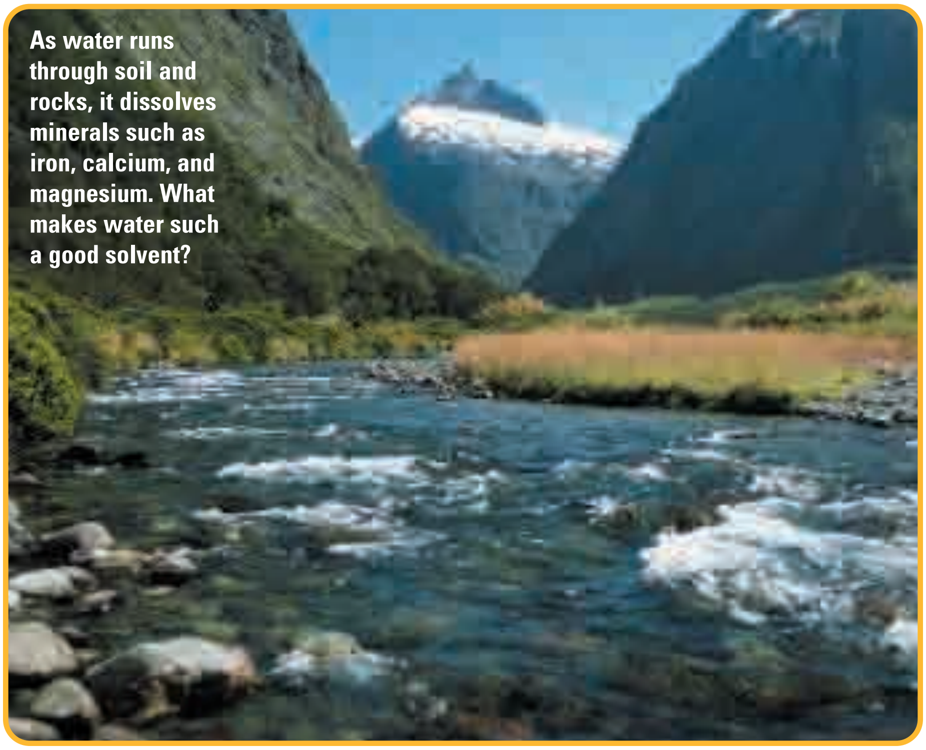
CHAPTER 8

Solutions and Their Concentrations

Your environment is made up of many important solutions, or homogeneous mixtures. The air you breathe and the liquids you drink are solutions. So are many of the metallic objects that you use every day. The quality of a solution, such as tap water, depends on the substances that are dissolved in it. "Clean" water may contain small amounts of dissolved substances, such as iron and chlorine. "Dirty" water may have dangerous chemicals dissolved in it.

The difference between clean water and undrinkable water often depends on concentration: the amount of a dissolved substance in a particular quantity of a solution. For example, tap water contains a low concentration of fluoride to help keep your teeth healthy. Water with a high concentration of fluoride, however, could be harmful to your health.

Water is a good solvent for many substances. You may have noticed, however, that grease-stained clothing cannot be cleaned by water alone. Grease is one substance that does not dissolve in water. Why doesn't it dissolve? In this chapter, you will find out why. You will learn how solutions form. You will explore factors that affect a substance's ability to dissolve. You will find out more about the concentration of solutions, and you will have a chance to prepare your own solutions as well.



As water runs through soil and rocks, it dissolves minerals such as iron, calcium, and magnesium. What makes water such a good solvent?

Chapter Preview

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- classifying mixtures (Chapter 1, section 1.3)
- predicting molecular polarity (Chapter 3, section 3.3)
- distinguishing between intermolecular and intramolecular forces (Chapter 3, section 3.2)
- describing the shape and bonding of the water molecule (Chapter 3, section 3.3)
- calculating molar mass (Chapter 5, section 5.3)
- calculating molar amounts (Chapter 5, section 5.3)

8.1 Types of Solutions

Section Preview/ Specific Expectations

In this section, you will

- **explain** solution formation, referring to polar and non-polar solvents
- **identify examples** of solid, liquid, and gas solutions from everyday life
- **communicate** your understanding of the following terms: *solution*, *solvent, solutes, variable composition, aqueous solution, miscible, immiscible, alloys, solubility, saturated solution, unsaturated solution*

As stated in the chapter opener, a **solution** is a homogeneous mixture. It is uniform throughout. If you analyze any two samples of a solution, you will find that they contain the same substances in the same relative amounts. The simplest solutions contain two substances. Most common solutions contain many substances.

A **solvent** is any substance that has other substances dissolved in it. In a solution, the substance that is present in the largest amount (whether by volume, mass, or number of moles) is usually referred to as the solvent. The other substances that are present in the solution are called the **solutes**.

Pure substances (such as pure water, H_2O) have fixed composition. You cannot change the ratio of hydrogen, H, to oxygen, O, in water without producing an entirely new substance. Solutions, on the other hand, have **variable composition**. This means that different ratios of solvent to solute are possible. For example, you can make a weak or a strong solution of sugar and water, depending on how much sugar you add. Figure 8.1 shows a strong solution of tea and water on the left, and a weak solution of tea and water on the right. The ratio of solvent to solute in the strong solution is different from the ratio of solvent to solute in the weak solution.

Figure 8.1 How can a solution have variable composition yet be uniform throughout?

When a solute dissolves in a solvent, no chemical reaction occurs. Therefore, the solute and solvent can be separated using physical properties, such as boiling point or melting point. For example, water and ethanol have different boiling points. Using this property, a solution of water and ethanol can be separated by the process of distillation. Refer back to Chapter 1, section 1.2. What physical properties, besides boiling point, can be used to separate the components of solutions and other mixtures?

COURSE CHALLENGE

How can you find out what solutes are dissolved in a sample of water? What physical properties might be useful? You will need answers to these questions when you do your Chemistry Course Challenge.

Figure 8.2 **Can you identify the components of some of these solutions?**

A solution can be a gas, a liquid, or a solid. Figure 8.2 shows some examples of solutions. Various combinations of solute and solvent states are possible. For example, a gas can be dissolved in a liquid, or a solid can be dissolved in another solid. Solid, liquid, and gaseous solutions are all around you. Steel is a solid solution of carbon in iron. Juice is a liquid solution of sugar and flavouring dissolved in water. Air is an example of a gaseous solution. The four main components of dry air are nitrogen (78%), oxygen (21%), argon (0.9%), and carbon dioxide (0.03%). Table 8.1 lists some other common solutions.

Table 8.1 Types of Solutions

Original state of solute	Solvent	Examples
gas	gas	air; natural gas; oxygen-acetylene mixture used in welding
gas	liquid	carbonated drinks; water in rivers and lakes containing oxygen
gas	solid	hydrogen in platinum
liquid	gas	water vapour in air; gasoline-air mixture
liquid	liquid	alcohol in water; antifreeze in water
liquid	solid	amalgams, such as mercury in silver
solid	gas	mothballs in air
solid	liquid	sugar in water; table salt in water; amalgams
solid	solid	alloys, such as the copper-nickel alloy used to make coins

mind STRETCH

Take another look at the four components of dry air. Which component would you call the solvent? Which components are the solutes?

Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner can help you learn more about the properties of water.

CHEM FACT

An alloy that is made of a metal dissolved in mercury is called an *amalgam*. A traditional dental amalgam, used to fill cavities in teeth, contains 50% mercury. Due to concern over the use of mercury, which is toxic, dentists now use other materials, such as ceramic materials, to fill dental cavities.

You are probably most familiar with liquid solutions, especially aqueous solutions. An **aqueous solution** is a solution in which water is the solvent. Because aqueous solutions are so important, you will focus on them in the next two sections of this chapter and again in Chapter 9.

Some liquids, such as water and ethanol, dissolve readily in each other in any proportion. That is, any amount of water dissolves in any amount of ethanol. Similarly, any amount of ethanol dissolves in any amount of water. Liquids such as these are said to be **miscible** with each other. Miscible liquids can be combined in any proportions. Thus, either ethanol or water can be considered to be the solvent. Liquids that do not readily dissolve in each other, such as oil and water, are said to be **immiscible**.

As you know from Chapter 4, solid solutions of metals are called **alloys**. Adding even small quantities of another element to a metal changes the properties of the metal. Technological advances throughout history have been linked closely to the discovery of new alloys. For example, bronze is an alloy of copper and tin. Bronze contains only about 10% tin, but it is much stronger than copper and more resistant to corrosion. Also, bronze can be melted in an ordinary fire so that castings can be made, as shown in Figure 8.3.

Solubility and Saturation

The ability of a solvent to dissolve a solute depends on the forces of attraction between the particles. There is always some attraction between solvent and solute particles, so some solute always dissolves. The **solubility** of a solute is the amount of solute that dissolves in a given quantity of solvent, at a certain temperature. For example, the solubility of sodium chloride in water at 20°C is 36 g per 100 mL of water.

A **saturated solution** is formed when no more solute will dissolve in a solution, and excess solute is present. For example, 100 mL of a saturated solution of table salt (sodium chloride, NaCl) in water at 20°C contains 36 g of sodium chloride. The solution is saturated with respect to sodium chloride. If more sodium chloride is added to the solution, it will not dissolve. The solution may still be able to dissolve other solutes, however.

An **unsaturated solution** is a solution that is not yet saturated. Therefore, it can dissolve more solute. For example, a solution that contains 20 g of sodium chloride dissolved in 100 mL of water at 20°C is unsaturated. This solution has the potential to dissolve another 16 g of salt, as Figure 8.4 demonstrates.

Figure 8.3 The introduction of the alloy bronze around 3000 BCE led to the production of better-quality tools and weapons.

Figure 8.4 **At 20°C, the solubility of table salt in water is 36 g/100 mL.**

A 20 g of NaCl dissolve to form an unsaturated solution.

B 36 g of NaCl dissolve to form a saturated solution.

C 40 g of NaCl are added to 100 mL of water. 36 g dissolve to form a saturated solution. 4 g of undissolved solute are left.

Suppose that a solute is described as *soluble* in a particular solvent. This generally means that its solubility is greater than 1 g per 100 mL of solvent. If a solute is described as *insoluble*, its solubility is less than 0.1 g per 100 mL of solvent. Substances with solubility between these limits are called *sparingly soluble*, or *slightly soluble*. Solubility is a relative term, however. Even substances such as oil and water dissolve in each other to some extent, although in very tiny amounts.

The general terms that are used to describe solubility for solids and liquids do not apply to gases in the same way. For example, oxygen is described as soluble in water. Oxygen from the air dissolves in the water of lakes and rivers. The solubility of oxygen in fresh water at 20°C is only 9 mg/L, or 0.0009 g/100 mL. This small amount of oxygen is enough to ensure the survival of aquatic plants and animals. A solid solute with the same solubility, however, would be described as insoluble in water.

mind STRETCH

Imagine that you are given a filtered solution of sodium chloride. How can you decide whether the solution is saturated or unsaturated?

