily "divided by eye," such as 1, 2, 5, or 10, rather than a value such as 6, 7, 9, or 14.
- An axis scale does not have to start at "zero," particularly if the plotted values cluster in a narrow range not near zero.
- Label each axis with the quantity and unit being graphed.
- Plot each point. Then draw a small symbol around each point, like this: ⊙. If you plot more than one set of data on the same graph, distinguish each set of points by using a different color or geometric shape, such as: ⊡, ▽, or △.
- Give your graph a title that will readily convey its meaning and purpose to a reader.
- If you use technology—such as a graphing calculator or computer program—to prepare your graphs, ensure that you follow the guidelines just given. Different devices or programs have different ways of processing data. Choose the appropriate type of graph (scatter plot, bar graph, and so forth) for your data.

Procedure

1. Prepare a graph for each of your group's Snake River data sets. Label the *x,* or independent, axis with the consecutive months indicated in the data. Label the *y,* or dependent, axis with the water factor measured, accompanied by its units.
2. Plot each data point and connect the consecutive points with straight lines.

Questions

1. Is any pattern apparent in your group's plotted data?
2. Can you offer possible explanations for any pattern or irregularities that you detect?
3. Do you think the data analyzed by your group might help to account for the Snake River fish kill? Why? How?
4. Prepare to share your group's findings regarding Questions 1 through 3 in a class discussion. During the class discussion, take notes on key findings reported by each data-analysis group. Also note and record significant points raised in the data-analysis discussions.

Questions & Answers CD-ROM WWW.

Your class will reassemble several times during your study of Section D to discuss and consider implications of the water-analysis data that you have just processed. In particular, guided by the patterns and irregularities found in your analysis of Snake River data, your class will seek an explanation or scenario that accounts for the observed data and for the resulting Riverwood fish kill. Good luck!

SECTION SUMMARY

Reviewing the Concepts

◆ **The solubility of a substance in water can be expressed as the quantity of that substance that will dissolve in a certain quantity of water at a specified temperature.**

1. If the solubility of sugar (sucrose) in water is 2.0 g/mL at room temperature, what is the maximum amount of sugar that will dissolve in 946 mL (1 quart) water?
2. Explain why three teaspoons of sugar will completely dissolve in a serving of hot tea, but not in an equally sized serving of iced tea.
3. Rank the substances in Figure 20 (page 46) from most soluble to least soluble

   a. at 20 °C.
   b. at 80 °C.

◆ **Solutions can be described qualitatively or quantitatively. In qualitative terms, solutions can be classified as unsaturated, saturated, or supersaturated. In quantitative terms, the concentration of a solution expresses the relative quantities of solute and solvent in a particular solution.**

4. A 35-g sample of ethanol is dissolved in 115 g water. What is the concentration of the ethanol, expressed as grams ethanol per 100 g solution?
5. Calculate the masses of water and sugar in a 55-g sugar solution that is labeled 20.0% sugar.
6. Using the graph on page 46, answer the following questions about the solubility of potassium nitrate, KNO_3.

   a. What is the maximum mass of KNO_3 that can dissolve in 100 g water if the water temperature is 20 °C?
   b. At 30 °C, 55 g KNO_3 is dissolved in 100 g water. Is this solution saturated, unsaturated, or supersaturated?
   c. A saturated solution of KNO_3 is formed in 100 g water at 75 °C. If the saturated solution is cooled to 40 °C, how many grams of solid KNO_3 should form?

◆ **Polar bonds have an uneven distribution of electrical charge. Polar O–H bonds in water help explain water's ability to dissolve many ionic solids.**

7. Draw a model that shows how molecules in liquid water generally arrange themselves relative to one another.
8. Why does table salt (NaCl) dissolve in water but not in cooking oil?

◆ **Heavy-metal ions, such as Pb^{2+} and Hg^{2+}, are useful resources but can be toxic if introduced into biological systems, even in small amounts.**

9. a. What are heavy metals?
   b. List some general effects of heavy-metal poisoning.
10. What are some possible sources of human exposure to heavy metals?
11. What items might be found in an urban landfill that could contribute heavy-metal ions to groundwater?

◆ **Water solutions can be characterized as acidic, basic, or chemically neutral on the basis of their chemical properties.**

12. Classify each sample as acidic, basic, or chemically neutral:
    a. seawater (pH = 8.6)
    b. drain cleaner (pH = 13.0)
    c. vinegar (pH = 2.7)
    d. pure water (pH = 7.0)

13. Which is more acidic, a tomato or a soft drink? (See Figure 25 on page 57)

14. How many times more acidic is a solution of pH 2 than a solution of pH 4?

◆ **The solubility of a molecular substance in water depends on the relative strength of attractive forces between solute and water molecules, compared with the strength of attractive forces between solute molecules and between water molecules.**

15. Would you select ethanol, water, or lamp oil to dissolve a nonpolar molecular substance? Explain.

16. Explain the phrase "Like dissolves like."

17. Explain why greasy dishes cannot be satisfactorily cleaned with pure water.

◆ **The solubility of a gaseous substance in water depends on the temperature of the water and the external pressure of the gas.**

18. As scuba divers descend, the pressure increases on the gases that they are breathing. How does the increasing pressure affect the amount of gas dissolved in their blood?

19. Given your knowledge of gas solubility, explain why a bottle of warm cola produces more "fizz" when opened than does a bottle of cold cola.

Connecting the Concepts

20. At room temperature, C_2H_6 is a gas but C_2H_5OH is a liquid. Suggest an explanation for this difference.

21. Predict the relative solubilities of C_2H_6 and C_2H_5OH in water.

22. From each of the following pairs, select the water source more likely to contain the higher concentration of dissolved oxygen. Give a reason for each choice.
    a. a river with rapids or a calm lake
    b. a lake in spring or the same lake in summer
    c. a lake containing only sunfish or a lake containing only pike

Extending the Concepts

23. Read the label on a container of baking soda. Compare it with the label on a can of baking powder. Which one contains an acid ingredient? Suggest a reason for including the acid in the mixture.

24. Describe how changes in solubility due to temperature could be used to separate two solid, water-soluble substances.

25. The continued health of an aquarium depends on the balance of the solubilities of several substances. Investigate how this balance is maintained in a freshwater aquarium.

26. The pH of rainwater is approximately 5.5. Rainwater flows into the ocean. The pH of ocean water, however, is approximately 8.7. Investigate reasons for the difference in pH.