Factors That Affect Rate of Dissolving and Solubility

As you learned in section 8.1, the solubility of a solute is the *amount* of solute that dissolves in a given volume of solvent at a certain temperature. Solubility is determined by the intermolecular attractions between solvent and solute particles. You will learn more about solubility and the factors that affect it later in this section. First, however, you will look at an important property of a solution: the **rate of dissolving**, or how quickly a solute dissolves in a solvent.

The rate of dissolving depends on several factors, including temperature, agitation, and particle size. You have probably used these factors yourself when making solutions like the fruit juice shown in Figure 8.5.

Figure 8.5 Fruit juice is soluble in water. The concentrated juice in this photograph, however, will take a long time to dissolve. Why?

Factors That Affect the Rate of Dissolving

You may have observed that a solute, such as sugar, dissolves faster in hot water than in cold water. In fact, *for most solid solutes, the rate of dissolving is greater at higher temperatures.* At higher temperatures, the solvent molecules have greater kinetic energy. Thus, they collide with the undissolved solid molecules more frequently. This increases their rate of dissolving.

Suppose that you are dissolving a spoonful of sugar in a cup of hot coffee. How can you make the sugar dissolve even faster? You can stir the coffee. *Agitating a mixture by stirring or by shaking the container increases the rate of dissolving.* Agitation brings fresh solvent into contact with undissolved solid.

Finally, you may have noticed that a large lump of solid sugar dissolves more slowly than an equal mass of powdered sugar. *Decreasing the size of the particles increases the rate of dissolving.* When you break up a large mass into many smaller masses, you increase the surface area that is in contact with the solvent. This allows the solid to dissolve faster. Figure 8.6 shows one way to increase the rate of dissolving.

Figure 8.6 Chemists often grind solids into powders using a mortar and pestle. This increases the rate of dissolving.

Solubility and Particle Attractions

By now, you are probably very familiar with the process of dissolving. You already know what it looks like when a solid dissolves in a liquid. Why, however, does something dissolve? What is happening at the molecular level?

The reasons why a solute may or may not dissolve in a solvent are related to the forces of attraction between the solute and solvent particles. These forces include the attractions between two solute particles, the attractions between two solvent particles, and the attractions between a solute particle and a solvent particle. When the forces of attraction between *different* particles in a mixture are stronger than the forces of attraction between *like* particles in the mixture, a solution forms. The strength of each attraction influences the solubility, or the amount of a solute that dissolves in a solvent.

To make this easier to understand, consider the following three steps in the process of dissolving a solid in a liquid.

The Process of Dissolving at the Molecular Level

Step 1 The forces between the particles in the solid must be broken. This step always requires energy. In an ionic solid, the forces that are holding the ions together must be broken. In a molecular solid, the forces between the molecules must be broken.

Step 2 Some of the intermolecular forces between the particles in the liquid must be broken. This step also requires energy.

Step 3 There is an attraction between the particles of the solid and the particles of the liquid. This step always gives off energy.

The solid is more likely to dissolve in the liquid if the energy change in step 3 is greater than the sum of the energy changes in steps 1 and 2. (You will learn more about energy and dissolving in Unit 5.)

Polar and Non-Polar Substances

In the Thought Lab in section 8.1, you observed that solid iodine is insoluble in water. Only a weak attraction exists between the non-polar iodine molecules and the polar water molecules. On the other hand, the intermolecular forces between the water molecules are very strong. As a result, the water molecules remain attracted to each other rather than attracting the iodine molecules.

You also observed that iodine is soluble in kerosene. Both iodine and kerosene are non-polar substances. The attraction that iodine and kerosene molecules have for each other is greater than the attraction between the iodine molecules in the solid and the attraction between the kerosene molecules in the liquid.

The Concept Organizer shown on the next page summarizes the behaviour of polar and non-polar substances in solutions. You will learn more about polar and non-polar substances later in this section.