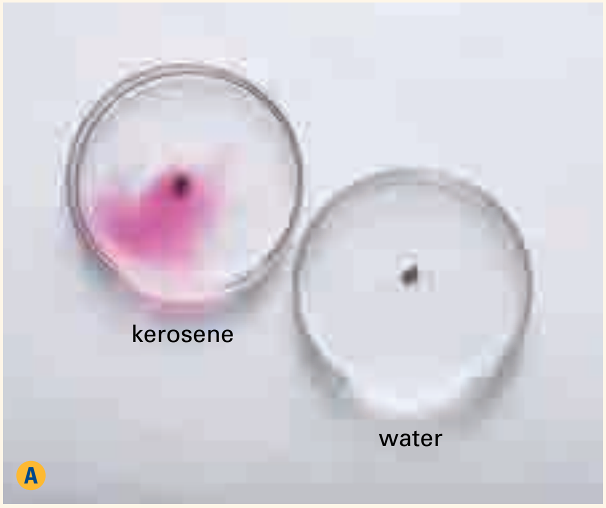
Identifying Suitable Solvents

Water is a good solvent for many compounds, but it is a poor solvent for others. If you have grease on your hands after adjusting a bicycle chain, you cannot use water to dissolve the grease and clean your hands. You need to use a detergent, such as soap, to help dissolve the grease in the water. You can also use another solvent to dissolve the grease. How can you find a suitable solvent? How can you predict whether a solvent will dissolve a particular solute? Try the Thought Lab on the next page to find out for yourself.

Electronic Learning Partner

Go to your Chemistry 11 Electronic Learning Partner to find out more about the properties of two solvents: water and benzene.

ThoughtLab Matching Solutes and Solvents



Although there is a solvent for every solute, not all mixtures produce a solution. Table salt dissolves in water but not in kerosene. Oil dissolves in kerosene but not in water. What properties must a solvent and a solute share in order to produce a solution?

In an investigation, the bottom of a Petri dish was covered with water, as shown in photo A. An equal amount of kerosene was added to a second Petri dish. When a crystal of iodine was added to the water, it did not dissolve. When a second crystal of iodine was added to the kerosene, however, it did dissolve.

Procedure

Classify each compound as ionic (containing ions), polar (containing polar molecules), or non-polar (containing non-polar molecules).

(a) iodine, I_2

(b) cobalt(II) chloride, $CoCl_2$

(c) sucrose, $C_{12}H_{22}O_{11}$ (**Hint:** Sucrose contains 8 O–H bonds.)

In photo B, the same experiment was repeated with crystals of cobalt(II) chloride. This time, the crystal dissolved in the water but not in the kerosene.

Analysis

1. Water is a polar molecule. Therefore, it acts as a polar solvent.
   (a) Think about the compounds you classified in the Procedure. Which compounds are soluble in water?
   (b) Assume that the interaction between solutes and solvents that you examine here applies to a wide variety of substances. Make a general statement about the type of solute that dissolves in polar solvents.
2. Kerosene is non-polar. It acts as a non-polar solvent.
   (a) Which of the compounds you classified in the Procedure is soluble in kerosene?
   (b) Make a general statement about the type of solute that dissolves in non-polar solvents.

Section Wrap-up

In this section, you learned the meanings of several important terms, such as *solvent*, *solute*, *saturated solution*, *unsaturated solution*, *aqueous solution*, and *solubility*. You need to know these terms in order to understand the material in the rest of the chapter. In section 8.2, you will examine the factors that affect the rate at which a solute dissolves in a solvent. You will also learn about factors that affect solubility.

Section Review

1. K/U Name the two basic components of a solution.
2. K/U Give examples of each type of solution.
   (a) solid solution
   (b) liquid solution
   (c) gaseous solution (at room temperature)
3. K/U Explain the term "homogeneous mixture."
4. C How do the properties of a homogeneous mixture differ from the properties of a heterogeneous mixture, or mechanical mixture? Use diagrams to explain.
5. K/U Give examples of each type of mixture.
   (a) homogeneous mixture
   (b) mechanical mixture (heterogeneous mixture)
6. K/U Distinguish between the following terms: soluble, miscible, and immiscible.
7. K/U Distinguish between an alloy and an amalgam. Give one example of each.
8. K/U What type of solute dissolves in a polar solvent, such as water? Give an example.
9. I Potassium bromate, $KBrO_3$, is sometimes added to bread dough to make it easier to work with. Suppose that you are given an aqueous solution of potassium bromate. How can you determine if the solution is saturated or unsaturated?
10. K/U Two different clear, colourless liquids were gently heated in an evaporating dish. Liquid A left no residue, while liquid B left a white residue. Which liquid was a solution, and which was a pure substance? Explain your answer.
11. I You are given three liquids. One is a pure substance, and the second is a solution of two miscible liquids. The third is a solution composed of a solid solute dissolved in a liquid solvent. Describe the procedure you would follow to distinguish between the three solutions.
12. MC In 1989, the oil tanker Exxon Valdez struck a reef in Prince William Sound, Alaska. The accident released 40 million litres of crude oil. The oil eventually covered 26 000 km^2 of water.
    (a) Explain why very little of the spilt oil dissolved in the water.
    (b) The density of crude oil varies. Assuming a value of 0.86 g/mL, estimate the average thickness of the oil slick that resulted from the Exxon Valdez disaster.
    (c) How do you think most of the oil from a tanker accident is dispersed over time? Why would this have been a slow process in Prince William Sound?
13. MC Food colouring is often added to foods such as candies, ice cream, and icing. Are food colouring dyes more likely to be polar or non-polar molecules? Explain your answer.

8.2 Factors That Affect Rate of Dissolving and Solubility

Section Preview/ Specific Expectations

In this section, you will

- **explain** some important properties of water
- **explain** solution formation in terms of intermolecular forces between polar, ionic, and non-polar substances
- **describe** the relationship between solubility and temperature for solids, liquids, and gases
- **communicate** your understanding of the following terms: *rate of dissolving, dipole, dipole-dipole attraction, hydrogen bonding, ion-dipole attractions, hydrated, electrolyte, non-electrolytes*

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.

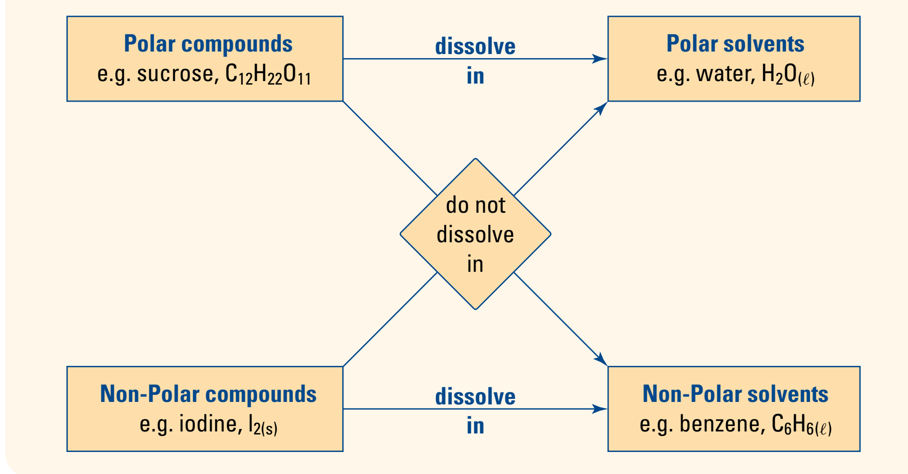
Math LINK

Calculate the surface area of a cube with the dimensions 5.0 cm × 5.0 cm × 5.0 cm. Now imagine cutting this cube to form smaller cubes with the dimensions 1.0 cm × 1.0 cm × 1.0 cm. How many smaller cubes could you make? Calculate their total surface area.

CHECKPOINT

Remember that the *rate* at which a solute dissolves is different from the *solubility* of the solute. In your notebook, explain briefly and clearly the difference between rate of dissolving and solubility.

Concept Organizer — Polar and Non-Polar Compounds



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.

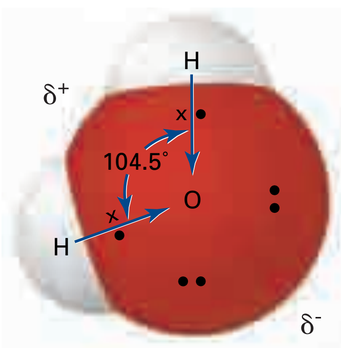


Figure 8.7 **The bent shape and polar bonds of a water molecule give it a permanent dipole.**

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.

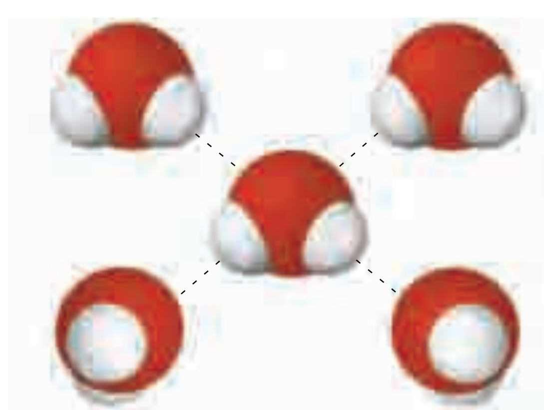


Figure 8.8 **Hydrogen bonding between water molecules is shown as dotted lines. The H atoms on each molecule are attracted to O atoms on other water molecules.**

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.

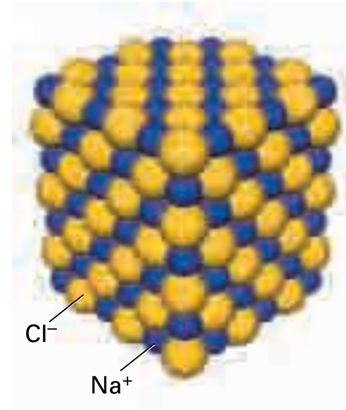


Figure 8.9 Ionic crystals have very ordered structures.

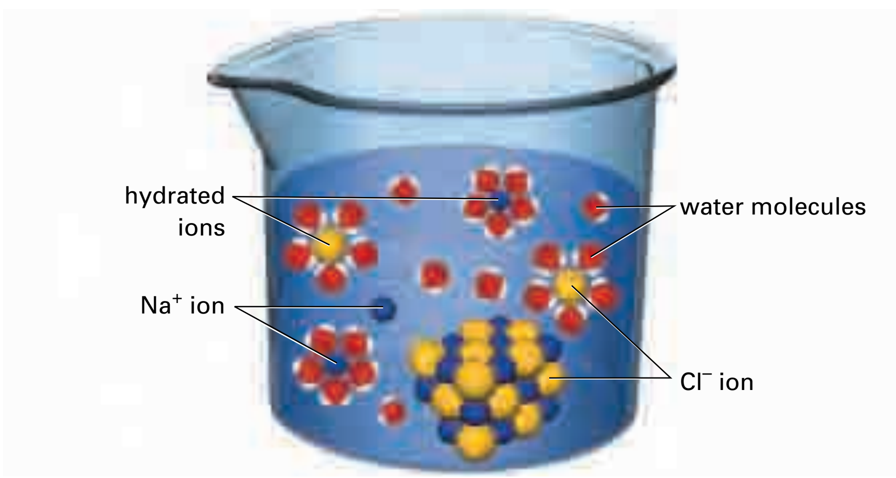


Figure 8.10 Ion-dipole attractions help to explain why sodium chloride dissolves in water.

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.

CHECKPOINT

Look back at the Concept Organizer on page 292. Where do ionic compounds belong in this diagram?

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**.

