CHAPTER 13

SOLUTIONS

Ocean water is an excellent example of a solution. A solution is a homogeneous mixture of two or more substances. *Homogeneous* means that the solution looks the same throughout, even under a microscope. In this case, water is the solvent—the substance in excess. Many substances are dissolved in the water. The most abundant substance is ordinary table salt, NaCl, which is present as the ions Na^+ and Cl^-. Other ions present include Mg^{2+}, Ca^{2+}, and Br^-. Oxygen gas is dissolved in the water. Without oxygen, the fish could not live. Because humans do not have gills like fish do, the diver needs scuba gear.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Exploring Types of Mixtures

PROCEDURE

1. Prepare five mixtures in **five different 250 mL beakers.** Each should contain about 200 mL of water and one of the following substances: **12 g of sucrose, 3 g of soluble starch, 5 g of clay, 2 mL of food coloring,** and **20 mL of cooking oil.**
2. Observe the five mixtures and their characteristics. Record the appearance of each mixture after stirring.
3. Note which mixtures do not separate after standing. Transfer 10 mL of each of these mixtures to an individual **test tube.** Shine a **flashlight** through each mixture in a darkened room. Make a note of the mixture(s) in which the path of the light beam is visible.

ANALYSIS

1. What characteristic did the mixture(s) that separated after stirring share?
2. How do you think the mixture(s) in which the light's path was visible differed from those in which it was not?

Pre-Reading Questions

1. **Give three examples of solutions you find in everyday life.**
2. **What main components do these solutions consist of?**
3. **How do you know that each of these examples is actually a solution?**

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SECTION 1

What Is a Solution?

KEY TERMS

- **solution**
- **suspension**
- **solvent**
- **solute**
- **colloid**

OBJECTIVES

1. **Distinguish** between solutions, suspensions, and colloids.
2. **Describe** some techniques chemists use to separate mixtures.

Mixtures

You have learned about the difference between pure substances and mixtures. Mixtures can either be *heterogeneous* or *homogeneous*. The particles of a heterogeneous mixture are large enough to see under a microscope. In a homogeneous mixture, however, the particles are molecule-sized, so the mixture appears uniform, even under a microscope. A homogenous mixture is also known as a **solution.**

solution

a homogeneous mixture of two or more substances uniformly dispersed throughout a single phase

Suspensions Are Temporary Heterogeneous Mixtures

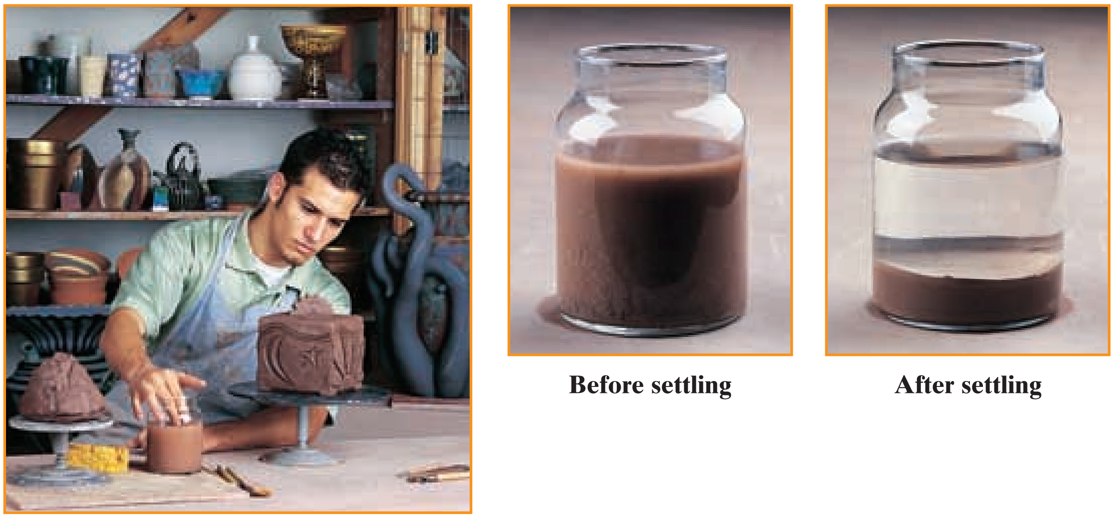
The potter shown in **Figure 1** uses water to help sculpt the sides of a clay pot. As he dips his clay-covered fingers into a container of water, the water turns brown. The clay-water mixture appears uniform. However, if the container sits overnight, the potter will see a layer of mud on the bottom and clear water on top. The clay does not dissolve in water. The clay breaks up into small pieces that are of such low mass that, for a while, they remain suspended in the water. This type of mixture, in which the different parts spontaneously separate over time, is called a **suspension.** In a suspension, the particles may remain mixed with the liquid while the liquid is being stirred, but later they settle to the bottom.

suspension

a mixture in which particles of a material are more or less evenly dispersed throughout a liquid or gas

Topic Link

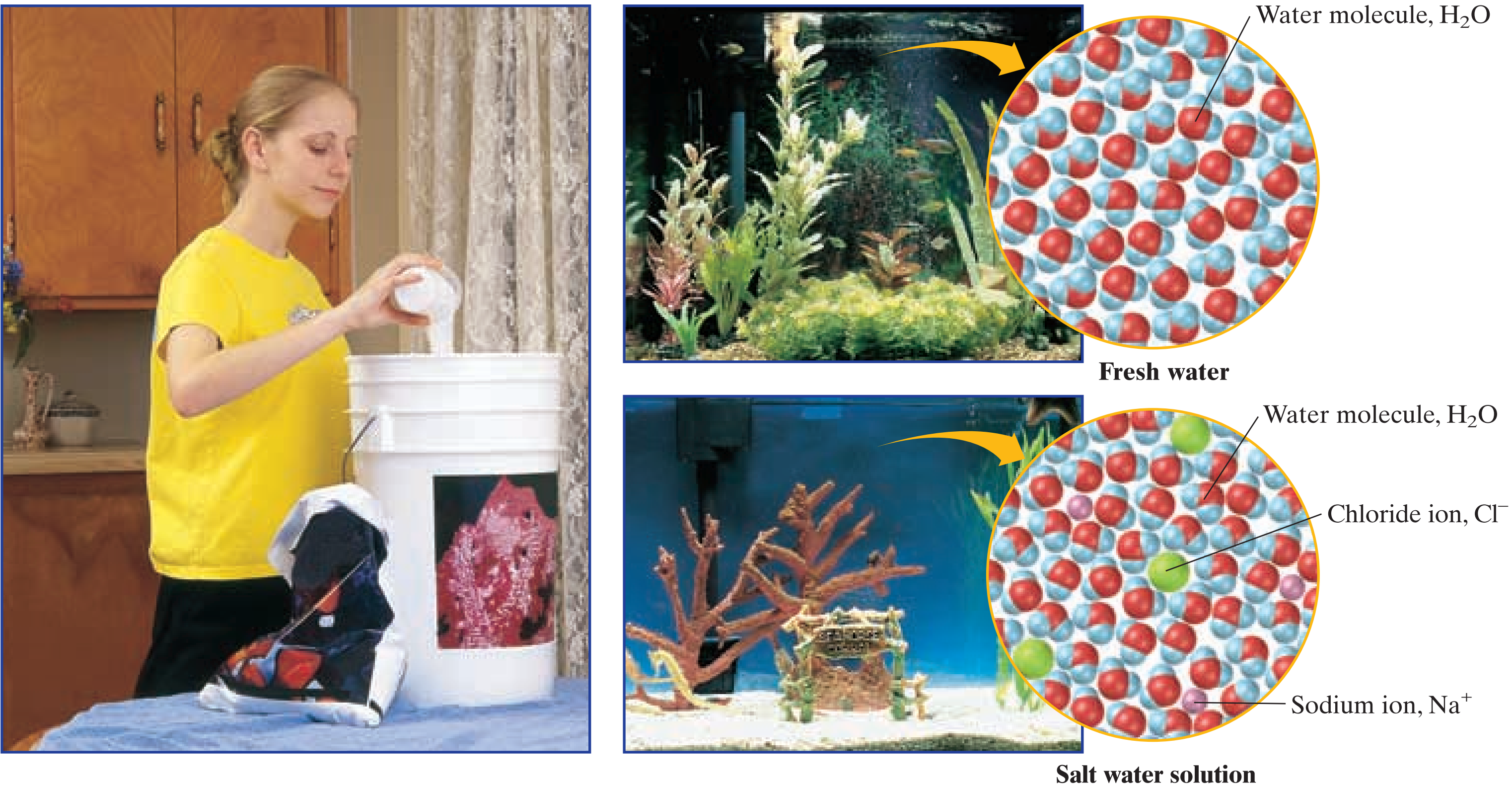
Refer to the "The Science of Chemistry" chapter for a discussion of the classification of mixtures.



Before settling

After settling

Figure 1 Clay particles suspended in water form a suspension.



Solutions Are Stable Homogeneous Mixtures

A student working in a pet shop is asked to prepare some water for a salt water aquarium. She prepares a bucket of fresh water. Then she adds a carefully measured quantity of salt crystals to the water, as shown in **Figure 2,** and stirs the water. These crystals consist of a mixture of salts which, when dissolved in water, will produce a solution with the same composition as sea water. After stirring, the student can no longer see any grains of salt in the water. No matter how long she waits, the salt will not separate from the water. The salt has dissolved in the water to form a stable homogeneous mixture. The particles are evenly distributed throughout the mixture, making it a true solution. The dissolved particles, which are ions in this case, are close to the size of the water molecules and are not clustered together.

Figure 2
Fresh water is stable and homogeneous. The saltwater mixture is also stable and homogeneous because mixing occurs between molecules and ions.

Solution Is a Broad Term

Any mixture that is homogeneous on a microscopic level is a solution. According to that definition, air is a gaseous solution. However, when most people use the word *solution,* they are usually referring to a homogeneous *liquid* mixture. A homogeneous liquid mixture has one main component—a liquid—as well as one or more additional ingredients that are usually present in smaller amounts. The primary ingredient in a solution is called the **solvent,** and the other ingredients are the **solutes** and are said to be dissolved in the solvent. Water is the most common solvent. Although it is a very common substance, water is a unique solvent because so many substances can dissolve in it. Solutions in which water is the solvent are called *aqueous* solutions.

solvent
in a solution, the substance in which the solute is dissolved

solute
in a solution, the substance that is dissolved in the solvent

Figure 3
Brass is a solid solution of zinc atoms in copper atoms.

Another type of solution involves one solid mixed with another solid. Examples include solid alloys, such as brass, bronze, and steel. Brass, shown in **Figure 3,** is a mixture of copper and zinc. Brass is widely used in musical instruments because it is harder and more resistant to corrosion than pure copper.

Colloids Are Stable Heterogeneous Mixtures

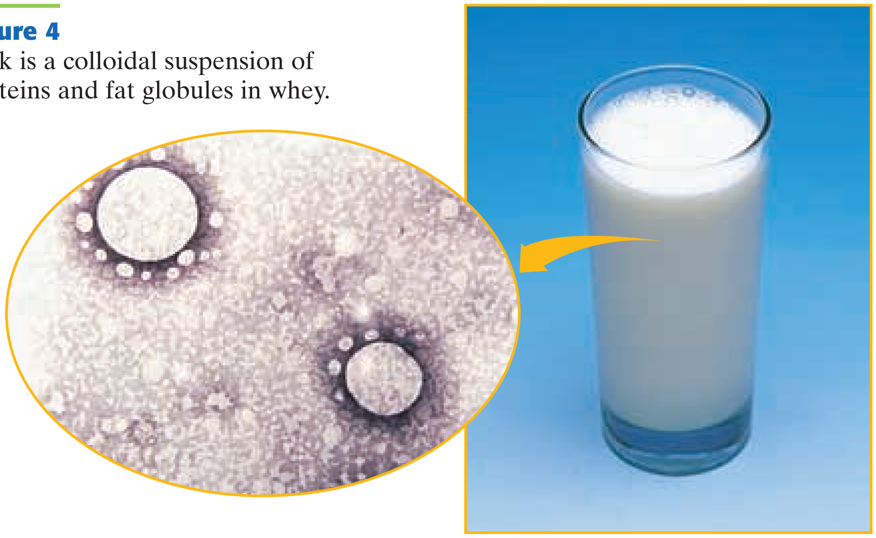
Milk appears to be homogeneous. But as **Figure 4** shows, under a microscope you see that milk contains globules of fat and small lumps of the protein casein dispersed in a liquid called *whey*. Milk is actually a **colloid,** and not a solution. The particles of casein do not settle out after standing.