Polar compounds e.g. sucrose, $C_{12}H_{22}O_{11}$	dissolve in	Polar solvents e.g. water, $H_2O_{(\ell)}$

do not dissolve in

Non-Polar compounds e.g. iodine, $I_{2(s)}$	dissolve in	Non-Polar solvents e.g. benzene, $C_6H_{6(\ell)}$
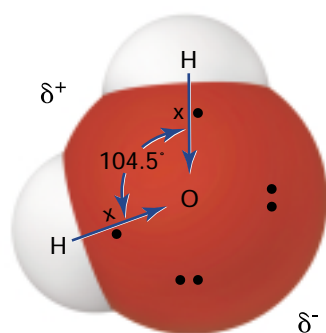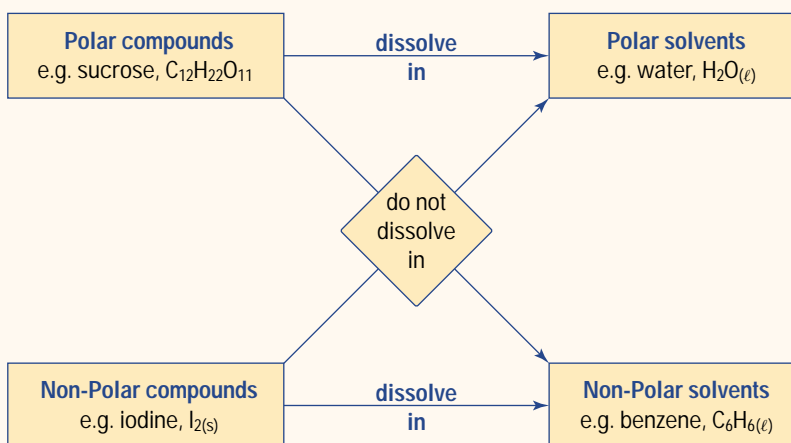
Figure 8.7 The bent shape and polar bonds of a water molecule give it a permanent dipole.

Solubility and Intermolecular Forces

You have learned that solubility depends on the forces between particles. Thus, polar substances dissolve in polar solvents, and non-polar substances dissolve in non-polar solvents. What are these forces that act between particles?

In Chapter 3, section 3.3, you learned that a water molecule is polar. It has a relatively large negative charge on the oxygen atom, and positive charges on both hydrogen atoms. Molecules such as water, which have charges separated into positive and negative regions, are said to have a permanent dipole. A **dipole** consists of two opposite charges that are separated by a short distance. Figure 8.7 shows the dipole of a water molecule.

Dipole-Dipole Attractions

The attraction between the opposite charges on two different polar molecules is called a **dipole-dipole attraction**. Dipole-dipole attractions are intermolecular. This means that they act *between* molecules. Usually they are only about 1% as strong as an ionic or covalent bond. In water, there is a special dipole-dipole attraction called **hydrogen bonding**. It occurs between the oxygen atom on one molecule and the hydrogen atoms on a nearby molecule. Hydrogen bonding is much stronger than an ordinary dipole-dipole attraction. It is much weaker, however, than the covalent bond between the oxygen and hydrogen atoms in a water molecule. Figure 8.8 illustrates hydrogen bonding between water molecules.

Electronic Learning Partner

Go to the Chemistry 11 Electronic Learning Partner to find out how hydrogen bonding leads to water's amazing surface tension.

Figure 8.8 Hydrogen bonding between water molecules is shown as dotted lines. The H atoms on each molecule are attracted to O atoms on other water molecules.

Ion-Dipole Attractions

Ionic crystals consist of repeating patterns of oppositely charged ions, as shown in Figure 8.9. What happens when an ionic compound comes in contact with water? The negative end of the dipole on some water molecules attracts the cations on the surface of the ionic crystal. At the same time, the positive end of the water dipole attracts the anions. These attractions are known as **ion-dipole attractions**: attractive forces between an ion and a polar molecule. If ion-dipole attractions can replace the ionic bonds between the cations and anions in an ionic compound, the compound will dissolve. *Generally an ionic compound will dissolve in a polar solvent.* For example, table salt (sodium chloride, NaCl) is an ionic compound. It dissolves well in water, which is a polar solvent.