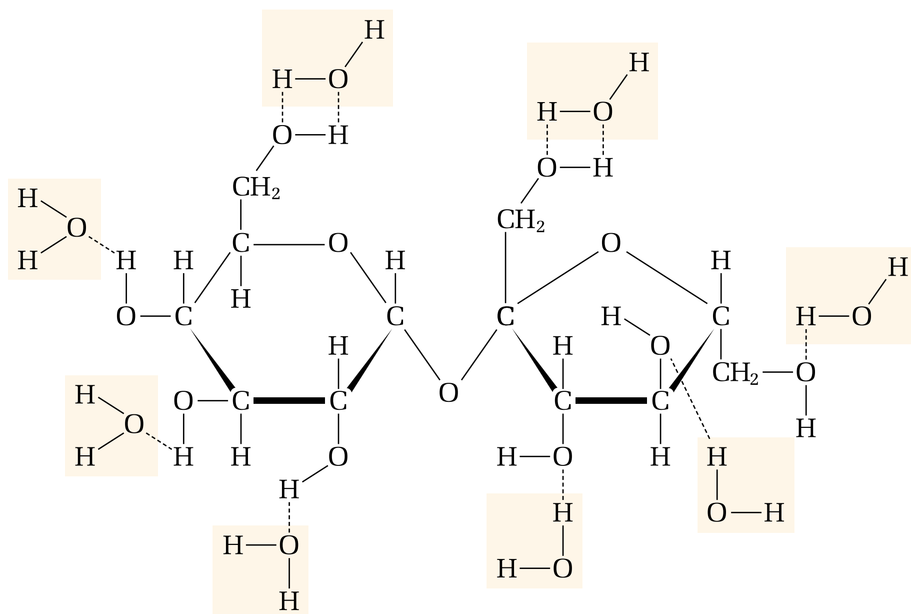


Figure 8.11 A sucrose molecule contains several O–H atom connections. The O–H bond is highly polar, with the H atom having the positive charge. The negative charges on water molecules form hydrogen bonds with a sucrose molecule, as shown by the dotted lines.

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.

Investigation 8-A

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

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 KNO_3?

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

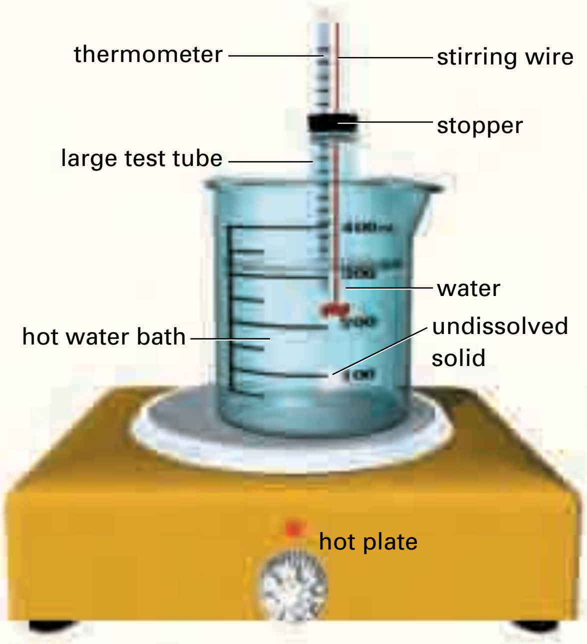
- Before lighting the Bunsen burner, check that there are no flammable solvents nearby. If you are using a Bunsen burner, tie back long hair and loose clothing. Be careful of the open flame.
- After turning it on, be careful not to touch the hot plate.

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, KNO_3
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.



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:
$$\frac{x\text{ g}}{100\text{ mL}} = \frac{14.0\text{ g}}{\text{your volume}}$$
This equation represents the solubility of KNO_3 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

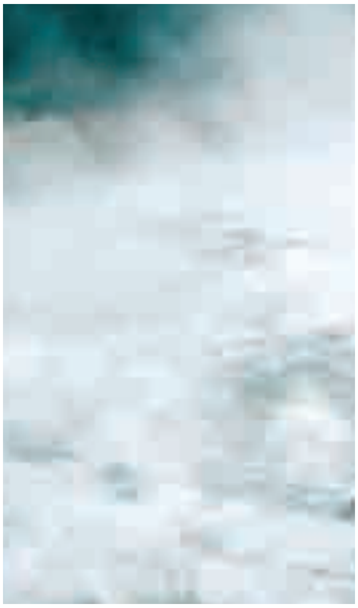


Figure 8.12 This image shows the result of heat pollution. Warmer water contains less dissolved oxygen.

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.

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.

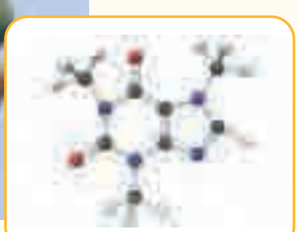
In the common Swiss Water Process, coffee beans are soaked in hot water. This dissolves the caffeine and the flavouring compounds from the beans. The liquid is passed through activated carbon filters. The filters retain the caffeine, but let the flavouring compounds pass through. The filtered liquid, 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.

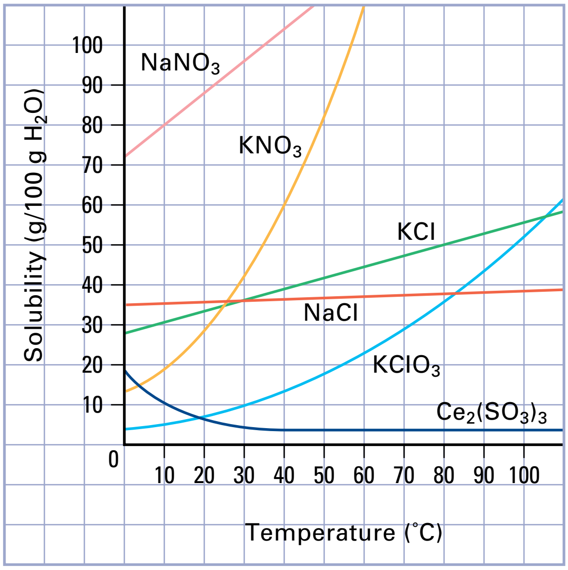
2 K/U When water vaporizes, which type of attraction, intramolecular or intermolecular, is broken? Explain.

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?

8.3 The Concentration of Solutions

Section Preview/ Specific Expectations

In this section, you will

- **solve problems** involving the concentration of solutions
- **express** concentration as grams per 100 mL, mass and volume percents, parts per million and billion, and moles per litre
- **communicate** your understanding of the following terms: *concentration, mass/volume percent, mass/mass percent, volume/volume percent, parts per million, parts per billion, molar concentration*

Material Safety Data Sheet

Component Name	CAS Number
PHENOL, 100% Pure	108952

SECTION III: Hazards Identification

- Very hazardous in case of ingestion, inhalation, skin contact, or eye contact.
- Product is corrosive to internal membranes when ingested.
- Inhalation of vapours may damage central nervous system. Symptoms: nausea, headache, dizziness.
- Skin contact may cause itching and blistering.
- Eye-contact may lead to corneal damage or blindness.
- Severe over-exposure may lead to lung-damage, choking, or coma.

Figure 8.15 Should phenol be banned from drugstores?

Phenol is a hazardous liquid, especially when it is at room temperature. It is a volatile chemical. Inhaling phenol adversely affects the central nervous system, and can lead to a coma. Inhalation is not the only danger. Coma and death have been known to occur within 10 min after phenol has contacted the skin. Also, as little as 1 g of phenol can be fatal if swallowed.

Would you expect to find such a hazardous chemical in over-the-counter medications? Check your medicine cabinet at home. You may find phenol listed as an ingredient in throat sprays and in lotions to relieve itching. You may also find it used as an antiseptic or disinfectant. Is phenol a hazard or a beneficial ingredient in many medicines? This depends entirely on **concentration**: the amount of solute per quantity of solvent. At high concentrations, phenol can kill. At low concentrations, it is a safe component of certain medicines.

Modern analytical tests allow chemists to detect and measure almost any chemical at extremely low concentrations. In this section, you will learn about various ways that chemists use to express the concentration of a solution. As well, you will find the concentration of a solution by experiment.

Concentration as a Mass/Volume Percent

Recall that the solubility of a compound at a certain temperature is often expressed as the mass of the solute per 100 mL of solvent. For example, you know that the solubility of sodium chloride is 36 g/100 mL of water at room temperature. The final volume of the sodium chloride solution may or may not be 100 mL. It is the volume of the solvent that is important.

Chemists often express the concentration of an *unsaturated* solution as the mass of solute dissolved per volume of the *solution*. This is different from solubility. It is usually expressed as a percent relationship. A **mass/volume percent** gives the mass of solute dissolved in a volume of solution, expressed as a percent. The mass/volume percent is also referred to as the *percent (m/v)*.

$$\text{Mass/volume percent} = \frac{\text{Mass of solute (in g)}}{\text{Volume of solution (in mL)}} \times 100\%$$

Suppose that a hospital patient requires an intravenous drip to replace lost body fluids. The intravenous fluid may be a saline solution that contains 0.9 g of sodium chloride dissolved in 100 mL of solution, or 0.9% (m/v). *Notice that the number of grams of solute per 100 mL of solution is numerically equal to the mass/volume percent.* Explore this idea further in the following problems.

PROBEWARE

If you have access to probeware, do the Chemistry 11 lab, "Concentration of Solutions" now.

Sample Problem

Solving for a Mass/Volume Percent

Problem

A pharmacist adds 2.00 mL of distilled water to 4.00 g of a powdered drug. The final volume of the solution is 3.00 mL. What is the concentration of the drug in g/100 mL of solution? What is the percent (m/v) of the solution?

What Is Required?

You need to calculate the concentration of the solution, in grams of solute dissolved in 100 mL of solution. Then you need to express this concentration as a mass/volume percent.

What Is Given?

The mass of the dissolved solute is 4.00 g. The volume of the solution is 3.00 mL.

Plan Your Strategy

There are two possible methods for solving this problem.

Method 1

Use the formula

$$\text{Mass/volume percent} = \frac{\text{Mass of solute (in g)}}{\text{Volume of solution (in mL)}} \times 100\%$$

Method 2

Let x represent the mass of solute dissolved in 100 mL of solution. The ratio of the dissolved solute, x, in 100 mL of solution must be the same as the ratio of 4.00 g of solute dissolved in 3.00 mL of solution. The concentration, expressed in g/100 mL, is numerically equal to the percent (m/v) of the solution.

Act on Your Strategy

Method 1

$$\begin{aligned}\text{Percent (m/v)} &= \frac{4.00\text{ g}}{3.00\text{ mL}} \times 100\% \\ &= 133\%\end{aligned}$$

Continued ...

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Method 2

$$\frac{x}{100\text{ mL}} = \frac{4.00\text{ g}}{3.00\text{ mL}}$$

$$\frac{x}{100\text{ mL}} = 1.33\text{ g/mL}$$

$$x = 100\ \cancel{\text{mL}} \times 1.33\text{ g/}\cancel{\text{mL}}$$

$$= 133\text{ g}$$

The concentration of the drug is 133 g/100 mL of solution, or 133% (m/v).

Check Your Solution

The units are correct. The numerical answer is large, but this is reasonable for an extremely soluble solute.

Sample Problem

Finding Mass for an (m/v) Concentration

Problem

Many people use a solution of trisodium phosphate, Na_3PO_4 (commonly called TSP), to clean walls before putting up wallpaper. The recommended concentration is 1.7% (m/v). What mass of TSP is needed to make 2.0 L of solution?