The particles in a colloid usually have an electric charge. These like-charged particles repel each other, so they do not collect into larger particles that would settle out. If you added acid to milk, the acid would neutralize the charge, and the particles would coagulate and settle to the bottom of the container.

colloid

a mixture consisting of tiny particles that are intermediate in size between those in solutions and those in suspensions and that are suspended in a liquid, solid, or gas

Figure 4
Milk is a colloidal suspension of proteins and fat globules in whey.



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Separating Mixtures

There are many ways to separate mixtures into their components. The best method to use in a particular case depends on the kind of mixture and on the properties of the ingredients. The methods shown in **Figure 5** rely on the physical properties of the ingredients to separate them. Notice that some of these methods can be used outside a laboratory. In fact, you may use one or more of these methods in your own home. For example, you may use filtration to make coffee or evaporation when you cook. Centrifuges are used in dairies to separate the cream from the milk to make fat-free milk.

Figure 5

a Decanting separates a liquid from solids that have settled. To decant a mixture, carefully pour off the liquid to leave the solids behind.

b A centrifuge is used to separate substances of different densities. The centrifuge spins rapidly, and the denser substances collect at the bottom of the tube.

c Ground coffee is separated from liquid coffee by filtration. The filtrate—the liquid and whatever passes through the filter—collects in the coffeepot. The solid grounds stay on the filter.

d In saltwater ponds such as this one, sea water evaporates, and salts, mainly sodium chloride, are left behind.

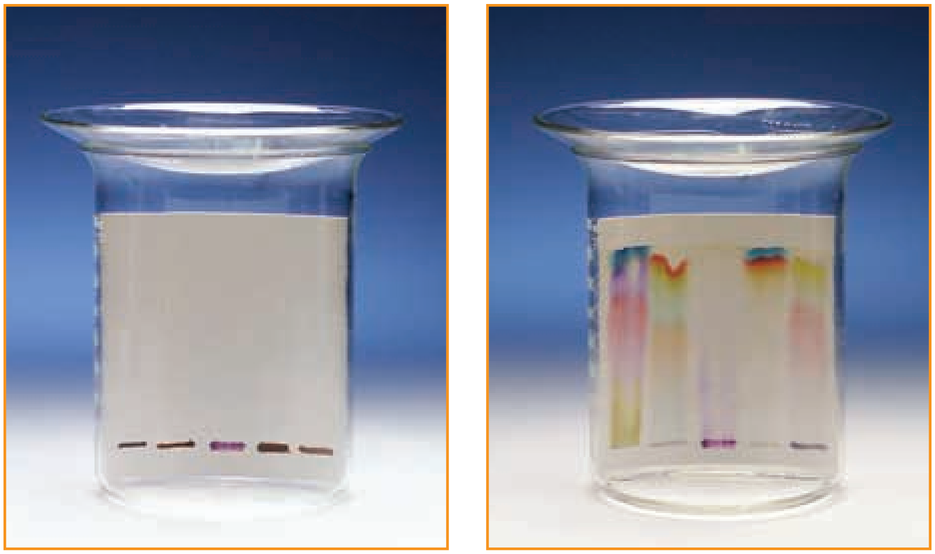


Figure 6
By using paper chromatography, you can separate different dyes mixed together in a product.

Chromatography Separates by Differences in Attraction

You may have noticed that stains stick to some fabrics more than other fabrics. Also, different processes are used to remove different stains. This illustrates the principles used in *chromatography.* Chromatography separates components of a mixture based on how quickly molecules dissolved in a *mobile phase* solvent move along a *solid phase.*

Paper chromatography is a powerful technique for separating solutions. For example, it can be used to separate the dyes in ink. The dyes are blotted onto the paper (solid phase), and a solvent such as water (mobile phase) travels up the paper. The solvent dissolves the ink as it travels up the paper. Dyes that are attracted more strongly to the paper than other dyes travel more slowly along the paper. The right-hand photo in **Figure 6** shows the separation of dyes that make up different colors of ink.

Quick LAB

SAFETY PRECAUTIONS

Design Your Own Lab: The Colors of Candies

In this activity you will investigate whether the colors of candy-coated chocolates are single dyes or mixtures of dyes.

PROCEDURE

Design an experiment that uses **candy-coated chocolates, chromatography paper, small beakers,** and **0.1% NaCl developing solution** to determine whether the dyes in the candies are mixtures of dyes or are single dyes. For example, is the green color a result of mixing two rimary colors? Are the other colors mixtures?

ANALYSIS

1. Prepare a report that includes your experimental procedure, a data table that summarizes your results, and the experimental evidence.
2. Be sure to answer the questions posed in the procedure.

Distillation Separates by Differences in Boiling Point

Sometimes mixtures of liquids need to be purified or have their components separated. If the boiling points of the components are different, *distillation* can separate them based on their boiling points. As one component reaches its boiling point, it evaporates from the mixture and is allowed to cool and condense. This process continues until all the desired components have been separated from the mixture. For example, fermentation produces a solution of alcohol in water. If this is placed in a pot and boiled, the alcohol boils first. This alcohol-rich *distillate* can be collected by a distilling column. The distilling column is a cooler surface upon which the distillate recondenses, and can be collected as a liquid.

In some places where fresh water is scarce, distillation is used to obtain drinking water from sea water. However, because distillation requires a lot of energy, the process is expensive. Distillation is also used in the petroleum industry to separate crude oil into fractions according to their boiling points. The first fractions to distill are fluids with low boiling points, used as raw material in the plastics industry. Next comes gasoline, then at higher temperatures diesel fuel, then heating oil, then kerosene distill. What remains is the basis for lubricating greases.

1 Section Review

UNDERSTANDING KEY IDEAS

1. Explain why a suspension is considered a heterogeneous mixture.

2. Classify the following mixtures as homogeneous or heterogeneous:

a. lemon juice **c.** blood

b. tap water **d.** house paint

3. In a solution, which component is considered the solvent? Which is the solute?

4. Name the solvent and solute(s) in the following solutions:

a. carbonated water **c.** coffee

b. apple juice **d.** salt water

5. Does a solution have to involve a liquid? Explain your answer.

6. How is a colloid distinguished from a solution or a suspension?

7. What is the basic physical principle that chromatography is based upon?

8. How can distillation be used to prepare pure water from tap water?

CRITICAL THINKING

9. Explain how you could determine that brass is a solution, and not a colloid or suspension.

10. Explain why fog is a colloid.

11. You get a stain on a table cloth. Soapy water will not take the stain out, but rubbing alcohol will. How does this relate to chromatography?

12. If you allow a container of sea water to sit in the sun, the liquid level gets lower and lower, and finally crystals appear. What is happening?

SECTION 2

Concentration and Molarity

KEY TERMS

- **concentration**
- **molarity**

OBJECTIVES

1. **Calculate** concentration using common units.
2. **Define** molarity, and calculate the molarity of a solution.
3. **Describe** the procedure for preparing a solution of a certain molarity.
4. **Use** molarity in stoichiometric calculations.

Concentration

In a solution, the solute is distributed evenly throughout the solvent. This means that any part of a solution has the same ratio of solute to solvent as any other part of the solution. This ratio is the **concentration** of the solution. Some common ways of expressing concentration are given in **Table 1.**

concentration

the amount of a particular substance in a given quantity of a solution

Calculating Concentration

Concentrations can be expressed in many forms. One unit of concentration used in pollution measurements that involve very low concentrations is *parts per million,* or ppm. Parts per million is the number of grams of solute in 1 million grams of solution. For example, the concentration of lead in drinking water may be given in parts per million.

When you want to express concentration, you will begin with analytical data which may be expressed in units other than the units you want to use. In that case, each value must be converted into the appropriate units. Then, you must be sure to express the concentration in the correct ratio. **Sample Problem A** shows a typical calculation.

Table 1 Some Measures of Concentration

Name	Abbreviation or symbol	Units	Areas of application
Molarity	M	$\frac{\text{mol solute}}{\text{L solution}}$	in solution stoichiometry calculations
Molality	*m*	$\frac{\text{mol solute}}{\text{kg solvent}}$	with calculation of properties such as boiling-point elevation and freezing-point depression
Parts per million	ppm	$\frac{\text{g solute}}{\text{1000 000 g solution}}$	to express small concentrations

See Appendix A for additional units of concentration.

SAMPLE PROBLEM A

Calculating Parts per Million

A chemical analysis shows that there are 2.2 mg of lead in exactly 500 g of water. Convert this measurement to parts per million.

1 Gather information.

mass of solute: 2.2 mg
mass of solvent: 500 g
parts per million = ?

2 Plan your work.

First, change 2.2 mg to grams:

$$2.2 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 2.2 \times 10^{-3} \text{ g}.$$

Divide this by 500 g to get the amount of lead in 1 g water, then multiply by 1000 000 to get the amount of lead in 1000 000 g water.

3 Calculate.

$$\frac{0.0022 \cancel{\text{ g}} \text{ Pb}}{500 \cancel{\text{ g}} \text{ H}_2\text{O}} \times \frac{1000\,000 \text{ parts}}{1 \text{ million}} =$$

$$4.4 \text{ ppm (parts Pb per million parts H}_2\text{O)}$$

4 Verify your results.

Work backwards. If you divide 4.4 by 1000 000 you get 4.4×10^{-6}. This result is the mass in grams of lead found per gram of water in the sample. Multiply by 500 g to find the total amount of lead in the sample. The result is 2.2×10^{-3}, which is the given number of grams of lead in the sample.

PRACTICE HINT

Be sure to keep track of units in all concentration calculations. In the case of mass-to-mass ratios, such as parts per million, the masses of solute and solvent must be expressed in the same units to obtain a correct ratio.

PRACTICE

PROBLEM SOLVING SKILL

1. Helium gas, 3.0×10^{-3} g, is dissolved in 200.0 g of water. Express this concentration in parts per million.
2. A sample of 300.0 g of drinking water is found to contain 38 mg Pb. What is this concentration in parts per million?
3. A solution of lead sulfate contains 0.425 g of lead sulfate in 100.0 g of water. What is this concentration in parts per million?
4. A 900.0 g sample of sea water is found to contain 6.7×10^{-3} g Zn. Express this concentration in parts per million.
5. A 365.0 g sample of water, contains 23 mg Au. How much gold is present in the sample in parts per million?
6. A 650.0 g hard-water sample contains 101 mg Ca. What is this concentration in parts per million?
7. An 870.0 g river water sample contains 23 mg of cadmium. Express the concentration of cadmium in parts per million.