When ions are present in an aqueous solution, each ion is **hydrated**. This means that it is surrounded by water molecules. Hydrated ions can move through a solution and conduct electricity. A solute that forms an aqueous solution with the ability to conduct electricity is called an **electrolyte**. Figure 8.10 shows hydrated sodium chloride ions, which are electrolytes.
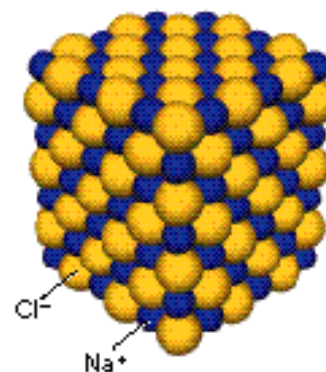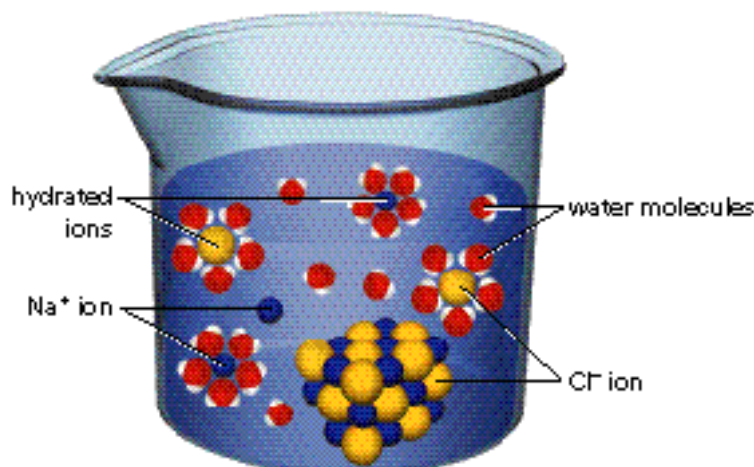

Figure 8.9 Ionic crystals have very ordered structures.

Figure 8.10 Ion-dipole attractions help to explain why sodium chloride dissolves in water.

An Exception: Insoluble Ionic Compounds

Although most ionic compounds are soluble in water, some are not very soluble at all. The attraction between ions is difficult to break. As a result, compounds with very strong ionic bonds, such as silver chloride, tend to be less soluble in water than compounds with weak ionic bonds, such as sodium chloride.

Predicting Solubility

You can predict the solubility of a binary compound, such as mercury(II) sulfide, HgS, by comparing the electronegativity of each element in the compound. If there is a large difference in the two electronegativities, the bond between the elements is polar or even ionic. This type of compound probably dissolves in water. If there is only a small difference in the two electronegativities, the bond is not polar or ionic. This type of compound probably does not dissolve in water. For example, the electronegativity of mercury is 1.9. The electronegativity of sulfur is 2.5. The difference in these two electronegativities is small, only 0.6. Therefore, you can predict that mercury(II) sulfide is insoluble in water. In Chapter 9, you will learn another way to predict the solubility of ionic compounds in water.
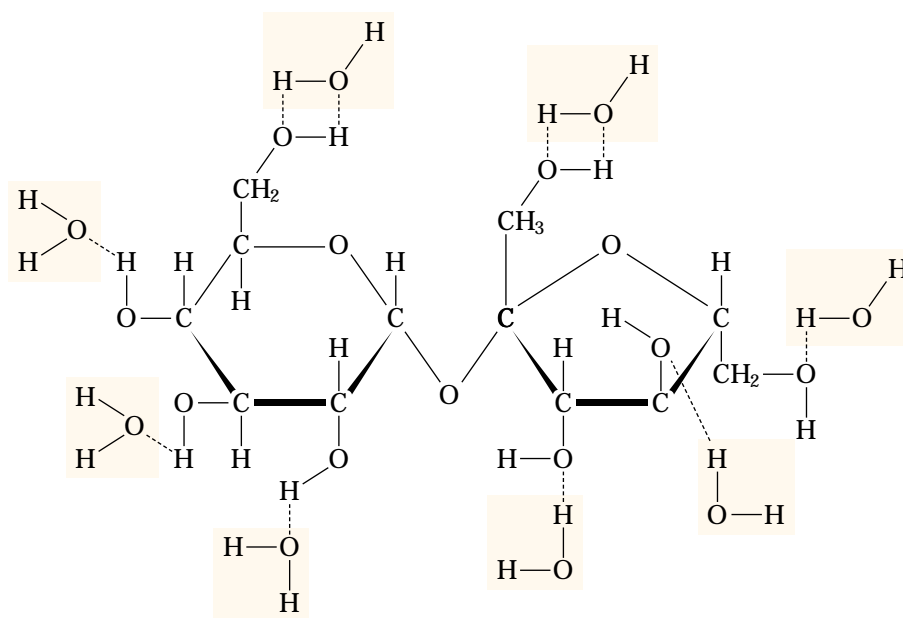
Electronic Learning Partner

The Chemistry 11 Electronic Learning Partner contains a video clip describing how water dissolves ionic and some covalent compounds. This will be useful if you are having difficulty visualizing particle attractions.