What Is Required?

You need to find the mass of TSP needed to make 2.0 L of solution.

What Is Given?

The concentration of the solution should be 1.7% (m/v). The volume of solution that is needed is 2.0 L.

Plan Your Strategy

There are two different methods you can use.

Method 1

Use the formula for (m/v) percent. Rearrange the formula to solve for mass. Then substitute in the known values.

Method 2

The percent (m/v) of the solution is numerically equal to the concentration in g/100 mL. Let x represent the mass of TSP dissolved in 2.0 L of solution. The ratio of dissolved solute in 100 mL of solution must be the same as the ratio of the mass of solute, x, dissolved in 2.0 L (2000 mL) of solution.

Continued ...

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Act on Your Strategy

Method 1

$$(\text{m/v})\ \text{percent} = \frac{\text{Mass of solute (in g)}}{\text{Volume of solution (in mL)}} \times 100\%$$

$$\therefore\ \text{Mass of solute} = \frac{(\text{m/v})\ \text{percent} \times \text{Volume of solution}}{100\%}$$

$$= \frac{1.7\% \times 2000\ \text{mL}}{100\%}$$

$$= 34\ \text{g}$$

Method 2

A TSP solution that is 1.7% (m/v) contains 1.7 g of solute dissolved in 100 mL of solution.

$$\frac{1.7\ \text{g}}{100\ \text{mL}} = \frac{x}{2000\ \text{mL}}$$

$$0.017\ \text{g/mL} = \frac{x}{2000\ \text{mL}}$$

$$x = 0.017\ \text{g/mL} \times 2000\ \text{mL}$$

$$= 34\ \text{g}$$

Therefore, 34 g of TSP are needed to make 2.0 L of cleaning solution.

Check Your Solution

The units are appropriate for the problem. The answer appears to be reasonable.

Practice Problems

1. What is the concentration in percent (m/v) of each solution?
- **(a)** 14.2 g of potassium chloride, KCl (used as a salt substitute), dissolved in 450 mL of solution
- **(b)** 31.5 g of calcium nitrate, $Ca(NO_3)_2$ (used to make explosives), dissolved in 1.80 L of solution
- **(c)** 1.72 g of potassium permanganate, $KMnO_4$ (used to bleach stone-washed blue jeans), dissolved in 60 mL of solution

2. A solution of hydrochloric acid was formed by dissolving 1.52 g of hydrogen chloride gas in enough water to make 24.1 mL of solution. What is the concentration in percent (m/v) of the solution?

3. At 25°C, a saturated solution of carbon dioxide gas in water has a concentration of 0.145% (m/v). What mass of carbon dioxide is present in 250 mL of the solution?

4. Ringer's solution contains three dissolved salts in the same proportions as they are found in blood. The salts and their concentrations (m/v) are as follows: 0.86% NaCl, 0.03% KCl, and 0.033% $CaCl_2$. Suppose that a patient needs to receive 350 mL of Ringer's solution by an intravenous drip. What mass of each salt does the pharmacist need to make the solution?

Concentration as a Mass/Mass Percent

The concentration of a solution that contains a solid solute dissolved in a liquid solvent can also be expressed as a mass of solute dissolved in a mass of solution. This is usually expressed as a percent relationship. A **mass/mass percent** gives the mass of a solute divided by the mass of solution, expressed as a percent. The mass/mass percent is also referred to as the *percent (m/m)*, or the *mass percent*. It is often inaccurately referred to as a weight (w/w) percent, as well. Look at your tube of toothpaste, at home. The percent of sodium fluoride in the toothpaste is usually given as a w/w percent. This can be confusing, since weight (*w*) is not the same as mass (*m*). In fact, this concentration should be expressed as a mass/mass percent.

$$\text{Mass/mass percent} = \frac{\text{Mass of solute (in g)}}{\text{Mass of solution (in g)}} \times 100\%$$

For example, 100 g of seawater contains 0.129 g of magnesium ion (along with many other substances). The concentration of Mg^{2+} in seawater is 0.129 (m/m). *Notice that the number of grams of solute per 100 g of solution is numerically equal to the mass/mass percent.*

The concentration of a solid solution, such as an alloy, is usually expressed as a mass/mass percent. Often the concentration of a particular alloy may vary. Table 8.3 gives typical compositions of some common alloys.

Table 8.3 The Composition of Some Common Alloys

Alloy	Uses	Typical percent (m/m) composition
brass	ornaments, musical instruments	Cu (85%) Zn (15%)
bronze	statues, castings	Cu (80%) Zn (10%) Sn (10%)
cupronickel	“silver” coins	Cu (75%) Ni (25%)
dental amalgam	dental fillings	Hg (50%) Ag (35%) Sn (15%)
duralumin	aircraft parts	Al (93%) Cu (5%) other (2%)
pewter	ornaments	Sn (85%) Cu (7%) Bi (6%) Sb ((2%)
stainless steel	cutlery, knives	Fe (78%) Cr (15%) Ni (7%)
sterling silver	jewellery	Ag (92.5%) Cu (7.5%)

Figure 8.16, on the following page, shows two objects made from brass that have distinctly different colours. The difference in colours reflects the varying concentrations of the copper and zinc that make up the objects.

Figure 8.16 Brass can be made using any percent from 50% to 85% copper, and from 15% to 50% zinc. As a result, two objects made of brass can look very different.

Sample Problem

Solving for a Mass/Mass Percent

Problem

Calcium chloride, $CaCl_2$, can be used instead of road salt to melt the ice on roads during the winter. To determine how much calcium chloride had been used on a nearby road, a student took a sample of slush to analyze. The sample had a mass of 23.47 g. When the solution was evaporated, the residue had a mass of 4.58 g. (Assume that no other solutes were present.) What was the mass/mass percent of calcium chloride in the slush? How many grams of calcium chloride were present in 100 g of solution?

What Is Required?

You need to calculate the mass/mass percent of calcium chloride in the solution (slush). Then you need to use your answer to find the mass of calcium chloride in 100 g of solution.

What Is Given?

The mass of the solution is 23.47 g. The mass of calcium chloride that was dissolved in the solution is 4.58 g.

Plan Your Strategy

There are two methods that you can use to solve this problem.

Method 1

Use the formula for mass/mass percent.

$$\text{Mass/mass percent} = \frac{\text{Mass of solute (in g)}}{\text{Mass of solution (in g)}} \times 100\%$$

Continued ...

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The mass of calcium chloride in 100 g of solution will be numerically equal to the mass/mass percent.

Method 2

Use ratios, as in the previous Sample Problems.

Act on Your Strategy

Method 1

$$\text{Mass/mass percent} = \frac{4.58\ \cancel{g}}{23.47\ \cancel{g}} \times 100\%$$
$$= 19.5\%$$

Method 2

$$\frac{x\ g}{100\ g} = \frac{4.58\ g}{23.47\ g}$$
$$\frac{x\ \cancel{g}}{100\ \cancel{g}} = 0.195$$
$$x = 19.5\%$$

The mass/mass percent was 19.5% (m/m). 19.5 g of calcium chloride was dissolved in 100 g of solution.

Check Your Solution

The mass units divide out properly. The final answer has the correct number of significant digits. It appears to be reasonable.

Practice Problems

5. Calculate the mass/mass percent of solute for each solution.
   (a) 17 g of sulfuric acid in 65 g of solution
   (b) 18.37 g of sodium chloride dissolved in 92.2 g of water
   Hint: Remember that a solution consists of both solute and solvent.
   (c) 12.9 g of carbon tetrachloride dissolved in 72.5 g of benzene
6. If 55 g of potassium hydroxide is dissolved in 100 g of water, what is the concentration of the solution expressed as mass/mass percent?
7. Steel is an alloy of iron and about 1.7% carbon. It also contains small amounts of other materials, such as manganese and phosphorus. What mass of carbon is needed to make a 5.0 kg sample of steel?
8. Stainless steel is a variety of steel that resists corrosion. Your cutlery at home may be made of this material. Stainless steel must contain at least 10.5% chromium. What mass of chromium is needed to make a stainless steel fork with a mass of 60.5 g?
9. 18-carat white gold is an alloy. It contains 75% gold, 12.5% silver, and 12.5% copper. A piece of jewellery, made of 18-carat white gold, has a mass of 20 g. How much pure gold does it contain?

Concentration as a Volume/Volume Percent

When mixing two liquids to form a solution, it is easier to measure their volumes than their masses. A **volume/volume percent** gives the volume of solute divided by the volume of solution, expressed as a percent. The volume/volume percent is also referred to as the *volume percent concentration*, *volume percent*, *percent (v/v)*, or the *percent by volume*. You can see this type of concentration on a bottle of rubbing alcohol from a drugstore. (See Figure 8.17.)

$$\text{Volume/volume percent} = \frac{\text{Volume of solute (in mL)}}{\text{Volume of solution (in mL)}} \times 100\%$$

Read through the Sample Problem below, and complete the Practice Problems that follow. You will then have a better understanding of how to calculate the volume/volume percent of a solution.

Figure 8.17 The concentration of this solution of isopropyl alcohol in water is expressed as a volume/volume percent.

Sample Problem

Solving for a Volume/Volume Percent

Problem

Rubbing alcohol is commonly used as an antiseptic for small cuts. It is sold as a 70% (v/v) solution of isopropyl alcohol in water. What volume of isopropyl alcohol is used to make 500 mL of rubbing alcohol?

What Is Required?

You need to calculate the volume of isopropyl alcohol (the solute) used to make 500 mL of solution.

What Is Given?

The volume/volume percent is 70% (v/v). The final volume of the solution is 500 mL.

Plan Your Strategy

Method 1

Rearrange the following formula to solve for the volume of the solute. Then substitute the values that you know into the rearranged formula.

$$\text{Volume/volume percent} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100\%$$

Method 2

Use ratios to solve for the unknown volume.

Continued ...

CHEM FACT

Archaeologists can learn a lot about ancient civilizations by chemically analyzing the concentration of ions in the soil where the people lived. When crops are grown, the crops remove elements such as nitrogen, magnesium, calcium, and phosphorus from the soil. Thus, soil with a lower-than-average concentration of these elements may have held ancient crops. Chlorophyll, which is present in all plants, contains magnesium ions. In areas where ancient crops were processed, the soil has a higher concentration of Mg^{2+}.

Continued ...
FROM PAGE 309

Act on Your Strategy

Method 1

$$\text{Volume/volume percent} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100\%$$

$$\begin{aligned}\text{Volume of solute} &= \frac{\text{Volume/volume percent} \times \text{Volume of solution}}{100\%} \\ &= \frac{70\% \times 500\text{ mL}}{100\%} \\ &= 350\text{ mL}\end{aligned}$$

Method 2

$$\begin{aligned}\frac{x\text{ mL}}{500\text{ mL}} &= \frac{70\text{ mL}}{100\text{ mL}} \\ x &= 0.7 \times 500\text{ mL} \\ &= 350\text{ mL}\end{aligned}$$

Therefore, 350 mL of isopropyl alcohol is used to make 500 mL of 70% (v/v) rubbing alcohol.

Check Your Solution

The answer seems reasonable. It is expressed in appropriate units.

Practice Problems

10. 60 mL of ethanol is diluted with water to a final volume of 400 mL. What is the percent by volume of ethanol in the solution?