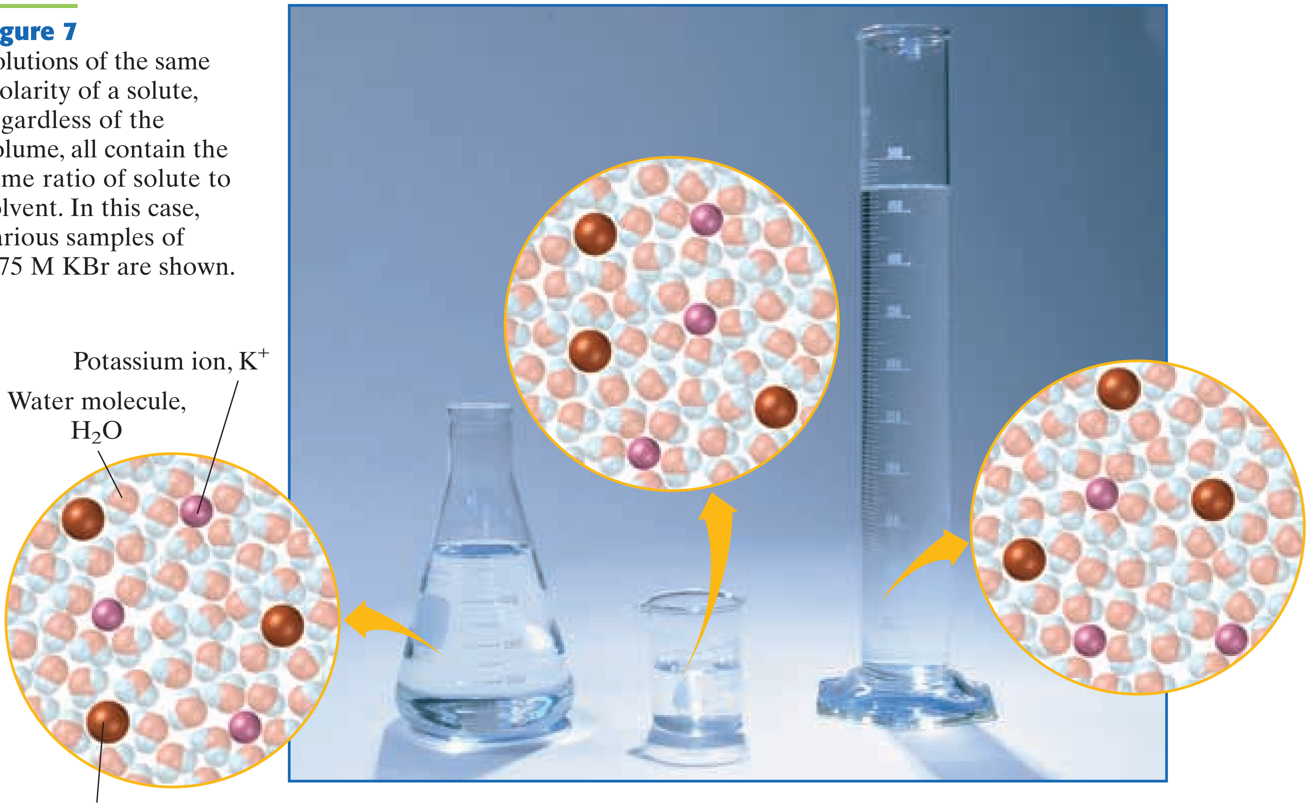


Figure 7
Solutions of the same molarity of a solute, regardless of the volume, all contain the same ratio of solute to solvent. In this case, various samples of 0.75 M KBr are shown.

Molarity

It is often convenient for chemists to discuss concentrations in terms of the number of solute particles in solution rather than the mass of particles in solution. Since the mole is the unit chemists use to measure the number of particles, they often specify concentrations using **molarity.** Molarity describes how many moles of solute are in each liter of solution.

molarity

a concentration unit of a solution expressed as moles of solute dissolved per liter of solution

Suppose that 0.30 moles of KBr are present in 0.40 L of solution. The molarity of the solution is calculated as follows:

$$\frac{0.30 \text{ mol KBr}}{0.40 \text{ L solution}} = 0.75 \text{ M KBr}$$

The symbol M is read as "molar" or as "moles per liter." Any amount of this solution has the same ratio of solute to solution, as shown in **Figure 7.**

Chemists often refer to the molarity of a solution by placing the formula for the solute in brackets. For example, $[CuSO_4]$ would be read as "the molarity of copper sulfate."

Topic Link

Refer to the "The Mole and Chemical Composition" chapter for a discussion of the use of the mole to express chemical amounts.

Preparing a Solution of Specified Molarity

Note that molarity describes concentration in terms of *volume of solution,* not volume of solvent. If you simply added 1.000 mol solute to 1.000 L solvent, the solution would not be 1.000 M. The added solute will change the volume, so the solution would not have a concentration of 1.000 M. The solution must be made to have exactly the specified volume of solution. The process of preparing a solution of a certain molarity is described in **Skills Toolkit 1.**

Preparing 1.000 L of a 0.5000 M Solution

Copper(II) sulfate, $CuSO_4$, is one of the compounds used to produce the chemiluminescence in light sticks. To make a 0.5000 M $CuSO_4$ solution, you need 0.5000 mol of the hydrate, $CuSO_4 \bullet 5H_2O$, for each liter of solution. To convert this amount of $CuSO_4 \bullet 5H_2O$ to a mass, multiply by the molar mass of $CuSO_4 \bullet 5H_2O$ (mass of $CuSO_4 \bullet 5H_2O$ = 0.5000 mol × 249.68 g/mol = 124.8 g).

Add some solvent (water) to the calculated mass of solute in the beaker to dissolve it, and then pour the solution into a 1.000 L volumetric flask.

Rinse the beaker with more water several times, and each time pour the rinse water into the flask until the solution almost reaches the neck of the flask.

Stopper the flask, and swirl thoroughly until all of the solid is dissolved.

Carefully fill the flask with water to the 1.000 L mark.

Restopper the flask, and invert the flask at least 10 more times to ensure complete mixing.

The solution that results has 0.5000 mol of $CuSO_4$ dissolved in 1.000 L of solution—a 0.5000 M concentration.

Calculating with Molarity

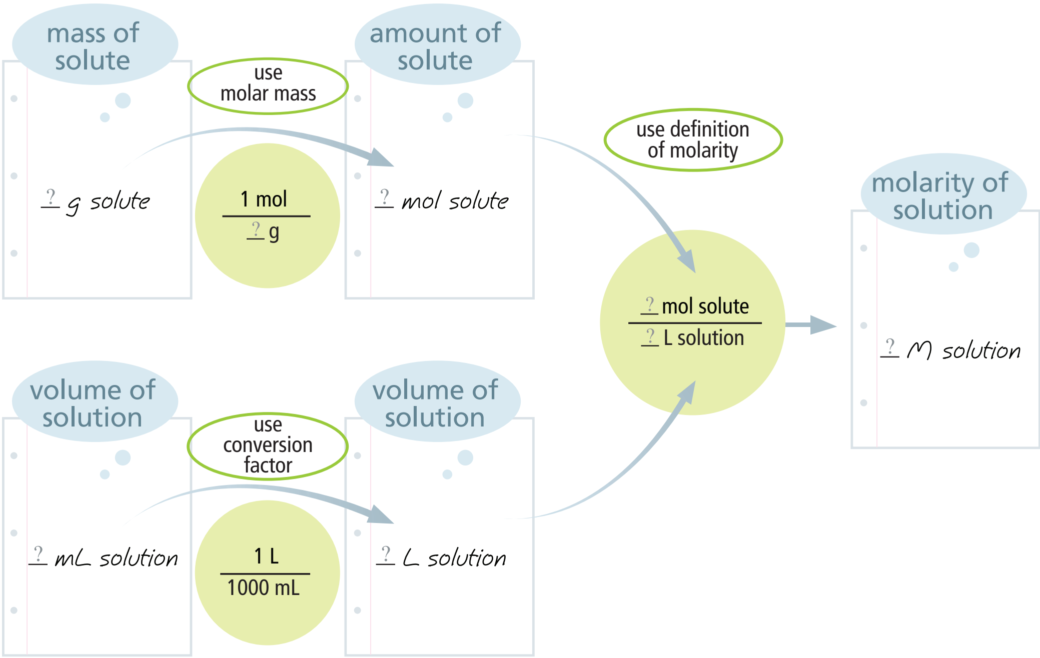
In working with solutions in chemistry, you will find that numerical calculations often involve molarity. The key to all such calculations is the definition of molarity, which is stated as an equation below.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

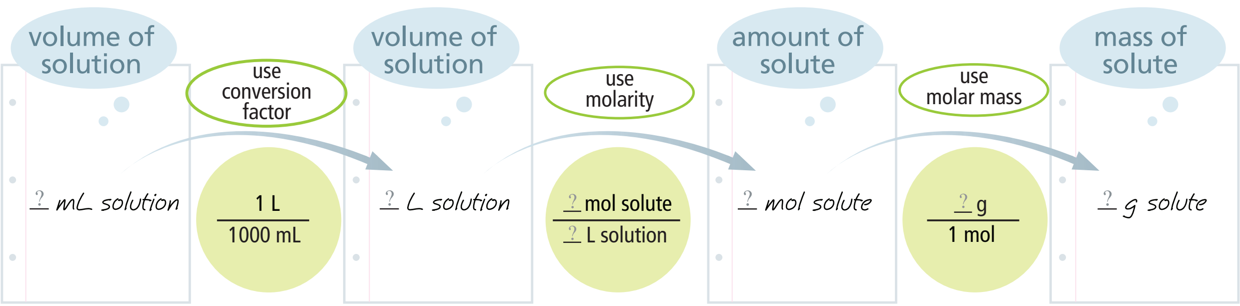
Skills Toolkit 2, below, shows how to use this equation in two common types of problems.

2 SKILLS Toolkit

1. Calculating the molarity of a solution when given the mass of solute and volume of solution



2. Calculating the mass of solute when given the molarity and volume of solution



SAMPLE PROBLEM B

Calculating Molarity

What is the molarity of a potassium chloride solution that has a volume of 400.0 mL and contains 85.0 g KCl?

1 Gather information.

volume of solution = 400.0 mL
mass of solute = 85.0 g KCl
molarity of KCl solution = ?

2 Plan your work.

Convert the mass of KCl into moles of KCl by using the molar mass:

$$85.0 \cancel{\text{g KCl}} \times \frac{1 \text{ mol}}{74.55 \cancel{\text{g KCl}}} = 1.14 \text{ mol KCl}$$

Convert the volume in milliliters into volume in liters:

$$400.0 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 0.4000 \text{ L}$$

3 Calculate.

Molarity is moles of solute divided by volume of solution:

$$\frac{1.14 \text{ mol KCl}}{0.4000 \text{ L}} = 2.85 \text{ mol/L} = 2.85 \text{ M KCl}$$

4 Verify your results.

As a rough estimate, 85 g divided by 75 g/mol is about 1 mol. If you divide 1 mol by 0.4 L you get 2.5 M. This approximation agrees with the answer of 2.85 M.

PRACTICE HINT

Remember to check that any solution volumes are converted to liters before you begin calculations that involve molarity.

PRACTICE

PROBLEM SOLVING SKILL

1. Vinegar contains 5.0 g of acetic acid, CH_3COOH, in 100.0 mL of solution. Calculate the molarity of acetic acid in vinegar.
2. If 18.25 g HCl is dissolved in enough water to make 500.0 mL of solution, what is the molarity of the HCl solution?
3. If 20.0 g H_2SO_4 is dissolved in enough water to make 250.0 mL of solution, what is the molarity of the sulfuric acid solution?
4. A solution of $AgNO_3$ contains 29.66 g of solute in 100.0 mL of solution. What is the molarity of the solution?
5. A solution of barium hydroxide, $Ba(OH)_2$, contains 4.285 g of barium hydroxide in 100.0 mL of solution. What is the molarity of the solution?
6. What mass of KBr is present in 25 mL of a 0.85 M solution of potassium chloride?
7. If all the water in 430.0 mL of a 0.45 M NaCl solution evaporates, what mass of NaCl will remain?

Topic Link

Refer to the "Stoichiometry" chapter for a discussion of stoichiometric calculations.

Using Molarity in Stoichiometric Calculations

There are many instances in which solutions of known molarity are used in chemical reactions in the laboratory. Instead of starting with a known mass of reactants or with a desired mass of product, the process involves a solution of known molarity. The substances are measured out by volume, instead of being weighed on a balance. An example of such an application in stoichiometry is shown in **Sample Problem C.**

SAMPLE PROBLEM C

Solution Stoichiometry

What volume (in milliliters) of a 0.500 M solution of copper(II) sulfate, $CuSO_4$, is needed to react with an excess of aluminum to provide 11.0 g of copper?