The Solubility of Covalent Compounds

Many covalent compounds do not have negative and positive charges to attract water molecules. Thus they are not soluble in water. There are some exceptions, however. Methanol (a component of windshield washer fluid), ethanol (the "alcohol" in alcoholic beverages), and sugars (such as sucrose) are examples of covalent compounds that are extremely soluble in water. These compounds dissolve because their molecules contain polar bonds, which are able to form hydrogen bonds with water.

For example, sucrose molecules have a number of sites that can form a hydrogen bond with water to replace the attraction between the sucrose molecules. (See Figure 8.11.) The sucrose molecules separate and become hydrated, just like dissolved ions. The molecules remain neutral, however. As a result, sucrose and other soluble covalent compounds do not conduct electricity when dissolved in water. They are **non-electrolytes**.

Figure 8.11 A sucrose molecule contains several O–H atom connections. The O–H bond is highly polar, with the H atom having the positive charge. The negative charges on water molecules form hydrogen bonds with a sucrose molecule, as shown by the dotted lines.

Insoluble Covalent Compounds

The covalent compounds that are found in oil and grease are insoluble in water. They have no ions or highly polar bonds, so they cannot form hydrogen bonds with water molecules. Non-polar compounds tend to be soluble in non-polar solvents, such as benzene or kerosene. The forces between the solute molecules are replaced by the forces between the solute and solvent molecules.

In general, *ionic solutes and polar covalent solutes both dissolve in polar solvents. Non-polar solutes dissolve in non-polar solvents.* The phrase *like dissolves like* summarizes these observations. It means that solutes and solvents that have similar properties form solutions.

If a compound has both polar and non-polar components, it may dissolve in both polar and non-polar solvents. For example, acetic acid, CH_3COOH, is a liquid that forms hydrogen bonds with water. It is fully miscible with water. Acetic acid also dissolves in non-polar solvents, such as benzene and carbon tetrachloride, because the CH_3 component is non-polar.

Factors That Affect Solubility

You have taken a close look at the attractive forces between solute and solvent particles. Now that you understand why solutes dissolve, it is time to examine the three factors that affect solubility: molecule size, temperature, and pressure. Notice that these three factors are similar to the factors that affect the rate of dissolving. Be careful not to confuse them.

Molecule Size and Solubility

Small molecules are often more soluble than larger molecules. Methanol, CH_3OH, and ethanol, CH_3CH_2OH, are both completely miscible with water. These compounds have OH groups that form hydrogen bonds with water. Larger molecules with the same OH group but more carbon atoms, such as pentanol, $CH_3CH_2CH_2CH_2CH_2OH$, are far less soluble. All three compounds form hydrogen bonds with water, but the larger pentanol is less polar overall, making it less soluble. Table 8.2 compares five molecules by size and solubility.

Table 8.2 Solubility and Molecule Size

Name of compound	methanol	ethanol	propanol	butanol	pentanol
Chemical formula	CH_3OH	CH_3CH_2OH	$CH_3CH_2CH_2OH$	$CH_3(CH_2)_3OH$	$CH_3(CH_2)_4OH$
Solubility	infinitely soluble	infinitely soluble	very soluble	9 g/100 mL (at 25°C)	3 g/100 mL (at 25°C)

Temperature and Solubility

At the beginning of this section, you learned that temperature affects the rate of dissolving. Temperature also affects solubility. You may have noticed that solubility data always include temperature. The solubility of a solute in water, for example, is usually given as the number of grams of solute that dissolve in 100 mL of water at a specific temperature. (See Table 8.2 for two examples.) Specifying temperature is essential, since the solubility of a substance is very different at different temperatures.

When a solid dissolves in a liquid, energy is needed to break the strong bonds between particles in the solid. At higher temperatures, more energy is present. Thus, *the solubility of most solids increases with temperature*. For example, caffeine's solubility in water is only 2.2 g/100 mL at 25°C. At 100°C, however, caffeine's solubility increases to 40 g/100 mL.

The bonds between particles in a liquid are not as strong as the bonds between particles in a solid. When a liquid dissolves in a liquid, additional energy is not needed. Thus, *the solubility of most liquids is not greatly affected by temperature*.