11. Milk fat is present in milk. Whole milk usually contains about 5.0% milk fat by volume. If you drink a glass of milk with a volume of 250 mL, what volume of milk fat have you consumed?

12. Both antifreeze (shown in Figure 8.18) and engine coolant contain ethylene glycol. A manufacturer sells a concentrated solution that contains 75% (v/v) ethylene glycol in water. According to the label, a 1:1 mixture of the concentrate with water will provide protection against freezing down to a temperature of –37°C. A motorist adds 1 L of diluted solution to a car radiator. What is the percent (v/v) of ethylene glycol in the diluted solution?

13. The average adult human body contains about 5 L of blood. Of this volume, only about 0.72% consists of leukocytes (white blood cells). These essential blood cells fight infection in the body. What volume of pure leukocyte cells is present in the body of a small child, with only 2.5 L of blood?

14. Vinegar is sold as a 5% (v/v) solution of acetic acid in water. How much water should be added to 15 mL of pure acetic acid (a liquid at room temperature) to make a 5% (v/v) solution of acetic acid? **Note:** Assume that when water and acetic acid are mixed, the total volume of the solution is the sum of the volumes of each.

Figure 8.18 Antifreeze is a solution of ethylene glycol and water.

Concentration in Parts per Million and Parts per Billion

The concentration of a very small quantity of a substance in the human body, or in the environment, can be expressed in **parts per million (ppm)** and **parts per billion (ppb)**. Both parts per million and parts per billion are usually mass/mass relationships. They describe the amount of solute that is present in a solution. Notice that parts per million does not refer to the number of particles, but to the *mass* of the solute compared with the *mass* of the solution.

$$\text{ppm} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$

or

$$\frac{\text{Mass of solute}}{\text{Mass of solution}} = \frac{x\text{ g}}{10^6\text{ g of solution}}$$

$$\text{ppb} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^9$$

or

$$\frac{\text{Mass of solute}}{\text{Mass of solution}} = \frac{x\text{ g}}{10^9\text{ g of solution}}$$

Math LINK

One part per million is equal to 1¢ in $10 000. One part per billion is equal to 1 s in almost 32 years.

What distance (in km) would you travel if 1 cm represented 1 ppm of your journey?

A swimming pool has the dimensions 10 m × 5 m × 2 m. If the pool is full of water, what volume of water (in cm^3) would represent 1 ppb of the water in the pool?

Sample Problem

Parts per Billion in Peanut Butter

Problem

A fungus that grows on peanuts produces a deadly toxin. When ingested in large amounts, this toxin destroys the liver and can cause cancer. Any shipment of peanuts that contains more than 25 ppb of this dangerous fungus is rejected. A company receives 20 t of peanuts to make peanut butter. What is the maximum mass (in g) of fungus that is allowed?

What Is Required?

You need to find the allowed mass (in g) of fungus in 20 t of peanuts.

What Is Given?

The allowable concentration of the fungus is 25 ppb. The mass of the peanut shipment is 20 t.

Plan Your Strategy

Method 1

Convert 20 t to grams. Rearrange the formula below to solve for the allowable mass of the fungus.

$$\text{ppb} = \frac{\text{Mass of fungus}}{\text{Mass of peanuts}} \times 10^9$$

Method 2

Use ratios to solve for the unknown mass.

Continued ...

Continued ...
FROM PAGE 311

Act on Your Strategy

Method 1

First convert the mass in tonnes into grams.

$20\ \text{t} \times 1000\ \text{kg/t} \times 1000\ \text{g/kg} = 20 \times 10^6\ \text{g}$

Next rearrange the formula and find the mass of the fungus.

$$\text{ppb} = \frac{\text{Mass of fungus}}{\text{Mass of peanuts}} \times 10^9$$

$$\therefore \text{Mass of fungus} = \frac{\text{ppb} \times \text{Mass of peanuts}}{10^9}$$

$$= \frac{25\ \text{ppb} \times (20 \times 10^6\ \text{g})}{10^9}$$

$$= 0.5\ \text{g}$$

Method 2

$$\frac{x\ \text{g solute}}{20 \times 10^6\ \text{g solution}} = \frac{25\ \text{g solute}}{1 \times 10^9\ \text{g solution}}$$

$$x\ \text{g} = (20 \times 10^6\ \text{g solution}) \times \frac{25\ \cancel{\text{g}}\ \text{solute}}{1 \times 10^9\ \cancel{\text{g}}\ \text{solution}}$$

$$= 0.5\ \text{g}$$

The maximum mass of fungus that is allowed is 0.5 g.

Check Your Solution

The answer appears to be reasonable. The units divided correctly to give grams. **Note:** Parts per million and parts per billion have no units. The original units, g/g, cancel out.

Practice Problems

15. Symptoms of mercury poisoning become apparent after a person has accumulated more than 20 mg of mercury in the body.

- **(a)** Express this amount as parts per million for a 60 kg person.
- **(b)** Express this amount as parts per billion.
- **(c)** Express this amount as a (m/m) percent.

16. The use of the pesticide DDT has been banned in Canada since 1969 because of its damaging effect on wildlife. In 1967, the concentration of DDT in an average lake trout, taken from Lake Simcoe in Ontario, was 16 ppm. Today it is less than 1 ppm. What mass of DDT would have been present in a 2.5 kg trout with DDT present at 16 ppm?

17. The concentration of chlorine in a swimming pool is generally kept in the range of 1.4 to 4.0 mg/L. The water in a certain pool has 3.0 mg/L of chlorine. Express this value as parts per million. (**Hint:** 1 L of water has a mass of 1000 g.)

18. Water supplies with dissolved calcium carbonate greater than 500 mg/L are considered unacceptable for most domestic purposes. Express this concentration in parts per million.

Careers in Chemistry

Product Development Chemist

A solvent keeps paint liquefied so that it can be applied to a surface easily. After the paint has been exposed to the air, the solvent evaporates and the paint dries. Product development chemists develop and improve products such as paints. To work in product development, they require at least one university chemistry degree.

Chemists who work with paints must examine the properties of many different solvents. They must choose solvents that dissolve paint pigments well, but evaporate quickly and pose a low safety hazard.

Product development chemists must consider human health and environmental impact when choosing between solvents. Many solvents that have been used in the past, such as benzene and carbon tetrachloride, are now known to be harmful to the health and/or the environment. A powerful new solvent called *d-limonene* has been developed from the peel of oranges and lemons. This solvent is less harmful than many older solvents. It has been used successfully as a cleaner for airport runways and automotive parts, and as a pesticide. Chemists are now studying new applications for d-limonene.

Make Career Connections

1. Use reference books or the Internet to find the chemical structure of d-limonene. What else can you discover about d-limonene?
2. To learn more about careers involving work with solvents, contact the Canadian Chemical Producers Association (CCPA).

Molar Concentration

The most useful unit of concentration in chemistry is molar concentration. **Molar concentration** is the number of moles of solute in 1 L of solution. Notice that the volume of the *solution* in *litres* is used, rather than the volume of the *solvent* in *millilitres*. Molar concentration is also known as *molarity*.

$$\text{Molar concentration (in mol/L)} = \frac{\text{Amount of solute (in mol)}}{\text{Volume of solution (in L)}}$$

This formula can be shortened to give

$$C = \frac{n}{V}$$

Molar concentration is particularly useful to chemists because it is related to the number of particles in a solution. None of the other measures of concentration are related to the number of particles. If you are given the molar concentration and the volume of a solution, you can calculate the amount of dissolved solute in moles. This allows you to solve problems involving quantities in chemical reactions, such as the ones on the following pages.

Sample Problem

Calculating Molar Concentration

Problem

A saline solution contains 0.90 g of sodium chloride, NaCl, dissolved in 100 mL of solution. What is the molar concentration of the solution?

What Is Required?

You need to find the molar concentration of the solution in mol/L.

What Is Given?

You know that 0.90 g of sodium chloride is dissolved in 100 mL of solution.

Plan Your Strategy

Step 1 To find the amount (in mol) of sodium chloride, first determine its molar mass. Then divide the amount of sodium chloride (in g) by its molar mass (in g/mol).

Step 2 Convert the volume of solution from mL to L using this formula:

$$\text{Volume (in L)} = \text{Volume (in mL)} \times \frac{1.000\text{ L}}{1000\text{ mL}}$$

Step 3 Use the following formula to calculate the molar concentration:

$$\text{Molar concentration (in mol/L)} = \frac{\text{Amount of solute (in mol)}}{\text{Volume of solution (in L)}}$$

Act on Your Strategy

Step 1 Molar mass of NaCl $= 22.99 + 35.45$

$$= 58.44\text{ g/mol}$$

$$\text{Amount of NaCl} = \frac{0.90\text{ g}}{58.44\text{ g/mol}}$$

$$= 1.54 \times 10^{-2}\text{ mol}$$

Step 2 Convert the volume from mL to L.

$$\text{Volume} = 100\ \cancel{\text{mL}} \times \frac{1.000\text{ L}}{1000\ \cancel{\text{mL}}}$$

$$= 0.100\text{ L}$$

Step 3 Calculate the molar concentration.

$$\text{Molar concentration} = \frac{1.54 \times 10^{-2}\text{ mol}}{0.100\text{ L}}$$

$$= 1.54 \times 10^{-1}\text{ mol/L}$$

The molar concentration of the saline solution is 0.15 mol/L.

Check Your Solution

The answer has the correct units for molar concentration.

Sample Problem

Using Molar Concentration to Find Mass

Problem

At 20˚C, a saturated solution of calcium sulfate, $CaSO_4$, has a concentration of 0.0153 mol/L. A student takes 65 mL of this solution and evaporates it. What mass (in g) is left in the evaporating dish?

What Is Required?

You need to find the mass (in g) of the solute, calcium sulfate.

What Is Given?

The molar concentration is 0.0153 mol/L. The volume of the solution is 65 mL.

Plan Your Strategy

Step 1 Convert the volume from mL to L using the formula

$$\text{Volume (in L)} = \text{Volume (in mL)} \times \frac{1.000\ \text{L}}{1000\ \text{mL}}$$

Step 2 Rearrange the following formula to solve for the amount of solute (in mol).

$$\text{Molar concentration (in mol/L)} = \frac{\text{Amount of solute (in mol)}}{\text{Volume of solution (in L)}}$$

Step 3 Determine the molar mass of calcium sulfate. Use the molar mass to find the mass in grams, using the formula below:

$$\begin{aligned}&\text{Mass (in g) of } CaSO_4\\ &= \text{Amount (in mol)} \times \text{Molar mass of } CaSO_4 \text{ (in g/mol)}\end{aligned}$$

Act on Your Strategy

Step 1 Convert the volume from mL to L.

$$\begin{aligned}\text{Volume} &= 65\ \cancel{\text{mL}} \times \frac{1.000\ \text{L}}{1000\ \cancel{\text{mL}}}\\ &= 0.065\ \text{L}\end{aligned}$$

Step 2 Rearrange the formula to solve for the amount of solute.

$$\text{Molar concentration} = \frac{\text{Amount of solute}}{\text{Volume of solution}}$$

$$\begin{aligned}\therefore \text{Amount of solute} &= \text{Molar concentration} \times \text{Volume of solution}\\ &= 0.0153\ \text{mol/L} \times 0.065\ \text{L}\\ &= 9.94 \times 10^{-4}\ \text{mol}\end{aligned}$$

Step 3 Determine the molar mass. Then find the mass in grams.

$$\begin{aligned}\text{Molar mass of } CaSO_4 &= 40.08 + 32.07 + (4 \times 16.00)\\ &= 136.15\ \text{g/mol}\end{aligned}$$

$$\begin{aligned}\text{Mass (in g) of } CaSO_4 &= 9.94 \times 10^{-4}\ \text{mol} \times 136\ \text{g/mol}\\ &= 0.135\ \text{g}\end{aligned}$$

Continued ...