1 Gather information.

$[CuSO_4] = 0.500$ M
mass of product = 11.0 g Cu
solution volume = ? L

2 Plan your work.

Write the balanced chemical equation for the reaction:

$$3CuSO_4(aq) + 2Al(s) \longrightarrow 3Cu(s) + Al_2(SO_4)_3(aq)$$

Look up the molar mass of Cu:

$$\text{molar mass of Cu} = 63.55 \text{ g/mol}$$

Convert the mass of Cu to moles, and then use the mole ratio of $CuSO_4$:Cu from the balanced chemical equation to determine the number of moles of $CuSO_4$ needed. The moles of $CuSO_4$ can be converted into volume of solution using the reciprocal of molarity.

$$\text{g Cu} \times \frac{1 \text{ mol Cu}}{\text{g Cu}} \times \frac{\text{mol } CuSO_4}{\text{mol Cu}} \times \frac{\text{L solution}}{\text{mol } CuSO_4} = \text{L } CuSO_4$$

3 Calculate.

Substitute the values given:

$$11.0 \cancel{\text{ g Cu}} \times \frac{1 \cancel{\text{ mol Cu}}}{63.55 \cancel{\text{ g Cu}}} \times \frac{3 \cancel{\text{ mol } CuSO_4}}{3 \cancel{\text{ mol Cu}}} \times \frac{1 \cancel{\text{ L solution}}}{0.500 \cancel{\text{ mol } CuSO_4}} \times \frac{1000 \text{ mL solution}}{1 \cancel{\text{ L solution}}} = 346 \text{ mL } CuSO_4 \text{ solution}$$

4 Verify your results.

Work backwards. A volume of 0.346 L of a 0.500 M $CuSO_4$ solution contains 0.173 mol $CuSO_4$ (0.346 L × 0.500 M = 0.173 mol). A 0.173 mol sample of $CuSO_4$ contains 0.173 mol Cu, which has a mass of 11.0 g, so the answer is correct.

PRACTICE HINT

As in all stoichiometry problems, the mole ratio is the key. In solution stoichiometry, molarity provides the bridge between volume of solution and amount of solute.

PRACTICE

1 Commercial hydrochloric acid, HCl, is 12.0 molar. Calculate the mass of HCl in 250.0 mL of the solution.

2 An excess of zinc is added to 125 mL of 0.100 M HCl solution. What mass of zinc chloride is formed?

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

3 Yellow CdS pigment is prepared by reacting ammonium sulfide with cadmium nitrate. What mass of CdS can be prepared by mixing 2.50 L of a 1.25 M $Cd(NO_3)_2$ solution with an excess of $(NH_4)_2S$?

$$Cd(NO_3)_2(aq) + (NH_4)_2S(aq) \longrightarrow CdS(s) + 2NH_4NO_3(aq)$$

PROBLEM SOLVING SKILL

2 Section Review

UNDERSTANDING KEY IDEAS

1. Why did chemists develop the concept of molarity?

2. In what units is molarity expressed?

3. Describe in your own words how to prepare 100.0 mL of a 0.85 M solution of sodium chloride.

4. If you dissolve 2.00 mol KI in 1.00 L of water, will you get a 2.00 M solution? Explain your answer.

PRACTICE PROBLEMS

5. A sample of 400.0 g of water is found to contain 175 mg Cd. What is this concentration in parts per million?

6. If 1.63×10^{-4} g of helium dissolves in 100.0 g of water, what is the concentration in parts per million?

7. A standard solution of NaOH is 1.000 M. What mass of NaOH is present in 100.0 mL of the solution?

8. A 32 g sample of LiCl is dissolved in water to form 655 mL of solution. What is the molarity of the solution?

9. Most household bleach contains sodium hypochlorite, NaOCl. A 2.84 L bottle contains 177 g NaOCl. What is the molarity of the solution?

10. What mass of $AgNO_3$ is needed to prepare 250.0 mL of a 0.125 M solution?

11. Calcium phosphate used in fertilizers can be made in the reaction described by the following equation:

$$2H_3PO_4(aq) + 3Ca(OH)_2(aq) \longrightarrow Ca_3(PO_4)_2(s) + 6H_2O(aq)$$

What mass in grams of each product would be formed if 7.5 L of 5.00 M phosphoric acid reacted with an excess of calcium hydroxide?

CRITICAL THINKING

12. You have 1 L of 1 M NaCl, and 1 L of 1 M KCl. Which solution has the greater mass of solute?

13. Under what circumstances might it be easier to express solution concentrations in terms of molarity? in terms of parts per million?

14. One solution contains 55 g NaCl per liter, and another contains 55 g KCl per liter. Which solution has the higher molarity? How can you tell?

SECTION 3

Solubility and the Dissolving Process

KEY TERMS

- **solubility**
- **miscible**
- **immiscible**
- **dissociation**
- **hydration**
- **saturated solution**
- **unsaturated solution**
- **supersaturated solution**
- **solubility equilibrium**
- **Henry's law**

OBJECTIVES

1. **Identify** applications of solubility principles, and relate them to polarity and intermolecular forces.
2. **Explain** what happens at the particle level when a solid compound dissolves in a liquid.
3. **Predict** the solubility of an ionic compound by using a solubility table.
4. **Describe** solutions in terms of their degree of saturation.
5. **Describe** factors involved in the solubility of gases in liquids.

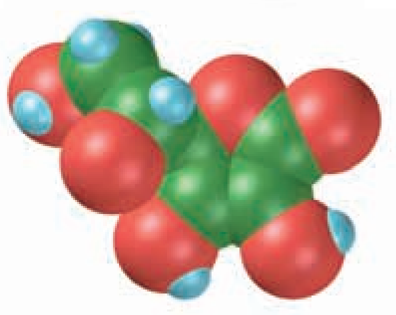
solubility

the ability of one substance to dissolve into another at a given temperature and pressure; expressed in terms of the amount of solute that will dissolve in a given amount of solvent to produce a saturated solution

Solubility and Polarity

Some pairs of liquids form a solution when they are mixed. For example, any amount of ethylene glycol, a common antifreeze, mixes with any amount of water to form antifreeze solutions in radiators. These two compounds are both very polar and have 100% **solubility** with each other.

Oils, such as cooking oil, do not mix with water. An oil is nonpolar, and water is polar. However, paint thinner is soluble with the oil in oil-based paints. Both the paint and paint thinner are nonpolar. Polar compounds tend to dissolve in other polar compounds, and nonpolar compounds tend to dissolve in other nonpolar compounds.



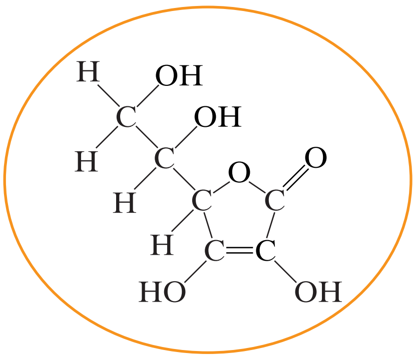


Figure 8
a The most common form of vitamin C is ascorbic acid, which is shown here.

b Because ascorbic acid is polar, it is very soluble in water but insoluble in fats and oils.

c Lemons, oranges, grapefruits, and limes are good sources of vitamin C.

Vitamin C Is a Water-Soluble Vitamin

The human body cannot make its own vitamin C; it must be obtained from external sources. Vitamin C also cannot be stored in the body. The disease scurvy, caused by a lack of vitamin C, has always been a threat to people with a limited diet. In 1747, Dr. James Lind studied the effect of diet on sailors who had scurvy. Those whose diet included citrus fruits recovered. In 1795, long before people knew that citrus fruits were rich in vitamin C, the British navy began to distribute lime juice during long sea voyages. For this reason, British sailors were often called "limeys."

Vitamin C was isolated and identified in the early 1930s by American chemists. The most important function of vitamin C is in the synthesis of collagen, a protein that makes up tendons and that enables muscle movements. **Figure 8** on the previous page shows that vitamin C has several –OH groups. These –OH groups form strong hydrogen bonds with the –OH groups in water, so vitamin C is very soluble in water. At room temperature, 33 g of vitamin C will dissolve in 100 g of water. Any excess in the diet is quickly eliminated by the kidneys. It is almost impossible to overdose on vitamin C.

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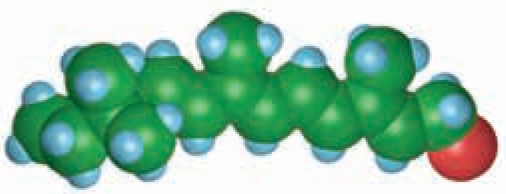
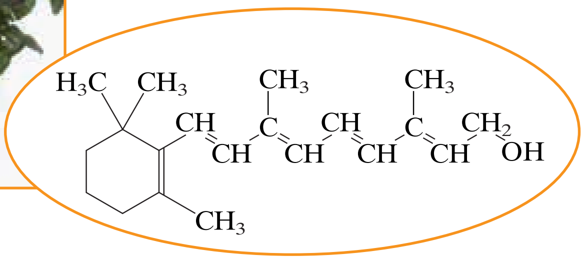
Vitamin A Is a Fat-Soluble Vitamin

Vitamin A also must be obtained in food, especially yellow vegetables. It has many functions in the body, and is essential for good vision. Vitamin A is also needed for the respiratory tract, skin, and for normal growth of bones. Fortunately, vitamin A is fairly abundant in foods.

Vitamin A has a long, nonpolar carbon-hydrogen chain, as shown in **Figure 9.** Consequently, it has very low solubility in water. Its nonpolarity makes it very soluble in fats and oils, which are also nonpolar. Any excess of vitamin A in the diet builds up in body fat and is not easily eliminated from the body. So much can accumulate in fat that the amount of vitamin A may become toxic. So, as with other fat-soluble vitamins, it is possible to take too much vitamin A.



b Vitamin A is also known as retinol because it plays a vital role in helping the retina of your eye detect light.



c The vitamin A molecule is composed of mostly carbon and hydrogen, which makes the molecule nonpolar.

Figure 9
a Sources of vitamin A include dark green leafy vegetables, carrots, broccoli, tomatoes, and egg yolks.

Figure 10
a The nonpolar hydrocarbons in paint dissolve in the nonpolar oil in the paint thinner.
b Oil and water do not mix because oil is nonpolar and water is very polar.

The Rule Is "Like Dissolves Like"

In nonpolar molecules, such as vitamin A, London forces are the only forces of attraction between molecules. When nonpolar molecules are mixed with other nonpolar molecule, the intermolecular forces of the molecules easily match. Thus, nonpolar molecules are generally soluble with each other, as shown in **Figure 10a.** This is one part of the rule "like dissolves like": liquids that are completely soluble with each other are described as being **miscible** in each other.

miscible

describes two or more liquids that are able to dissolve into each other in various proportions

If molecules are sufficiently polar, there is an additional electrical force pulling them toward each other. The negative partial charge on one side of a polar molecule attracts the positive partial charge on the other side of the next polar molecule. If you add polar molecules to other polar molecules, such as water, the attraction between the two is strong. An example is vitamin C dissolving in water. This is another part of the rule "like dissolves like": polar molecules dissolve other polar molecules.

Topic Link

Refer to the "States of Matter and Intermolecular Forces" chapter for a discussion of intermolecular forces.

However, if you try to mix oil and water, the nonpolar oil molecules do not mix with the polar water molecules. The two liquids are **immiscible.** They form two layers, as shown in **Figure 10b.** The polar water molecules attract each other, so they cannot be pushed apart by the nonpolar oil molecules to form a solution.

immiscible

describes two or more liquids that do not mix with each other

Miscibility can be difficult to determine for some substances. Ethyl alcohol, CH_3CH_2OH, is sufficiently polar to be completely miscible with water. But the alcohol octanol, $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2OH$, has a long nonpolar tail which causes it to be only slightly soluble in water.

Solubilities of Solid Compounds

Even two polar liquids placed in the same container may not dissolve in each other rapidly. Their strong intermolecular forces can only act on nearby molecules—not between molecules at the top of the container and those at the bottom.

The speed of the process can be increased by shaking the mixture. This action breaks the two liquids into small droplets and thereby increases the amount of contact between the surfaces of the liquids. This process works because the only place that dissolving can occur is at the surface between the two liquids, where the different molecules are near each other.

Similarly, in considering the solubility of solids in liquids, the only place where dissolution can occur is at the surface of the solid particles. The solid must be broken into smaller particles and then into molecules or ions, which can form a solution with the solvent molecules.

Greater Surface Area Speeds Up the Dissolving Process

As the discussion above indicated, the only place where dissolving can take place is at the surfaces where solute and solvent molecules are in contact. So if a solid has been broken into small particles, the surface area is much greater and the rate of the dissolving process is increased.

This is illustrated in **Figure 11.** The sugar granules dissolve more quickly than the sugar cubes. Because the sugar granules have more surface area, more of their molecules are directly exposed to the solvent and the dissolving process takes place faster. In the case of sugar cubes, most of the sugar molecules are inside the cubes and cannot dissolve until after the molecules at the outside of the cubes are dissolved.