Gas particles move quickly and have a great deal of kinetic energy. When a gas dissolves in a liquid, it loses some of this energy. At higher temperatures, the dissolved gas gains energy again. As a result, the gas comes out of solution and is less soluble. Thus, *the solubility of gases decreases with higher temperatures*.

In the next investigation, you will observe and graph the effect of temperature on the solubility of a solid dissolved in a liquid solvent, water. As you have learned, most solid solutes become more soluble at higher temperatures. By determining the solubility of a solute at various temperatures, you can make a graph of solubility against temperature. The curve of best fit, drawn through the points, is called the *solubility curve*. You can use a solubility curve to determine the solubility of a solute at any temperature in the range shown on the graph.

CHEM FACT

The link between cigarette smoking and lung cancer is well known. Other cancers are also related to smoking. It is possible that a smoker who consumes alcohol may be at greater risk of developing stomach cancer. When a person smokes, a thin film of tar forms inside the mouth and throat. The tar from cigarette smoke contains many carcinogenic (cancer-causing) compounds. These compounds are non-polar and do not dissolve in saliva. They are more soluble in alcohol, however. As a result, if a smoker drinks alcohol, carcinogenic compounds can be washed into the stomach.

Plotting Solubility Curves

In this investigation, you will determine the temperature at which a certain amount of potassium nitrate is soluble in water. You will then dilute the solution and determine the solubility again. By combining your data with other students' data, you will be able to plot a solubility curve.

Question

What is the solubility curve of ?

Prediction

Draw a sketch to show the shape of the curve you expect for the solubility of a typical solid dissolving in water at different temperatures. Plot solubility on the y-axis and temperature on the x-axis.
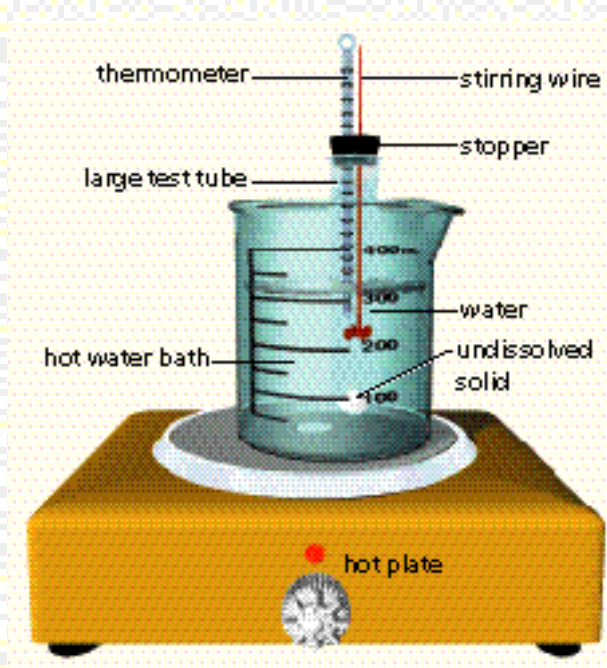
Safety Precautions

- Before lighting the Bunsen burner, check that there are no flammable solvents nearby. If you are using a Bunsen burner, tie back long hair and loose clothing. Be careful of the open flame.
- After turning it on, be careful not to touch the hot plate.

Materials

large test tube
balance
stirring wire
two-hole stopper to fit the test tube, with a
    thermometer inserted in one hole
400 mL beaker
graduated cylinder or pipette or burette
hot plate or Bunsen burner with ring clamps
    and wire gauze
retort stand and thermometer clamp
potassium nitrate,
distilled water

Procedure

1. Read through the steps in this Procedure. Prepare a data table to record the mass of the solute, the initial volume of water, the total volume of water after step 9, and the temperatures at which the solutions begin to crystallize.

2. Put the test tube inside a beaker for support. Place the beaker on a balance pan. Set the reading on the balance to zero. Then measure 14.0 g of potassium nitrate into the test tube.

3. Add one of the following volumes of distilled water to the test tube, as assigned by your teacher: 10.0 mL, 15.0 mL, 20.0 mL, 25.0 mL, 30.0 mL. (If you use a graduated cylinder, remember to read the volume from the bottom of the water meniscus. You can make a more accurate volume measurement using either a burette or a pipette.)

4. Pour about 300 mL of tap water into the beaker. Set up a hot-water bath using a hot plate, retort stand, and thermometer clamp. Alternatively, use a Bunsen burner, retort stand, ring clamp, thermometer clamp, and wire gauze.