Continued ...
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Therefore, 0.14 g of calcium sulfate are left in the evaporating dish.

Check Your Solution

The answer has the correct units and the correct number of significant figures.

Practice Problems

19. What is the molar concentration of each solution?

(a) 0.50 mol of NaCl dissolved in 0.30 L of solution

(b) 0.289 mol of iron(III) chloride, $FeCl_3$, dissolved in 120 mL of solution

(c) 0.0877 mol of copper(II) sulfate, $CuSO_4$, dissolved in 70 mL of solution

(d) 4.63 g of sugar, $C_{12}H_{22}O_{11}$, dissolved in 16.8 mL of solution

(e) 1.2 g of $NaNO_3$ dissolved in 80 mL of solution

20. What mass of solute is present in each aqueous solution?

(a) 1.00 L of 0.045 mol/L calcium hydroxide, $Ca(OH)_2$, solution

(b) 500 mL of 0.100 mol/L silver nitrate, $AgNO_3$, solution

(c) 2.5 L of 1.00 mol/L potassium chromate, K_2CrO_4, solution

(d) 40 mL of 6.0 mol/L sulfuric acid, H_2SO_4, solution

(e) 4.24 L of 0.775 mol/L ammonium nitrate, NH_4NO_3, solution

21. A student dissolves 30.46 g of silver nitrate, $AgNO_3$, in water to make 500 mL of solution. What is the molar concentration of the solution?

22. What volume of 0.25 mol/L solution can be made using 14 g of sodium hydroxide, NaOH?

23. A 100 mL bottle of skin lotion contains a number of solutes. One of these solutes is zinc oxide, ZnO. The concentration of zinc oxide in the skin lotion is 0.915 mol/L. What mass of zinc oxide is present in the bottle?

24. Formalin is an aqueous solution of formaldehyde, HCHO, used to preserve biological specimens. What mass of formaldehyde is needed to prepare 1.5 L of formalin with a concentration of 10 mol/L?

You have done many calculations for the concentration of various solutions. Now you are in a position to do some hands-on work with solution concentration. In the following investigation, you will use what you have learned to design your own experiment to determine the concentration of a solution.

SKILL FOCUS

- Initiating and planning
- Performing and recording
- Communicating results
- Predicting

Determining the Concentration of a Solution

Your teacher will give you a sample of a solution. Design and perform an experiment to determine the concentration of the solution. Express the concentration as

(a) mass of solute dissolved in 100 mL of *solution*

(b) mass of solute dissolved in 100 g of *solvent*

(c) amount of solute (in mol) dissolved in 1 L of *solution*

Safety Precautions

When you have designed your investigation, think about the safety precautions you will need to take.

Materials

any apparatus in the laboratory

solution containing a solid dissolved in water

Note: Your teacher will tell you the name of the solute.

Procedure

1. Think about what you need to know in order to determine the concentration of a solution. Then design your experiment so that you can measure each quantity you need. Assume that the density of pure water is 1.00 g/mL.
2. Write the steps that will allow you to measure the quantities you need. Design a data table for your results. Include a space for the name of the solute in your solution.
3. When your teacher approves your procedure, complete your experiment.
4. Dispose of your solution as directed by your teacher.

Analysis

1. Express the concentration of the solution you analyzed as

   (a) mass of solute dissolved in 100 mL of *solution*

   (b) mass of solute dissolved in 100 g of *solvent*

   (c) molar concentration

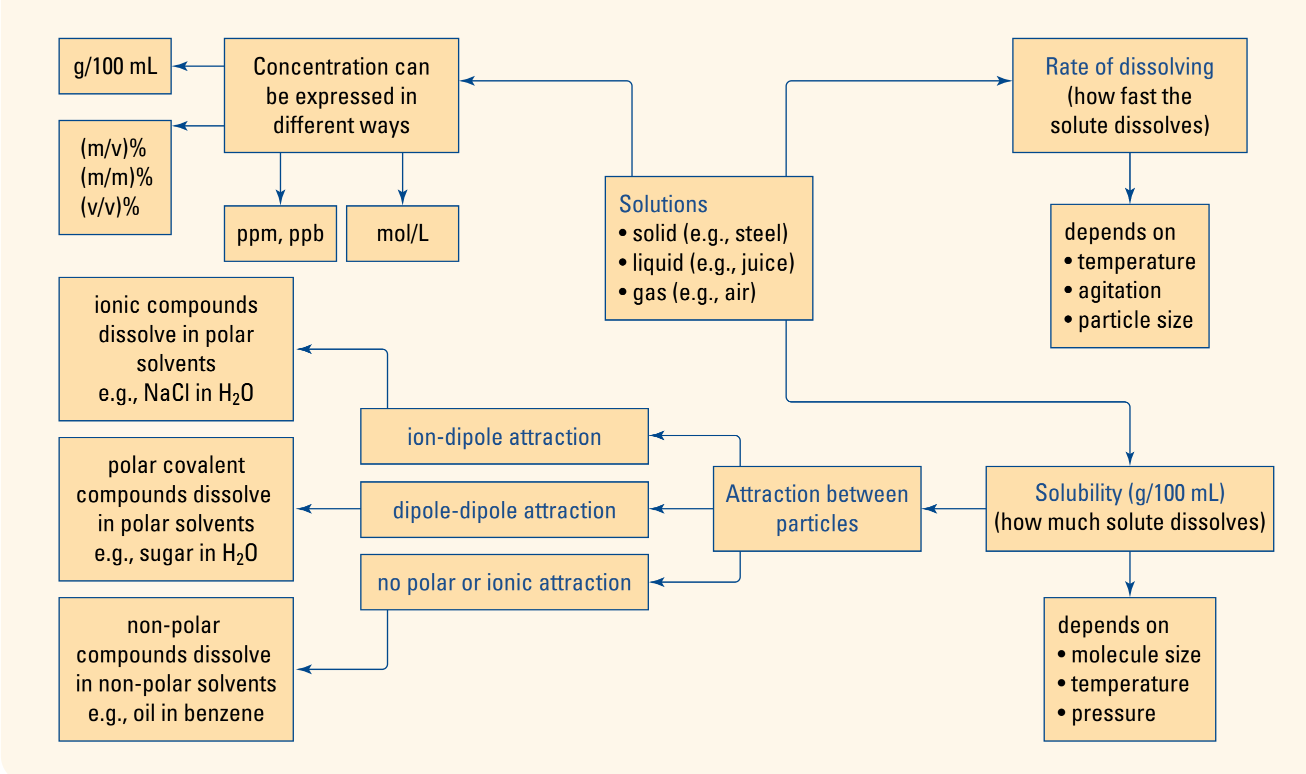
   Show your calculations.

Conclusions

2. List at least two important sources of error in your measurements.
3. List at least two important ways that you could improve your procedure.
4. Did the solute partially decompose on heating, producing a gas and another solid? If so, how do you think this affected the results of your experiment?

Concept Organizer

Summary of Sections 8.1 to 8.3



Section Wrap-up

You have learned about several different ways in which chemists express concentration: mass/volume, mass/mass, and volume/volume percent; parts per million and parts per billion; and molar concentration. The Concept Organizer above summarizes what you have learned in this chapter so far.

In section 8.4, you will learn how standard solutions of known concentration are prepared. You will also learn how to dilute a standard solution.

Section Review

1 I Ammonium chloride, NH_4Cl, is a very soluble salt. 300 g of ammonium chloride are dissolved in 600 mL of water. What is the percent (m/m) of the solution?

2 I A researcher measures 85.1 mL of a solution of liquid hydrocarbons. The researcher then distills the sample to separate the pure liquids. If 20.3 mL of the hydrocarbon hexane are recovered, what is its percent (v/v) in the sample?

3 I A stock solution of phosphoric acid is 85.0% (m/v) H_3PO_4 in water. What is its molar concentration?

4 MC Cytosol is an intracellular solution containing many important solutes. Research this solution. Write a paragraph describing the function of cytosol, and the solutes it contains.

Unit Issue Prep

The concentration of a pesticide or other contaminant dissolved in water determines the level of pollution. If you wish to work on your Unit 3 Issue now, do some research on pesticides such as DDT, diazinon, and MCPA. What concentrations of these pesticides are unacceptable in water systems?

Preparing Solutions

8.4

What do the effectiveness of a medicine, the safety of a chemical reaction, and the cost of an industrial process have in common? They all depend on solutions that are made carefully with known concentrations. A solution with a known concentration is called a **standard solution**. There are two ways to prepare an aqueous solution with a known concentration. You can make a solution by dissolving a measured mass of pure solute in a certain volume of solution. Alternatively, you can dilute a solution of known concentration.

Section Preview/ Specific Expectations

In this section, you will

- **prepare** solutions by dissolving a solid solute and then diluting a concentrated solution
- **communicate** your understanding of the following terms: *standard solution, volumetric flask*

Using a Volumetric Flask

A **volumetric flask** is a pear-shaped glass container with a flat bottom and a long neck. Volumetric flasks like the ones shown in Figure 8.19 are used to make up standard solutions. They are available in a variety of sizes. Each size can measure a fixed volume of solution to ±0.1 mL at a particular temperature, usually 20˚C. When using a volumetric flask, you must first measure the mass of the pure solute. Then you transfer the solute to the flask using a funnel, as shown in Figure 8.20. At this point, you add the solvent (usually water) to dissolve the solute, as in Figure 8.21. You continue adding the solvent until the bottom of the meniscus appears to touch the line that is etched around the neck of the flask. See Figure 8.22. This is the volume of the solution, within ±0.1 mL. If you were performing an experiment in which significant digits and errors were important, you would record the volume of a solution in a 500 mL volumetric flask as 500.0 mL ±0.1 mL. Before using a volumetric flask, you need to rinse it several times with a small quantity of distilled water and discard the washings. *Standard solutions are never stored in volumetric flasks*. Instead, they are transferred to another bottle that has a secure stopper or cap.

Figure 8.19 These volumetric flasks, from left to right, contain solutions of chromium(III) salts, iron(III) salts, and cobalt(II) salts.

Figure 8.20 Transfer a known mass of solid solute into the volumetric flask. Alternatively, dissolve the solid in a small volume of solvent. Then add the liquid to the flask.

Figure 8.21 Add distilled water until the flask is about half full. Swirl the mixture around in order to dissolve the solute completely. Rinse the beaker that contained the solute with solvent. Add the rinsing to the flask.

Figure 8.22 Add the rest of the water slowly. When the flask is almost full, add the water drop by drop until the bottom of the meniscus rests at the etched line.

Diluting a Solution

You can make a less concentrated solution of a known solution by adding a measured amount of additional solvent to the standard solution. The number of molecules, or moles, of solute that is present remains the same before and after the dilution. (See Figure 8.23.)

To reinforce these ideas, read through the Sample Problem below. Then try the Practice Problems that follow.