Figure 11
Sugar granules dissolve in water more quickly than sugar cubes, because sugar granules have more surface area than sugar cubes.

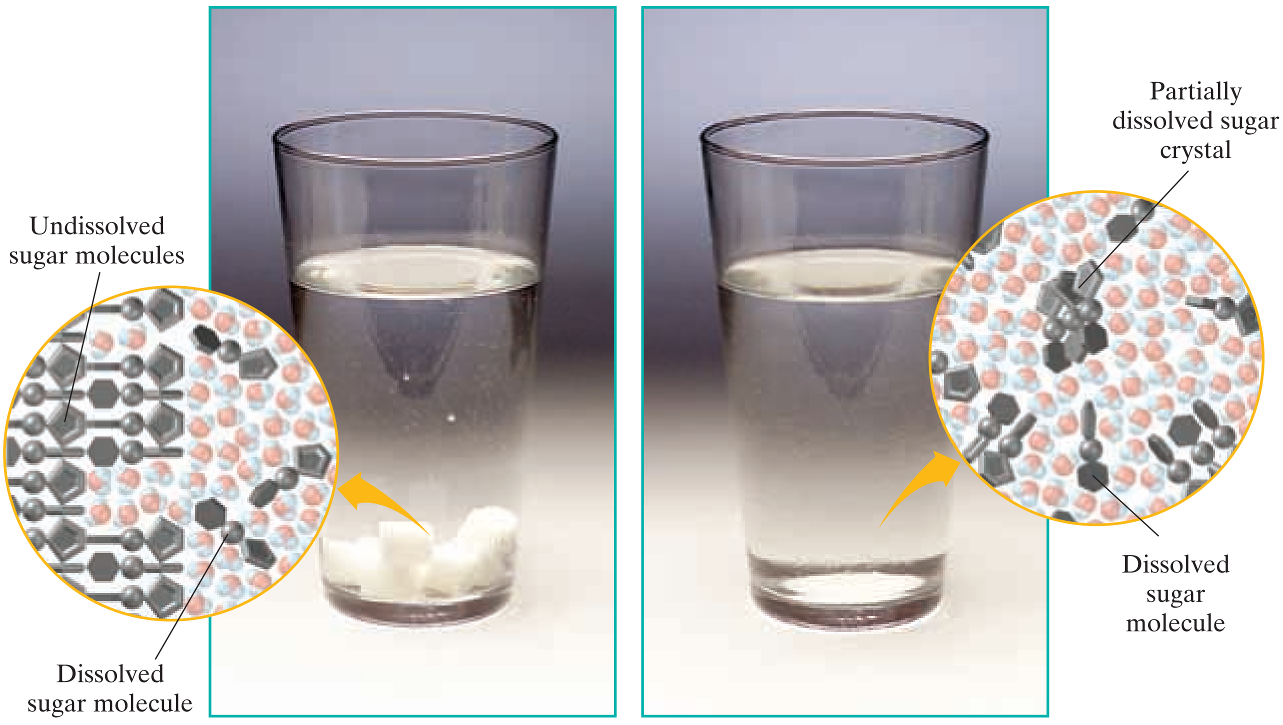
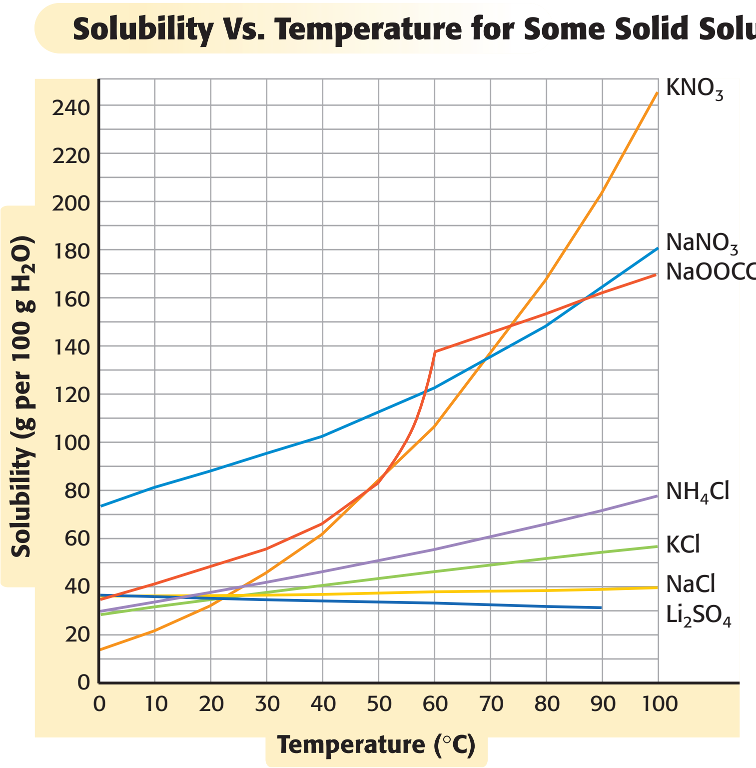


Figure 12
The effect of temperature on the solubility of some ionic solids is graphed here.



Solubilities of Solids Generally Increase with Temperature

Another way to make most solids dissolve more and faster is to increase the temperature. Increasing the temperature is effective because, in general, solvent molecules with greater kinetic energy can dissolve more solute particles. **Figure 12** shows how an increase in temperature affects the solubility of several ionic compounds. In most cases, such as in the case of KNO_3, the solubility increases with temperature. However, temperature has little effect on the solubility of NaCl. The solubility of Li_2SO_4 actually decreases slightly as temperature increases.

Topic Link

Refer to the "Causes of Change" chapter for a discussion of enthalpy and entropy.

dissociation

the separating of a molecule into simpler molecules, atoms, radicals, or ions

hydration

the strong affinity of water molecules for particles of dissolved or suspended substances that causes electrolytic dissociation

Both Enthalpy and Entropy Affect the Solubility of Salts

Until now, we have not made a distinction between the dissolving process of a covalent solid, such as sugar, and that of an ionic solid, such as table salt. Surface area and temperature affect both covalent and ionic solids. However, the dissolving of an ionic compound involves a unique factor: the separation of ions from the lattice into individual dissolved ions. This process, called **dissociation,** can be represented as an equation.

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

If water is the solvent, as above, dissociation involves **hydration,** the surrounding of the dissociated ions by water molecules.

The actual process of dissociation, however, is more complex. It takes a large amount of energy to separate the ions. The separation requires a large positive enthalpy change, ΔH. The polar ends of the water molecules approach the ions and release energy, and this ΔH is very negative. The ΔH changes nearly cancel.

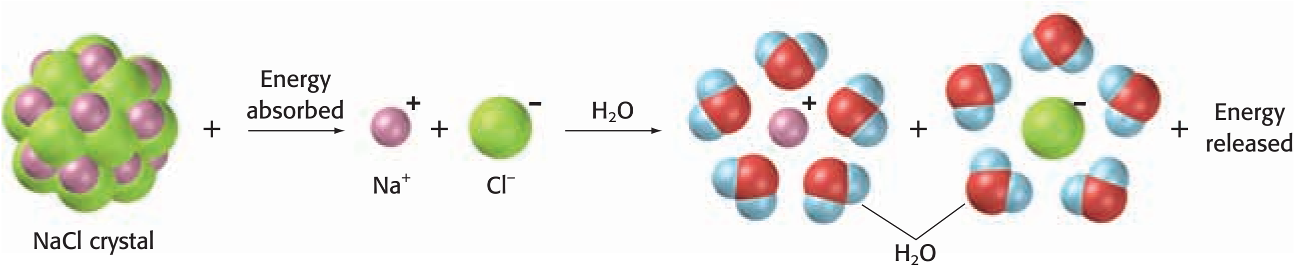


Figure 13
Individual ions are separated from the solid lattice by absorbed energy before they are hydrated by water molecules. Energy is released when hydrated ions are removed from the lattice.

Entropy increases as the ions are scattered throughout the solution. On the other hand, there is a large decrease in entropy as the water molecules are structured around the ions. Smaller ions, such as the Na^+ ions in **Figure 13,** have a greater decrease in entropy than larger ions. The net result of all of the enthalpy and entropy changes that accompany the dissolving process determines the solubility of an ionic solid.

Solubilities of Ionic Compounds

Solubilities are difficult to predict because of the many factors involved, so they must be measured experimentally. From experimental results of ionic solubilities in water, some patterns emerge, as shown in **Table 2.** Categories such as *soluble* and *insoluble* can be useful in many cases. The solubility of $Ba(OH)_2$ is 3.5 g per 100 g of water. It is described as slightly soluble. However, most substances are, at least to some extent, soluble in everything else. Even glass is very slightly soluble in water. In some delicate measurements, glass cannot be used as a container.

Table 2 Solubility Rules for Some Common Ionic Compounds

Compounds containing these ions are soluble in water:
Acetates, $CH_3CO_2^-$, except that of Fe^{3+}
Alkali metals (Group 1), except LiF
Ammonium, NH_4^+
Bromides, Br^-, except those of Ag^+, Pb^{2+}, and Hg_2^{2+}
Chlorides, Cl^-, except those of Ag^+, Pb^{2+}, and Hg_2^{2+}
Nitrates, NO_3^-
Sulfates, SO_4^{2-}, except those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, and Hg_2^{2+}
Compounds containing these ions are insoluble in water:
Carbonates, CO_3^{2-}, except those of Group 1 and NH_4^+
Chromates, CrO_4^{2-}, except those of Group 1 and NH_4^+
Hydroxides, OH^-, except those of Group 1
Oxides, O^{2-}, except those of Group 1, Ca^{2+}, Sr^{2+}, and Ba^{2+} (which form hydroxides)
Phosphates, PO_4^{3-}, except those of Group 1 and NH_4^+
Sulfides, S^{2-}, except those of Group 1, Mg^{2+}, Ca^{2+}, Ba^{2+}, and NH_4^+

For a more detailed solubility table, see Appendix A.

Figure 14
a When a solution is saturated, additional solute added to the solvent will remain undissolved.

b As long as a solution is unsaturated, more solute can be added to the solvent and be dissolved.

Saturation

If you look in a chemistry handbook, or check the table in Appendix A, you will find that the solubility of potassium chloride, KCl, is 23.8 grams per 100 grams of water, at 20°C. This suggests some sort of limit.

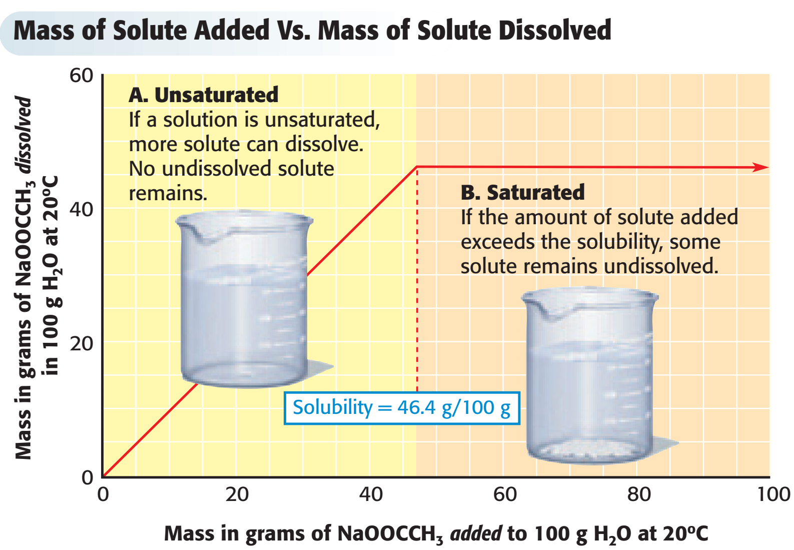
When the maximum amount of solute, such as the 23.8 g of KCl mentioned above, is dissolved in a solution, the solution is said to be a **saturated solution.** As shown in **Figure 14a,** if a solution is saturated, any additional solute that is added collects at the bottom of the container. If more solute can be added to a solution and dissolve, as in **Figure 14b,** the solution is considered to be an **unsaturated solution. Figure 15** illustrates the relationship between solute added and solute dissolved.

saturated solution
a solution that cannot dissolve any more solute under the given conditions

unsaturated solution
a solution that contains less solute than a saturated solution and that is able to dissolve additional solute

The amount of solute that can dissolve depends on the forces between the solute particles and on forces between the solute particles and the solvent particles. When a solute is placed in contact with a solvent, molecules or ions of solute dissolve into the solvent. As soon as this happens, these same dissolved ions or molecules are capable of rejoining the undissolved solute. As the concentration of solute increases, the rate of return to the solute increases.