5. Put the stirring wire through the second hole of the stopper. Insert the stopper, thermometer, and wire into the test tube. Make sure that the thermometer bulb is below the surface of the solution. (Check the diagram on the next page to make sure that you have set up the apparatus properly thus far.)

6. Place the test tube in the beaker. Secure the test tube and thermometer to the retort stand, using clamps. Begin heating the water bath gently.

7. Using the stirring wire, stir the mixture until the solute completely dissolves. Turn the heat source off, and allow the solution to cool.

Figure labels: thermometer, stirring wire, stopper, large test tube, hot water bath, water, undissolved solid, hot plate

8. Continue stirring. Record the temperature at which crystals begin to appear in the solution.

9. Remove the stopper from the test tube. Carefully add 5.0 mL of distilled water. The solution is now more dilute and therefore more soluble. Crystals will appear at a lower temperature.

10. Put the stopper, with the thermometer and stirring wire, back in the test tube. If crystals have already started to appear in the solution, begin warming the water bath again. Repeat steps 7 and 8.

11. If no crystals are present, stir the solution while the water bath cools. Record the temperature at which crystals first begin to appear.

12. Dispose of the aqueous solutions of potassium nitrate into the labelled waste container.

Analysis

1. Use the volume of water assigned by your teacher to calculate how much solute dissolved in 100 mL of water. Use the following equation to help you:

This equation represents the solubility of at the temperature at which you recorded the first appearance of crystals. Repeat your calculation to determine the solubility after the solution was diluted. Your teacher will collect and display all the class data for this investigation.

2. Some of your classmates were assigned the same volume of water that you were assigned. Compare the temperatures they recorded for their solutions with the temperatures you recorded. Comment on the precision of the data. Should any data be removed before averaging?

3. Average the temperatures at which crystal formation occurs for solutions that contain the same volume of water. Plot these data on graph paper. Set up your graph sideways on the graph paper (landscape orientation). Plot solubility on the vertical axis. (The units are grams of solute per 100 mL of water.) Plot temperature on the horizontal axis.

4. Draw the best smooth curve through the points. (Do not simply join the points.) Label each axis. Give the graph a suitable title.

Conclusions

5. Go back to the sketch you drew to predict the solubility of a typical solid dissolving in water at different temperatures. Compare the shape of your sketch with the shape of your graph.

6. Use your graph to *interpolate* the solubility of potassium nitrate at
   (a) 60°C   (b) 40°C

7. Use your graph to *extrapolate* the solubility of potassium nitrate at
   (a) 80°C   (b) 20°C

Application

8. At what temperature can 40 mL of water dissolve the following quantities of potassium nitrate?
   (a) 35.0 g   (b) 20.0 g

Heat Pollution: A Solubility Problem

For most solids, and almost all ionic substances, solubility increases as the temperature of the solution increases. Gases, on the other hand, always become *less* soluble as the temperature increases. This is why a refrigerated soft drink tastes fizzier than the same drink at room temperature. The warmer drink contains less dissolved carbon dioxide than the cooler drink.

This property of gases makes heat pollution a serious problem. Many industries and power plants use water to cool down overheated machinery. The resulting hot water is then returned to local rivers or lakes. Figure 8.12 shows steam rising from a "heat-polluted" river. Adding warm water into a river or lake does not seem like actual pollution. The heat from the water, however, increases the temperature of the body of water. As the temperature increases, the dissolved oxygen in the water decreases. Fish and other aquatic wildlife and plants may not have enough oxygen to breathe.

The natural heating of water in rivers and lakes can pose problems, too. Fish in warmer lakes and rivers are particularly vulnerable in the summer. When the water warms up even further, the amount of dissolved oxygen decreases.

Figure 8.12 This image shows the result of heat pollution. Warmer water contains less dissolved oxygen.

ExpressLab — The Effect of Temperature on Soda Water

In this Express Lab, you will have a chance to see how a change in temperature affects the dissolved gas in a solution. You will be looking at the pH of soda water. A low pH (1–6) indicates that the solution is acidic. You will learn more about pH in Chapter 10.

Safety Precautions

- If using a hot plate, avoid touching it when it is hot.
- If using a Bunsen burner, check that there are no flammable solvents nearby.