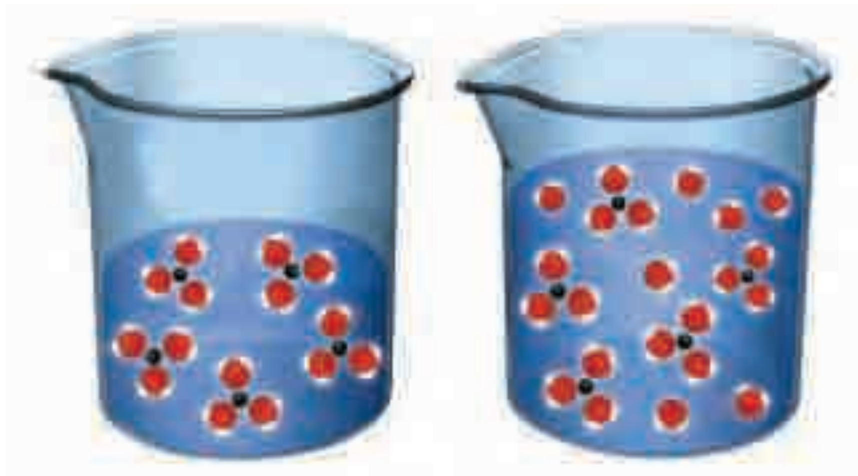


Figure 8.23 When a solution is diluted, the volume increases. However, the amount of solute remains the same.

Sample Problem

Diluting a Standard Solution

Problem

For a class experiment, your teacher must make 2.0 L of 0.10 mol/L sulfuric acid. This acid is usually sold as an 18 mol/L concentrated solution. How much of the concentrated solution should be used to make a new solution with the correct concentration?

What Is Required?

You need to find the volume of concentrated solution to be diluted.

What Is Given?

Initial concentration = 18 mol/L
Concentration of diluted solution = 0.10 mol/L
Volume of diluted solution = 2.0 L

Plan Your Strategy

Note: Amount of solute (mol) after dilution = Amount of solute (mol) before dilution

Step 1 Calculate the amount of solute (in mol) that is needed for the final dilute solution.

Step 2 Calculate the volume of the concentrated solution that will provide the necessary amount of solute.

Continued ...

Continued ...
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Act on Your Strategy

Step 1 Calculate the amount of solute that is needed for the final dilute solution.

$$\text{Molar concentration (in mol/L)} = \frac{\text{Amount of solute (in mol)}}{\text{Volume of solution (in L)}}$$

$$\therefore \text{Amount of solute} = \text{Molar concentration} \times \text{Volume of solution}$$

For the final dilute solution,

$$\begin{aligned}\text{Amount of solute} &= 0.10 \text{ mol/L} \times 2.0 \text{ L} \\ &= 0.20 \text{ mol}\end{aligned}$$

Step 2 Calculate the volume of the original concentrated solution that is needed.

Rearrange and use the molar concentration equation. Substitute in the amount of solute you calculated in step 1.

$$\begin{aligned}\text{Volume of solution (in L)} &= \frac{\text{Amount of solute (in mol)}}{\text{Molar concentration (in mol/L)}} \\ &= \frac{0.20 \cancel{\text{ mol}}}{18 \cancel{\text{ mol}}/\text{L}} \\ &= 0.011 \text{ L}\end{aligned}$$

Therefore, 0.011 L, or 11 mL, of the concentrated 18 mol/L solution should be used to make 2.0 L of 0.10 mol/L sulfuric acid.

Check Your Solution

The units are correct. The final solution must be much less concentrated. Thus, it is reasonable that only a small volume of concentrated solution is needed.

Practice Problems

25. Suppose that you are given a solution of 1.25 mol/L sodium chloride in water, $NaCl_{(aq)}$. What volume must you dilute to prepare the following solutions?

(a) 50 mL of 1.00 mol/L $NaCl_{(aq)}$

(b) 200 mL of 0.800 mol/L $NaCl_{(aq)}$

(c) 250 mL of 0.300 mol/L $NaCl_{(aq)}$

26. What concentration of solution is obtained by diluting 50.0 mL of 0.720 mol/L aqueous sodium nitrate, $NaNO_{3(aq)}$, to each volume?

(a) 120 mL **(b)** 400 mL **(c)** 5.00 L

27. A solution is prepared by adding 600 mL of distilled water to 100 mL of 0.15 mol/L ammonium nitrate. Calculate the molar concentration of the solution. Assume that the volume quantities can be added together.

Now that you understand how to calculate standard solutions and dilution, it is time for you to try it out for yourself. In the following investigation, you will prepare and dilute standard solutions.

SKILL FOCUS

Initiating and planning

Performing and recording

Analyzing and interpreting

Investigation 8-C

Estimating Concentration of an Unknown Solution

Copper(II) sulfate, $CuSO_4$, is a soluble salt. It is sometimes added to pools and ponds to control the growth of fungi. Solutions of this salt are blue in colour. The intensity of the colour increases with increased concentration. In this investigation, you will prepare copper(II) sulfate solutions with known concentrations. Then you will estimate the concentration of an unknown solution by comparing its colour intensity with the colour intensities of the known solutions.

Copper(II) sulfate pentahydrate is a *hydrate*. Hydrates are ionic compounds that have a specific amount of water molecules associated with each ion pair.

Question

How can you estimate the concentration of an unknown solution?

Part 1 Making Solutions with Known Concentrations

Safety Precautions

- Copper(II) sulfate is poisonous. Wash your hands at the end of this investigation.
- If you spill any solution on your skin, wash it off immediately with copious amounts of cool water.

Materials

graduated cylinder
6 beakers
chemical balance
stirring rod
copper(II) sulfate pentahydrate, $CuSO_4{\cdot}5H_2O$
distilled water
labels or grease marker

Procedure

1. With your partner, develop a method to prepare 100 mL of 0.500 mol/L aqueous $CuSO_4{\cdot}5H_2O$ solution. Include the water molecules that are hydrated to the crystals, as given in the molecular formula, in your calculation of the molar mass. Show all your calculations. Prepare the solution.
2. Save some of the solution you prepared in step 1, to be tested in Part 2. Use the rest of the solution to make the dilutions in steps 3 to 5. Remember to label the solutions.
3. Develop a method to dilute part of the 0.500 mol/L $CuSO_4$ solution, to make 100 mL of 0.200 mol/L solution. Show your calculations. Prepare the solution.
4. Show your calculations to prepare 100 mL of 0.100 mol/L solution, using the solution you prepared in step 3. (You do not need to describe the method because it will be similar to the method you developed in step 3. Only the volume diluted will be different.) Prepare the solution.
5. Repeat step 4 to make 100 mL of 0.050 mol/L $CuSO_4$, by diluting part of the 0.100 mol/L solution you made. Then make 50 mL of 0.025 mol/L solution by diluting part of the 0.050 mol/L solution.

Part 2 Estimating the Concentration of an Unknown Solution

Materials

paper towels
6 clean, dry, identical test tubes
medicine droppers
5 prepared solutions from Part 1
10 mL of copper(II) sulfate, $CuSO_4$, solution with an unknown concentration

Procedure

1. You should have five labelled beakers containing $CuSO_4$ solutions with the following concentrations: 0.50 mol/L, 0.20 mol/L, 0.10 mol/L, 0.05 mol/L, and 0.025 mol/L. Your teacher will give you a sixth solution of unknown concentration. Record the letter or number that identifies this solution.
2. Label each test tube, one for each solution. Pour a sample of each solution into a test tube. The height of the solution in the test tubes should be the same. Use a medicine dropper to add or take away solution as needed. (Be careful not to add water, or a solution of different concentration, to a test tube.)
3. The best way to compare colour intensity is by looking down through the test tube. Wrap each test tube with a paper towel to stop light from entering the side. Arrange the solutions of known concentration in order.
4. Place the solutions over a diffuse light source such as a lightbox. Compare the colour of the unknown solution with the colours of the other solutions.
5. Use your observations to estimate the concentration of the unknown solution.

6. Pour the solutions of $CuSO_4$ into a beaker supplied by your teacher. Wash your hands.

Analysis

1. Describe any possible sources of error for Part 1 of this investigation.
2. What is your estimate of the concentration of the unknown solution?

Conclusion

3. Obtain the concentration of the unknown solution from your teacher. Calculate the percentage error in your estimate.

Applications

4. Use your estimated concentration of the unknown solution to calculate the mass of $CuSO_4 \cdot 5H_2O$ that your teacher would need to prepare 500 mL of this solution.
5. If your school has a spectrometer or colorimeter, you can measure the absorption of light passing through the solutions. By measuring the absorption of solutions of copper(II) sulfate with different concentrations, you can draw a graph of absorption against concentration.

Diluting Concentrated Acids

Figure 8.24 When diluting acid, always add the acid to the water—never the reverse. Rubber gloves, a lab coat, and safety goggles or a face shield protect against acid splashes.

The acids that you use in your investigations are bought as concentrated standard solutions. Sulfuric acid is usually bought as an 18 mol/L solution. Hydrochloric acid is usually bought as a 12 mol/L solution. These acids are far too dangerous for you to use at these concentrations. Your teacher dilutes concentrated acids, following a procedure that minimizes the hazards involved.

Concentrated acids should be diluted in a fume hood because breathing in the fumes causes acid to form in air passages and lungs. Rubber gloves must be used to protect the hands. A lab coat is needed to protect clothing. Even small splashes of a concentrated acid will form holes in fabric. Safety goggles, or even a full-face shield, are essential.

Mixing a strong, concentrated acid with water is a very exothermic process. A concentrated acid is denser than water. Therefore, when it is poured into water, it sinks into the solution and mixes with the solution. The heat that is generated is spread throughout the solution. This is the only safe way to mix an acid and water. If you added water to a concentrated acid, the water would float on top of the solution. The heat generated at the acid-water layer could easily boil the solution and splatter highly corrosive liquid. The sudden heat generated at the acid-water boundary could crack the glassware and lead to a very dangerous spill. Figure 8.24 illustrates safety precautions needed to dilute a strong acid.

Section Wrap-up

In this section, you learned how to prepare solutions by dissolving a solid solute and then diluting a concentrated solution. In the next chapter, you will see how water is used as a solvent in chemistry laboratories. Many important reactions take place in water. You will also learn more about water pollution and water purification.

Section Review

1. I What mass of potassium chloride, KCl, is used to make 25.0 mL of a solution with a concentration of 2.00 mol/L?

2. I A solution is prepared by dissolving 42.5 g of silver nitrate, $AgNO_3$ in a 1 L volumetric flask. What is the molar concentration of the solution?

3. I The solution of aqueous ammonia that is supplied to schools has a concentration of 14 mol/L. Your class needs 3.0 L of a solution with a concentration of 0.10 mol/L.
   - **(a)** What procedure should your teacher follow to make up this solution?
   - **(b)** Prepare an instruction sheet or a help file for your teacher to carry out this dilution.

4. I 47.9 g of potassium chlorate, $KClO_3$, is used to make a solution with a concentration of 0.650 mol/L. What is the volume of the solution?

5. I Water and 8.00 mol/L potassium nitrate solution are mixed to produce 700 mL of a solution with a concentration of 6.00 mol/L. What volumes of water and potassium nitrate solution are used?