Figure 15
This graph shows the masses of solute that can be dissolved before saturation is reached. Any additional solute added beyond this point will remain undissolved.



Solubility Can Be Exceeded

In a saturated solution, some excess solute remains undissolved, and the mass that dissolves is equal to the solubility value for that temperature. Under special conditions, **supersaturated solutions** can also exist. Supersaturated solutions have more solute dissolved than the solubility indicates would normally be possible, but only as long as there is no excess undissolved solute remaining.

supersaturated solution

a solution holding more dissolved solute than what is required to reach equilibrium at a given temperature

Supersaturation is the reason why hand warmers, such as those shown in **Figure 16,** work. Inside the plastic pack, 60 g of sodium acetate, $NaOOCCH_3$, has been combined with 100 mL of water. This amount of sodium acetate is more than the amount that can dissolve in 100 mL of water at 20°C. As shown in **Figure 12,** only about 48 g $NaOOCCH_3$ can dissolve in 100 mL of water. When the solution is heated to 100°C, all of it dissolves, as shown in **Figure 16a.** When the solution is allowed to cool to 20°C, crystals of solute should form. **Figure 16b** shows that crystals do not form. Instead, the solution becomes supersaturated. However, if you disturb the cooled solution by clicking the disk in the center of the pack, crystallization immediately occurs, as shown in **Figure 16c.** The recrystallization of sodium acetate is exothermic, so its reappearance in **Figure 16c** releases heat.

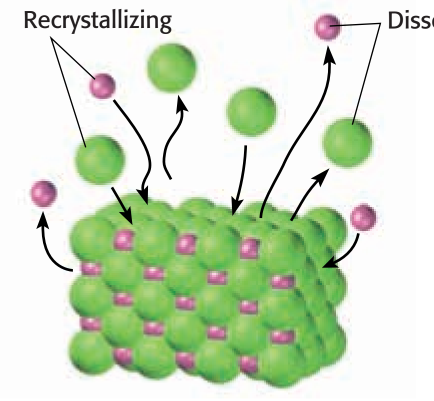
Figure 16
On a cold day it can be comforting to use a hand warmer. Hand warmers use the principle of supersaturation to provide heat.

a At 100°C, 60 g of $NaOOCCH_3$ will dissolve completely in the 100 mL of water contained in the pack.

b When the solution is cooled to 20°C, $NaOOCCH_3$ does not crystallize unless the solution is disturbed.

c Clicking the disk in the center of the pack triggers rapid recrystallization, and heat is released. The hand warmer can be reused if it is heated above the saturation point again.

Figure 17
In a saturated solution, the solute is recrystallizing at the same rate that it is dissolving.



Saturation Occurs at a Point of Solubility equilibrium

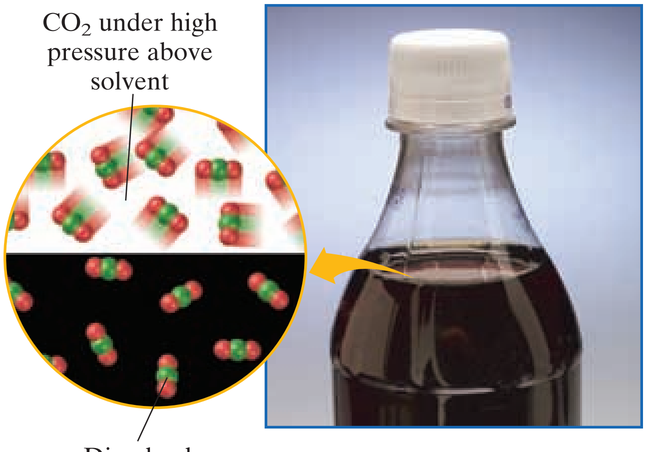
In a saturated solution, solute particles are dissolving and recrystallizing at the same rate. It is a state of *dynamic equilibrium.* There is constant exchange, yet there is no net change.

When the amount of solute added to a solvent has reached its solubility limit, it is understood that the particles of solute in solution are in dynamic equilibrium with excess solute. This is illustrated in **Figure 17.** Na^+ ions and Cl^- ions are leaving the solid surface at the same rate as ions are also returning to the pile of excess solute at the bottom. These ions are considered to be in **solubility equilibrium.**

solubility equilibrium
the physical state in which the opposing processes of dissolution and crystallization of a solute occur at equal rates

Gases Can Dissolve in Liquids

When you first look at an unopened bottle of soda, you see very few bubbles. The liquid is homogeneous, as shown in **Figure 18a.** But when you open the bottle, you can hear gas escaping. Then you see many bubbles rising in the liquid, as pictured in **Figure 18b.** You probably know that the bubbles from soda are carbon dioxide, CO_2.



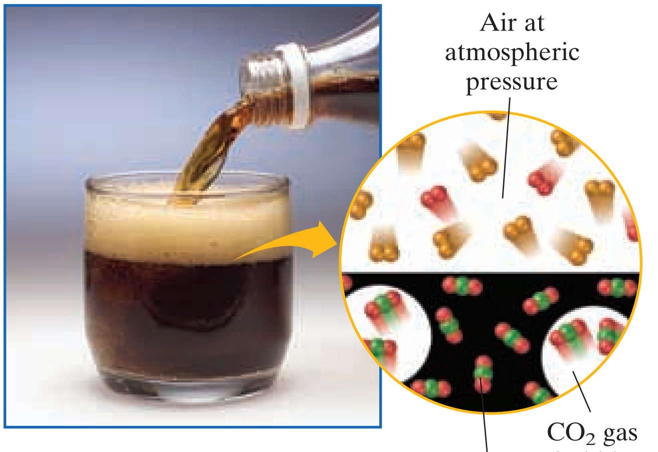


Figure 18
a There are no bubbles in an unopened bottle of soda because carbon dioxide is dissolved in the liquid.

b When the bottle cap is removed, the pressure inside the bottle decreases rapidly. Carbon dioxide escapes due to its lowered solubility.

Gas Solubility Depends on Pressure and Temperature

Because a gas escapes when you open the bottle, you know that there is gaseous carbon dioxide present above the solution. Eventually, almost all the gas escapes when the bottle is opened to the atmosphere. Why?

In a gas, there is low attraction between the molecules. Likewise, there is usually little attraction between molecules of a gas and molecules of a liquid solvent. **Henry's law** states that the solubility of a gas increases as the partial pressure of the gas on the surface of the liquid increases.

At the high pressure of CO_2 in the unopened can, the gaseous CO_2 is in equilibrium with the dissolved gas. Therefore, the solution is saturated. When the bottle is opened and the pressure is released, the solubility decreases, and CO_2 bubbles escape. Finally, the dissolved CO_2 comes to equilibrium with the carbon dioxide in the air. The solution is saturated, but at this lower pressure, it is at a much lower concentration.

Temperature also affects gas solubility. After the soda bottle is open and becomes warm, the soda forms fewer bubbles and tastes flat. Even if you compare the taste of a newly opened warm soda with that of a newly opened cold soda, you will find that the warm soda will taste somewhat flat. Warm soda tastes flat because there is less CO_2 dissolved in it. Gases are less soluble in a liquid of higher temperature because the increased molecular motion in the solution allows gas molecules to escape their loose association with the solvent molecules.

Henry's law

the law that states that at constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas on the surface of the liquid

Topic Link

Refer to the "Gases" chapter for a definition and discussion of partial pressure.

3 Section Review

UNDERSTANDING KEY IDEAS

1. Why is it possible to overdose on Vitamin A, but not on Vitamin C?

2. Why is ethanol miscible in water?

3. Why do sugar cubes dissolve more slowly than granulated sugar?

4. What factors are involved in determining the solubility of an ionic salt?

5. Would the compound $MgSO_4$ be considered soluble in water?

6. Would the compound PbS be considered soluble in water?

7. You keep adding sugar to a cold cup of coffee and stirring the coffee. Finally, solid sugar remains on the bottom of the cup. Explain why no more sugar dissolves.

8. What is the relation between supersaturation and the hand warmer in **Figure 16?**

CRITICAL THINKING

9. Ethylene glycol is represented by the formula $HOCH_2CH_2OH$. Is it likely to be soluble in water? in paint thinner? Explain your answer.

10. A solution of $BaCl_2$ is added to a solution of $AgNO_3$. Use **Table 2** to decide what reaction happens, and write a balanced equation.

11. Suppose a salt, when dissolving, has a positive ΔH and a negative ΔS. Is the solubility of the salt low or high?

12. A commercial "fizz saver" pumps helium under pressure into a soda bottle to keep gas from escaping. Will this keep CO_2 in the soda bottle? Explain your answer.

SECTION 4

Physical Properties of Solutions

KEY TERMS

- **conductivity**
- **electrolyte**
- **nonelectrolyte**
- **hydronium ion**
- **colligative property**
- **surfactant**
- **detergent**
- **soap**
- **emulsion**

OBJECTIVES

1. **Distinguish** between nonelectrolytes, weak electrolytes, and strong electrolytes.
2. **Describe** how a solute affects the freezing point and boiling point of a solution.
3. **Explain** how a surfactant stabilizes oil-in-water emulsions.

Electrical Conductivity in Solutions

Some substances conduct electricity and some cannot. The **conductivity** of a substance depends on whether it contains charged particles, and these particles must be able to move. Electrons move freely within a metal, thus allowing it to conduct electricity. Solid NaCl contains ions, but they cannot move, so solid NaCl is a nonconductor by itself. But an aqueous solution of ionic compounds such as NaCl contains charged ions, which can move about. Solutions of ionic compounds conduct electricity. Pure water does not conduct electricity.

conductivity
the ability to conduct an electric current

Electrolytes Provide Ions in Solution

An **electrolyte** is a substance that dissolves in a liquid solvent and provides ions that conduct electricity. For example, sports drinks such as the one pictured in **Figure 19** contain electrolytes that your body needs replenished after strenuous physical activity. Electrolytes are considered to belong to one of two classes depending on their tendency to dissociate.

electrolyte
a substance that dissolves in water to give a solution that conducts an electric current

Figure 19
A sports drink not only supplies water but also supplies electrolytes.

Strong electrolytes completely dissociate into ions and conduct electricity well. *Weak electrolytes* provide few ions in solution. Therefore, even in high concentrations, solutions of weak electrolytes conduct electricity weakly. Ionic compounds are usually strong electrolytes. Covalent compounds may be strong electrolytes, weak electrolytes, or nonconductors.

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Electrical Conductivities Span a Wide Range

As shown in **Figure 20,** the extent to which electrolytes dissociate into ions is indicated by the conductivity of their solutions. The apparatus shown has a light bulb attached to a battery, and there is a gap in the circuit between two electrodes. The electrodes are dipped in a solution. If the solution conducts electricity, the circuit is completed and the bulb lights. The amount of current that can be carried depends on the concentration of ions in the solution. A solution of a strong electrolyte has a high concentration of ions, so the bulb lights up brightly. A solution of a weak electrolyte has a low concentration of ions, so the bulb lights up dimly.

Virtually all of a strong electrolyte dissociates as it dissolves in a solvent. Sodium chloride, for example, ionizes completely in solution:

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

The ions in solution can move about. NaCl is a strong electrolyte, and a solution of NaCl can conduct electricity. The sugar sucrose, on the other hand, does not ionize at all in solution. It is a **nonelectrolyte,** and does not conduct electricity.

nonelectrolyte

a liquid or solid substance that does not alow the flow of an electric current, either in solution or in its pure state, such as water or sucrose

Figure 20
All four solutions have the same concentration. The brightness of the bulb indicates the degree of conduction and the degree of dissociation (ionization).