Procedure

1. Open a can of cool soda water. (Listen for the sound of excess carbon dioxide escaping.) Pour about 50 mL into each of two 100 mL beakers. Note the rate at which bubbles form. Record your observations.

2. Add a few drops of universal indicator to both beakers. Record the colour of the solutions. Then estimate the pH.

3. Measure and record the mass of each beaker. Measure and record the temperature of the soda water.

4. Place one beaker on a heat source. Heat it to about 50°C. Compare the rate of formation of the bubbles with the rate of formation in the beaker of cool soda water. Record any change in colour in the heated solution. Estimate its pH.

5. Allow the heated solution to cool. Again record any change in colour in the solution. Estimate its pH.

6. Measure and record the masses of both beakers. Determine any change in mass by comparing the final and initial masses.

Analysis

1. Which sample of soda water lost the most mass? Explain your observation.

2. Did the heated soda water become more or less acidic when it was heated? Explain why you think this change happened.

Pressure and Solubility

The final factor that affects solubility is pressure. Changes in pressure have hardly any effect on solid and liquid solutions. Such changes do affect the solubility of a gas in a liquid solvent, however. The solubility of the gas is directly proportional to the pressure of the gas above the liquid. For example, the solubility of oxygen in lake water depends on the air pressure above the lake.

When you open a carbonated drink, you can observe the effect of pressure on solubility. Figure 8.13 shows this effect. Inside a soft drink bottle, the pressure of the carbon dioxide gas is very high: about 400 kPa. When you open the bottle, you hear the sound of escaping gas as the pressure is reduced. Carbon dioxide gas escapes quickly from the bottle, since the pressure of the carbon dioxide in the atmosphere is much lower: only about 0.03 kPa. The solubility of the carbon dioxide in the liquid soft drink decreases greatly. Bubbles begin to rise in the liquid as gas comes out of solution and escapes. It takes a while for all the gas to leave the solution, so you have time to enjoy the taste of the soft drink before it goes "flat."

Figure 8.14 illustrates another example of dissolved gases and pressure. As a scuba diver goes deeper underwater, the water pressure increases. The solubility of nitrogen gas, which is present in the lungs, also increases. Nitrogen gas dissolves in the diver's blood. As the diver returns to the surface, the pressure acting on the diver decreases. The nitrogen gas in the blood comes out of solution. If the diver surfaces too quickly, the effect is similar to opening a soft drink bottle. Bubbles of nitrogen gas form in the blood. This leads to a painful and sometimes fatal condition known as "the bends." You will learn more about gases and deep-sea diving in Chapter 11.

Figure 8.13 What happens when the pressure of the carbon dioxide gas in a soft drink bottle is released? The solubility of the gas in the soft drink solution decreases.

Figure 8.14 Scuba divers must heed the effects of decreasing water pressure on dissolved nitrogen gas in their blood. They must surface slowly to avoid "the bends."

CHEM FACT

Do you crack your knuckles? The sound you hear is another example of the effect of pressure on solubility. Joints contain fluid. When a joint is suddenly pulled or stretched, the cavity that holds the fluid gets larger. This causes the pressure to decrease. A bubble of gas forms, making the sound you hear. You cannot repeatedly crack your knuckles because it takes some time for the gas to re-dissolve.

Chemistry Bulletin

Science | Technology | Society | Environment

Solvents and Coffee: What's the Connection?

The story of coffee starts with the coffee berry. First the pulp of the berry is removed. This leaves two beans, each containing 1% to 2% caffeine. The beans are soaked in water and natural enzymes to remove the outer parchment husk and to start a slight fermentation process. Once the beans have been fermented, they are dried and roasted. Then the coffee is ready for grinding. Grinding increases the surface area of the coffee. Thus, finer grinds make it easier to dissolve the coffee in hot water.

Decaffeinated coffee satisfies people who like the smell and taste of coffee but cannot tolerate the caffeine. How is caffeine removed from coffee?

All the methods of extracting caffeine take place before the beans are roasted. Caffeine and the other organic compounds that give coffee its taste are mainly non-polar. (Caffeine does contain some polar bonds, however, which allows it to dissolve in hot water.) Non-polar solvents, such as benzene and trichloroethene, were once used to dissolve and remove caffeine from the beans. These chemicals are now considered to be too hazardous. Today most coffee manufacturers use water or carbon dioxide as solvents.