CHAPTER 8 Review

Reflecting on Chapter 8

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Describe the difference between a saturated and an unsaturated solution.
- Explain how you can predict whether a solute will dissolve in a solvent.
- What factors affect the rate of dissolving?
- What factors affect solubility?
- How does temperature affect the solubility of a solid, a liquid, and a gas?
- Describe how particle attractions affect solubility.
- Explain how to plot a solubility curve.
- Write the formulas for (m/v) percent, (m/m) percent, (v/v) percent, ppm, ppb, and molar concentration.
- Explain how you would prepare a standard solution using a volumetric flask.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

solution	hydrogen bonding
solvent	ion-dipole attractions
solutes	hydrated
variable composition	electrolyte
aqueous solution	non-electrolytes
miscible	concentration
immiscible	mass/volume percent
alloys	mass/mass percent
solubility	volume/volume percent
saturated solution	parts per million
unsaturated solution	parts per billion
rate of dissolving	molar concentration
dipole	standard solution
dipole-dipole attraction	volumetric flask

Knowledge/Understanding

1. Identify at least two solutions in your home that are
   (a) beverages
   (b) found in the bathroom or medicine cabinet
   (c) solids
2. How is a solution different from a pure compound? Give specific examples.
3. Mixing 2 mL of linseed oil and 4 mL of turpentine makes a binder for oil paint. What term is used to describe liquids that dissolve in each other? Which liquid is the solvent?
4. How does the bonding in water molecules account for the fact that water is an exellent solvent?
5. Why does an aqueous solution of an electrolyte conduct electricity, but an aqueous solution of a non-electrolyte does not?
6. Use the concept of forces between particles to explain why oil and water are immiscible.
7. Explain the expression "like dissolves like" in terms of intermolecular forces.
8. What factors affect the rate of dissolving of a solid in a liquid?
9. Which of the following substances would you expect to be soluble in water? Briefly explain each answer.
   (a) potassium chloride, KCl
   (b) carbon tetrachloride, CCl_4
   (c) sodium sulfate, Na_2SO_4
   (d) butane, C_4H_{10}
10. Benzene, C_6H_6, is a liquid at room temperature. It is sometimes used as a solvent. Which of the following compounds is more soluble in benzene: naphthalene, $C_{10}H_8$, or sodium fluoride, NaF? Would you expect ethanol, CH_3CH_2OH, to be soluble in benzene? Explain your answers.

Inquiry

11. Boric acid solution is used as an eyewash. What mass of boric acid is present in 250 g of solution that is 2.25% (m/m) acid in water?
12. 10% (m/m) sodium hydroxide solution, $NaOH_{(aq)}$, is used to break down wood fibre to make paper.
    (a) What mass of solute is needed to make 250 mL of 10% (m/m) solution?
    (b) What mass of solvent is needed?
    (c) What is the molar concentration of the solution?
13. What volume of pure ethanol is needed to make 800 mL of a solution of ethanol in water that is 12% (v/v)?

14. Some municipalities add sodium fluoride, NaF, to drinking water to help protect the teeth of children. The concentration of sodium fluoride is maintained at 2.9×10^{-5} mol/L. What mass (in mg) of sodium fluoride is dissolved in 1 L of water? Express this concentration in ppm.

15. A saturated solution of sodium acetate, $NaCH_3COO$, can be prepared by dissolving 4.65 g in 10.0 mL of water at 20°C. What is the molar concentration of the solution?

16. What is the molar concentration of each of the following solutions?
  (a) 7.25 g of silver nitrate, $AgNO_3$, dissolved in 100 mL of solution
  (b) 80 g of glucose, $C_6H_{12}O_6$, dissolved in 70 mL of solution

17. Calculate the mass of solute that is needed to prepare each solution below.
  (a) 250 mL of 0.250 mol/L calcium acetate, $Ca(CH_3COO)_2$
  (b) 1.8 L of 0.35 mol/L ammonium sulfate, $(NH_4)_2SO_4$

18. Calculate the molar concentration of each solution formed after dilution.
  (a) 20 mL of 6.0 mol/L hydrochloric acid, $HCl_{(aq)}$, diluted to 70 mL
  (b) 300 mL of 12.0 mol/L ammonia, $NH_{3(aq)}$, diluted to 2.50 L

19. Calculate the molar concentration of each solution. Assume that the volumes can be added.
  (a) 85.0 mL of 1.50 mol/L ammonium chloride, $NH_4Cl_{(aq)}$, added to 250 mL of water
  (b) a 1:3 dilution of 1.0 mol/L calcium phosphate (that is, one part stock solution mixed with three parts water)

20. A standard solution of 0.250 mol/L calcium ion is prepared by dissolving solid calcium carbonate in an acid. What mass of calcium carbonate is needed to prepare 1.00 L of the solution?

21. Suppose that your teacher gives you three test tubes. Each test tube contains a clear, colourless liquid. One liquid is an aqueous solution of an electrolyte. Another liquid is an aqueous solution of a non-electrolyte. The third liquid is distilled water. Outline the procedure for an experiment to identify which liquid is which.

22. Fertilizers for home gardeners may be sold as aqueous solutions. Suppose that you want to begin a company that sells an aqueous solution of potassium nitrate, KNO_3, fertilizer. You need a solubility curve (a graph of solubility versus temperature) to help you decide what concentration to use for your solution. Describe an experiment that you might perform to develop a solubility curve for potassium nitrate. State which variables are controlled, which are varied, and which must be measured.

23. Potassium alum, $KAl(SO_4)_2 \cdot 12H_2O$, is used to stop bleeding from small cuts. The solubility of potassium alum, at various temperatures, is given in the following table.

Solubility of Potassium Alum

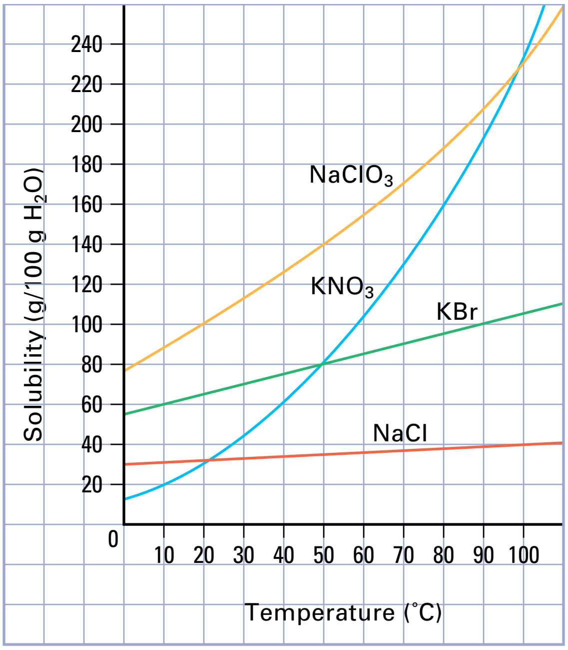
Solubility (g/100 g water)	Temperature (°C)
4	0
10	10
15	20
23	30
31	40
49	50
67	60
101	70
135	80

  (a) Plot a graph of solubility against temperature.
  (b) From your graph, interpolate the solubility of potassium alum at 67°C.
  (c) By extrapolation, estimate the solubility of potassium alum at 82°C.
  (d) Look at your graph. At what temperature will 120 g of potassium alum form a saturated solution in 100 g of water?

24. Use the graph on the next page to answer questions 24 and 25. At 80°C, what mass of sodium chloride dissolves in 1.0 L of water?

25. What minimum temperature is required to dissolve 24 g of potassium nitrate in 40 g of water?

26. A teacher wants to dilute 200 mL of 12 mol/L hydrochloric acid to make a 1 mol/L solution. What safety precautions should the teacher take?



This graph shows the solubility of four salts at various temperatures. Use it to answer questions 24 and 25.

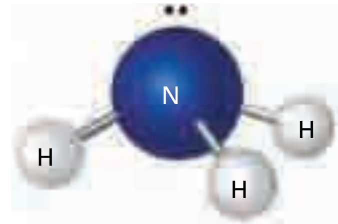
Communication

27. Suppose that you make a pot of hot tea. Later, you put a glass of the tea in the refrigerator to save it for a cool drink. When you take it out of the refrigerator some hours later, you notice that it is cloudy. How could you explain this to a younger brother or sister?

28. Define each concentration term.
- **(a)** percent (m/v)
- **(b)** percent (m/m)
- **(c)** percent (v/v)
- **(d)** parts per million, ppm
- **(e)** parts per billion, ppb

29. The concentration of iron in the water that is supplied to a town is 0.25 mg/L. Express this in ppm and ppb.

30. Ammonia is a gas at room temperature and pressure, but it can be liquefied easily. Liquid ammonia is probably present on some planets. Scientists speculate that it might be a good solvent. Explain why, based on the structure of the ammonia molecule shown above.



31. At 20°C, the solubility of oxygen in water is more than twice that of nitrogen. A student analyzed the concentration of dissolved gases in an unpolluted pond. She found that the concentration of nitrogen gas was greater than the concentration of oxygen. Prepare an explanation for the student to give to her class.

32. What is the concentration of pure water?

Making Connections

33. A bright red mineral called cinnabar has the chemical formula HgS. It can be used to make an artist's pigment, but it is a very insoluble compound. A saturated solution at 25°C has a concentration of 2×10^{-27} mol/L. In the past, why was heavy metal poisoning common in painters? Why did painters invariably waste more cinnabar than they used?

34. Vitamin A is a compound that is soluble in fats but not in water. It is found in certain foods, including yellow fruit and green vegetables. In parts of central Africa, children frequently show signs of vitamin A deficiency, although their diet contains a good supply of the necessary fruits and vegetables. Why?

Answers to Practice Problems and Short Answers to Section Review Questions:

Practice Problems: **1.(a)** 3.16% **(b)** 1.75% **(c)** 2.9% **2.** 6.31% **3.** 0.362 g **4.** 3.0 g, 0.1 g, 0.12 g **5.(a)** 26% **(b)** 16.6% **(c)** 15.1% **6.** 35% **7.** 85 g **8.** 6.35 g **9.** 15 g **10.** 15% **11.** 12 mL **12.** 38% **13.** 18 mL **14.** 285 mL **15.(a)** 0.33 ppm **(b)** 3.3×10^2 ppb **(c)** 0.000033% **16.** 0.040 g **17.** 3.0 ppm **18.** 500 ppm **19.(a)** 1.7 mol/L **(b)** 2.41 mol/L **(c)** 1.2 mol/L **(d)** 0.805 mol/L **(e)** 0.18 mol/L **20.(a)** 3.3 g **(b)** 8.49 g **(c)** 4.9×10^2 g **(d)** 24 g **(e)** 263 g **21.** 0.359 mol/L **22.** 1.4 L **23.** 7.45 g **24.** 4.5×10^2 g **25.(a)** 40 mL **(b)** 128 mL **(c)** 60.0 mL **26.(a)** 0.300 mol/L **(b)** 9.00×10^{-2} mol/L **(c)** 7.20×10^{-3} mol/L **27.** 2.1×10^{-2} mol/L

Section Review: **8.1:** **1.** solute, solvent **8.** polar, ionic **13.** non-polar **8.2:** **3.(a)** increases **(b)** decreases **6.(a)** $Ce_2(SO_4)_3$ **(b)** $Ce_2(SO_4)_3$, $NaNO_3$ **(c)** NaCl **(d)** 84°C **(e)** 10 g **8.3:** **1.** 33.3 **2.** 23.8% **3.** 8.67 mol/L **8.4:** **1.** 3.73 g **2.** 0.25 mol/L **4.** 601 mL **5.** 175 mL, 525 mL


500 mL
±5%
400
300