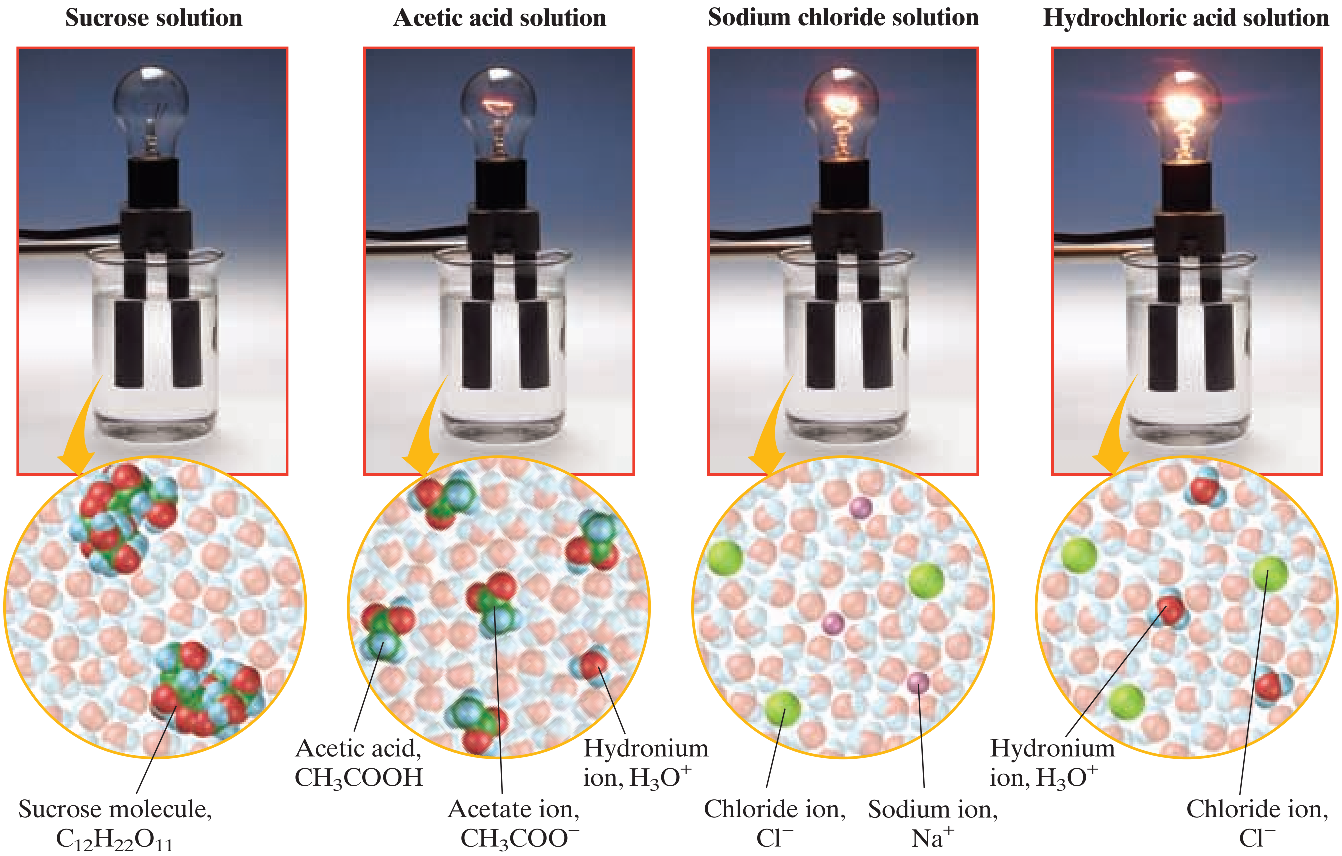
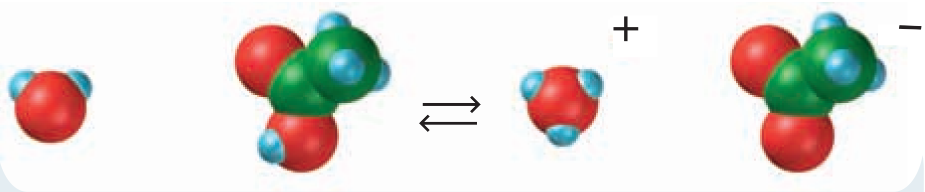


Figure 21 When acetic acid dissolves in water, very little of it is changed into ions.



Water		Acetic acid		Hydronium ion		Acetate ion
$H_2O(l)$	+	$CH_3COOH(aq)$	$\rightleftharpoons$	$H_3O^+(aq)$	+	$CH_3COO^-(aq)$

hydronium ion

an ion consisting of a proton combined with a molecule of water; H_3O^+

Acids react with water to form the **hydronium ion,** H_3O^+. The reaction of acetic acid with water is shown in **Figure 21.** Acetic acid is a weak electrolyte. In water, only about 1% of acetic acid molecules ionize.

Hydrogen chloride dissolves in water to form a strongly conducting solution called *hydrochloric acid.* Hydrogen chloride is a strong electrolyte because it ionizes completely, as shown by the following equation:

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Keep in mind that the use of the terms *strong* and *weak* have nothing to do with concentration. The term *strong* means that the substance provides a high *proportion* of ions in solution. Hydrochloric acid is a strong electrolyte at any concentration.

Tap Water Conducts Electricity

Have you wondered why you are warned not to use electrical appliances when you are near water? Have you wondered why you are also warned not to go swimming when a thunderstorm is near? The reason is because of the electrolytes in the water. Sea water, as shown in **Figure 22,** also conducts electricity.

Figure 22 The electricity from lightning is conducted through groundwater, ponds, or ocean water, which all contain electrolytes.

If you collect rainwater in a relatively unpolluted area, you will discover that the rainwater is essentially a nonconductor of electricity. A small concentration of carbonic acid from the carbon dioxide in the air added to the rainwater causes the rain water to be a weak conductor. Pure rainwater conducts almost as poorly as distilled water. However, most of the water we use comes from wells, lakes, or rivers. This water has been in contact with soil and rocks, which contain ionic compounds that dissolve in the water. Consequently, tap water conducts electricity. The conduction is not high, but the water can conduct enough current to stop a person's heart. So, for example, a person should not use an electrical appliance when in the bathtub or shower.

You should also not seek shelter under a tree during a thunderstorm. The tree not only sticks up like a lightning rod, but the sap in the tree also contains electrolytes, and conducts the electricity. Lightning finds a path to the ground through the trunk of the tree.

Colligative Properties

A solution made by dissolving a solute in a liquid, such as adding sulfuric acid to water, has particular chemical properties that the solvent alone did not have. The *physical properties* of water, such as how well the water mixes with other compounds, are also changed when substances dissolve in it.

As shown in **Figure 23,** salt can be added to icy sidewalks to melt the ice. The salt actually lowers the freezing point of water. Therefore, ice is able to melt at a lower temperature than it normally would. This change is called *freezing-point depression.* Nonvolatile solutes such as salt also increase the boiling point of a solvent. This change is called *boiling-point elevation.* For example, glycol in a car's radiator increases the boiling point of water in the radiator, which prevents overheating. It also lowers the freezing point, preventing freezing in cold weather.

Figure 23
Dissolved salt lowers the freezing point of water, causing it to melt at a lower temperature than it normally would.

Only the Concentration of Dissolved Particles Is Important

colligative property

a property of a substance or system that is determined by the number of particles present in the system but independent of the properties of the particles themselves.

Any physical effect of the solute on the solvent is a **colligative property.** The lowering of the freezing point and the raising of the boiling point are examples of colligative properties. Only nonvolatile solutes have predictable effects on boiling point, but besides that requirement, the identity of the solute is relatively unimportant.

Any solute, whether an electrolyte or a nonelectrolyte, contributes to the colligative properties of the solvent. The degree of the effect depends on the concentration of solute particles (either molecules or ions) in a certain mass of solvent. The greater the particle concentration is, the greater the boiling-point elevation or the freezing-point depression is. For example, based on the number of moles of solute particles, 1 mol of sodium chloride, NaCl, is expected to give twice the amount of change as 1 mol of sucrose, $C_{12}H_{22}O_{11}$. This result occurs because NaCl dissolves to give two moles of particles per mole, and sucrose dissolves to give only one. Likewise, 1 mol of calcium chloride, $CaCl_2$, has about three times the effect as 1 mol of sucrose because $CaCl_2$ dissolves to give three dissolved particles per mole. The following equations illustrate the logic.

$$C_{12}H_{22}O_{11}(s) \longrightarrow C_{12}H_{22}O_{11}(aq) \qquad \text{(1 dissolved particle)}$$

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq) \qquad \text{(2 dissolved particles)}$$

$$CaCl_2(s) \longrightarrow Ca^{2+}(aq) + 2Cl^-(aq) \qquad \text{(3 dissolved particles)}$$

Figure 24, below, illustrates the differing numbers of solute particles generated by equal concentrations of each compound in solution.

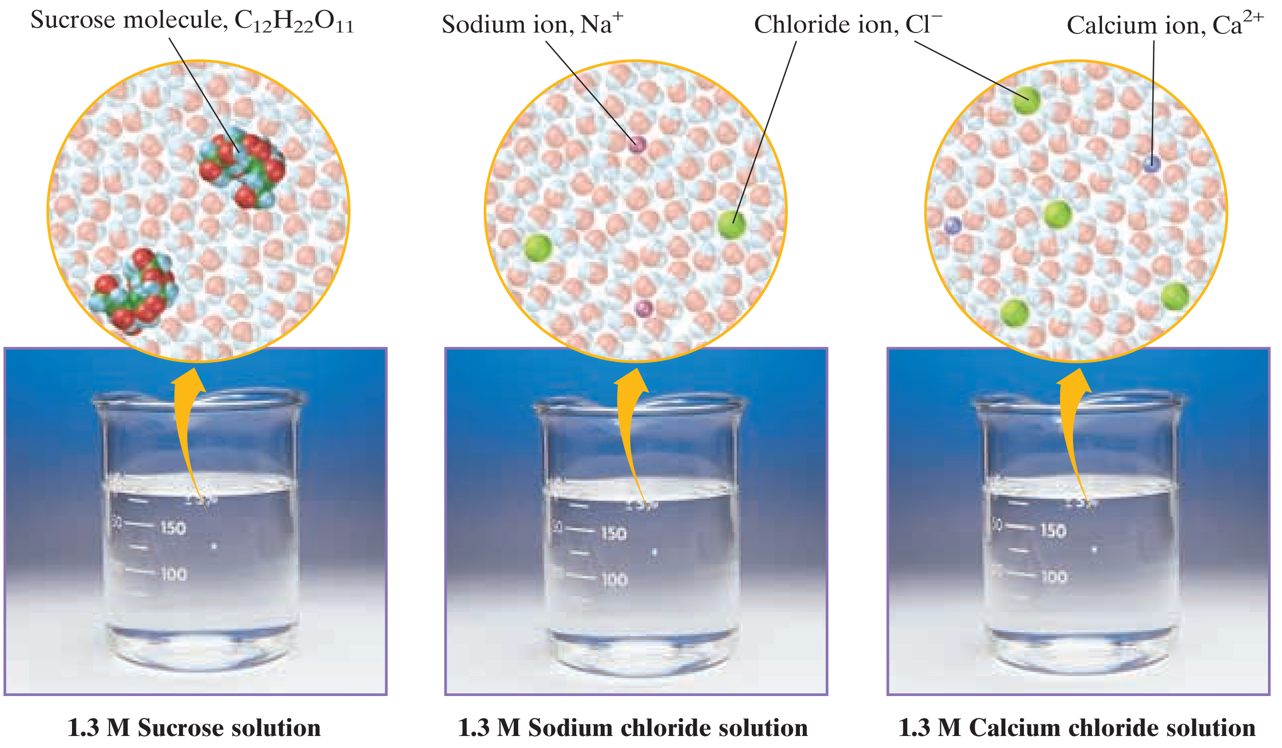


Figure 24
Colligative properties depend on the concentration of solute particles. Equal amounts in moles of sugar, table salt, and calcium chloride affect the solvent in different degrees because of the different numbers of particles they form when dissolved.