In the common Swiss Water Process, coffee beans are soaked in hot water. This dissolves the caffeine and the flavouring compounds from the beans. The liquid is passed through activated carbon filters. The filters retain the caffeine, but let the flavouring compounds pass through. The filtered liquid, now caffeine-free, is sprayed back onto the beans. The beans reabsorb the flavouring compounds. Now they are ready for roasting.

Carbon dioxide gas is a normal component of air. In the carbon dioxide decaffeination process, the gas is raised to a temperature of at least 32°C. Then it is compressed to a pressure of about 7400 kPa. At this pressure, it resembles a liquid but can flow like a gas. The carbon dioxide penetrates the coffee beans and dissolves the caffeine. When the pressure returns to normal, the carbon dioxide reverts to a gaseous state. The caffeine is left behind.

What happens to the caffeine that is removed by decaffeination? Caffeine is so valuable that it is worth more than the cost of taking it out of the beans. It is extensively used in the pharmaceutical industry, and for colas and other soft drinks.

Making Connections

1. As you have read, water is a polar liquid and the soluble fractions of the coffee grounds are non-polar. Explain, in chemical terms, how caffeine and the coffee flavour and aroma are transferred to hot coffee.

2. Why does hot water work better in the brewing process than cold water?

3. In chemical terms, explain why fine grinds of coffee make better coffee.

How can caffeine form hydrogen bonds with water?

Section Wrap-up

In this section, you examined the factors that affect the rate of dissolving: temperature, agitation, and particle size. Next you looked at the forces between solute and solvent particles. Finally, you considered three main factors that affect solubility: molecule size, temperature, and pressure. In section 8.3, you will learn about the effects of differing amounts of solute dissolved in a certain amount of solvent.

Section Review

1 **K/U** Describe the particle attractions that occur as sodium chloride dissolves in water.

2 **K/U** When water vaporizes, which type of attraction, intramolecular or intermolecular, is broken? Explain.

3 **K/U** Describe the effect of increasing temperature on the solubility of

(a) a typical solid in water

(b) a gas in water

4 **K/U** Sugar is more soluble in water than salt. Why does a salt solution (brine) conduct electricity, while a sugar solution does not?

5 **K/U** Dissolving a certain solute in water releases heat. Dissolving a different solute in water absorbs heat. Explain why.

6 **I** The graph below shows the solubility of various substances plotted against the temperature of the solution.

(a) Which substance decreases in solubility as the temperature increases?

(b) Which substance is least soluble at room temperature? Which substance is most soluble at room temperature?

(c) The solubility of which substance is least affected by a change in temperature?

(d) At what temperature is the solubility of potassium chlorate equal to 40 g/100 g of water?

(e) 20 mL of a saturated solution of potassium nitrate at 50°C is cooled to 20°C. Approximately what mass of solid will precipitate from the solution? Why is it not possible to use the graph to interpolate an accurate value?

7 **I** A saturated solution of potassium nitrate was prepared at 70°C and then cooled to 55°C. Use your graph from Investigation 8-A to predict the fraction of the dissolved solute that crystallized out of the solution.

8 **MC** Would you expect to find more mineral deposits near a thermal spring or near a cool mountain spring? Explain.
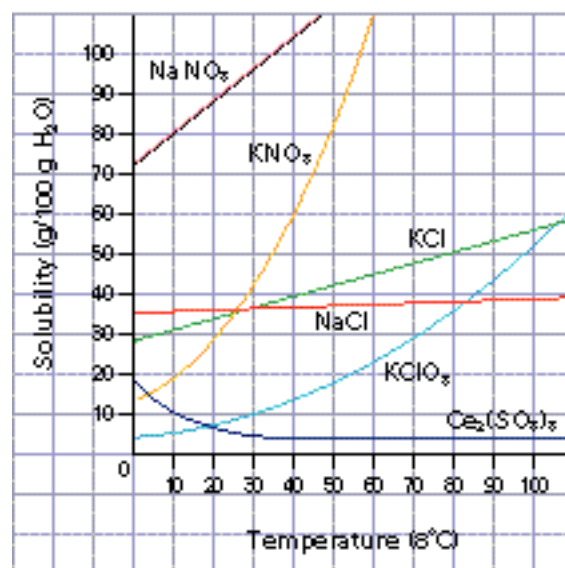
> **Unit Issue Prep**
>
> Think about how the properties of water affect its behaviour in the environment. Look ahead to the Unit 3 Issue. How could water's excellent ability as a solvent become a problem?