Dissolved Solutes Lower the Vapor Pressure of the Solvent

Colligative properties are all caused by a decrease in the vapor pressure of the solvent. Recall that all gases exert pressure, and vapor pressure is the pressure exerted by the vapor in equilibrium with its liquid state at a given temperature. The effect of a dissolved solute on the vapor pressure of a solvent can be understood when you consider the number of solvent particles in a solution. A solution has fewer solvent particles per volume than the pure solvent has, so fewer solvent particles are available to vaporize. Vapor pressure will therefore be decreased in proportion to the number of solute particles.

Figure 25 illustrates the difference between water's boiling point and freezing point (red lines), and the boiling point and freezing point of an aqueous solution (blue lines). Recall that the boiling point of a liquid is the temperature at which the liquid's vapor pressure is equal to the atmospheric pressure above the liquid. Because the vapor pressure of the water is lowered by the addition of a nonvolatile solute, the solution must be heated to a higher temperature for its vapor pressure to reach atmospheric pressure, at which point the solution boils.

The freezing point is the temperature at which water and ice are in equilibrium. Ice has a vapor pressure that is indicated by the line down to the left in **Figure 25.** The freezing point of water is the temperature at which the vapor pressure of pure water and ice are equal. Because the vapor pressure of the solution is lower, the vapor pressure of the solution intersects the line for the vapor pressure of ice at a lower temperature. Ice and water in the solution are in equilibrium at a lower temperature. The freezing point of the solution is therefore lower than that of pure water.

Topic Link

Refer to the "States of Matter and Intermolecular Forces" chapter for a discussion of vapor pressure.

Solute Effects on the Vapor Pressure of a Pure Solvent

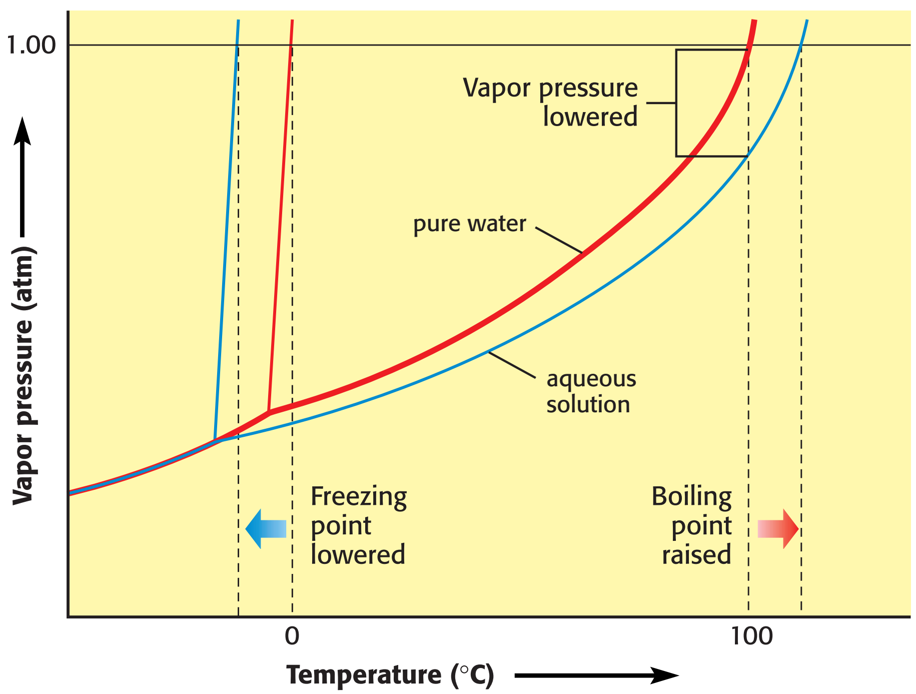


Figure 25
This is a modified phase diagram for pure water (red lines) and for an aqueous solution (blue lines). The addition of a solute has the effect of extending the range of the liquid phase.

Surfactants

Have you ever tried to wash your hands when they were very dirty without using soap? You were probably not very successful. Perspiration contains water and oils. The water evaporates, but the oil remains behind and coats the dirt particles. Oil and water do not mix, so washing without soap does not clean very well. However, if you use soap, the cleaning process is much more successful. Why?

The action of scrubbing your skin breaks the oil into tiny droplets. Soap molecules contain long nonpolar hydrocarbon chains, which are soluble in the nonpolar oil. As shown in **Figure 26,** soap molecules also have negatively charged ends, which are soluble in the water just outside the oil droplet. The negatively charged droplets repel each other and are carried away from the skin, along with any dirt that was on your skin.

Soap belongs to a general class of substances called **surfactants.** A surfactant is a substance that concentrates at the interface between two phases, either the solid-liquid, liquid-liquid, or gas-liquid phase. A **detergent** is a surfactant that is used for cleaning purposes. Usually, when we speak of detergents, we are talking about *synthetic detergents,* substances that are not natural products. A **soap** is a particular type of detergent and one that is a natural product. Soaps are sodium or potassium salts of fatty acids with long hydrocarbon chains. The formula for a typical soap, sodium palmitate, is shown below.

$$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2COO^-Na^+$$

Soap is an emulsifying agent. An **emulsion** is made of colloid-sized droplets suspended in a liquid in which they would ordinarily be insoluble, unless stabilized by an *emulsifying agent,* such as a soap. Without an emulsifying agent, polar and nonpolar molecules remain separate, as pictured in **Figure 27** on the next page.

surfactant
a compound that concentrates at the boundary surface between two immiscible phases, solid-liquid, liquid-liquid, or liquid-gas

detergent
a water-soluble cleaner that can emulsify dirt and oil

soap
a substance that is used as a cleaner and dissolves in water

emulsion
any mixture of two or more immiscible liquids in which one liquid is dispersed in the other

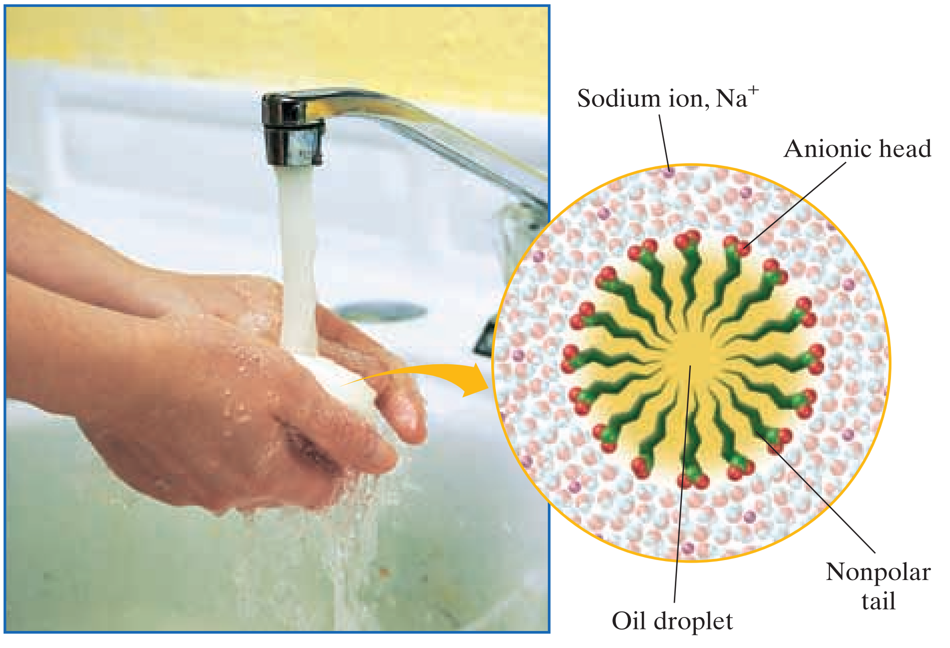


Figure 26
When you wash with soap, you create an emulsion of oil droplets dispersed in water, and stabilized by the soap.

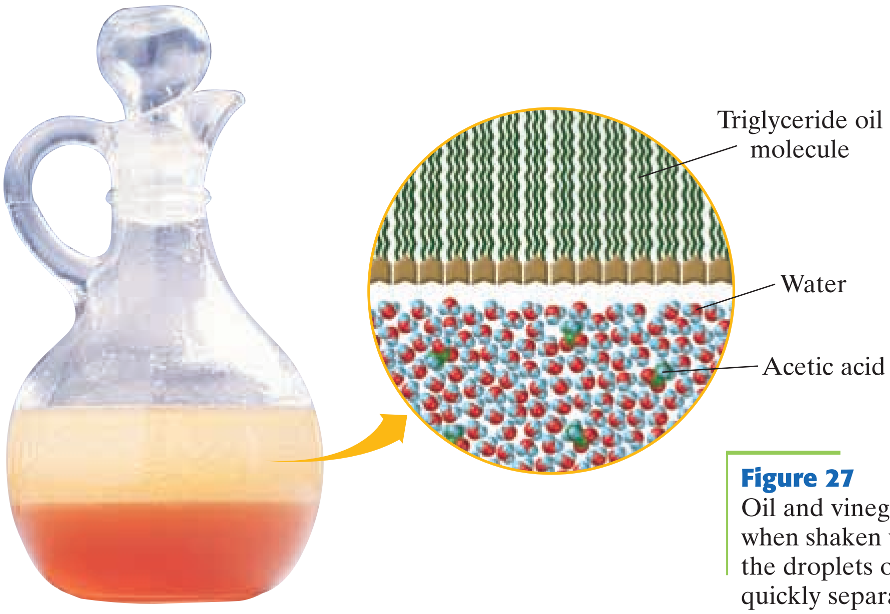


Figure 27
Oil and vinegar appear to mix when shaken vigorously. However, the droplets of oil and vinegar quickly separate into two layers.

Hard Water Limits Soap's Detergent Ability

Soaps are actually salts, although their physical properties are quite different from the salts you have studied so far. Like other salts, when dissolved, soaps form ions. Unlike other salts, the polyatomic anion of soap contains a long nonpolar part. It is this nonpolar hydrocarbon chain that is soluble with oils and dirt.

Soaps are not ideal cleansing agents because the salts of some of their anions are insoluble in water, especially salts of calcium, magnesium, and iron(II). Hard water has high concentrations of these cations, which react with anions such as the palmitate anion to form insoluble salts, such as the one shown in the following equation.

$$2C_{12}H_{25}COO^{-}(aq) + Ca^{2+}(aq) \longrightarrow (C_{12}H_{25}COO)_2Ca(s)$$

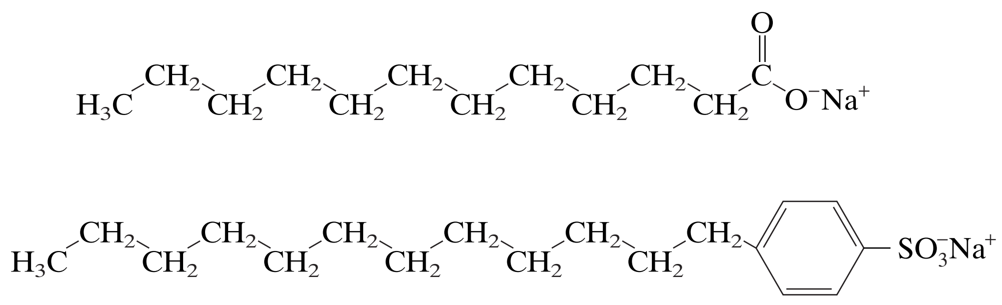
This is the type of substance responsible for bathtub rings. Before synthetic detergents were introduced for shampoos, some of the scum left over from soap would remain in people's hair after they washed it with soap. People would have to rinse their hair with vinegar to wash out the solid salts left over from the soap.

Synthetic Detergents Outperform Soaps in Hard Water

As noted above, soaps form precipitates when used in hard water. In the 1930s, chemists developed synthetic detergents as a substitute for soap to avoid this problem. Synthetic detergents can be used in hard water without forming precipitates. Today, almost all laundry products and shampoos contain synthetic detergents.

The early synthetic detergents were not biodegradable, gradually collected in the groundwater, and caused a serious pollution problem in some regions. Streams going over waterfalls developed huge mounds of soap suds. Most synthetic detergents are now made to biodegrade when they are disposed of.

Figure 28
The structure of sodium laurate (top), a typical soap, is shown here. Sodium dodecylbenzene sulfonate (bottom), is a synthetic detergent. It does not form an insoluble precipitate in hard water.



The basic structure of a synthetic detergent is the same as the structure of a soap. However, the long nonpolar tail of the detergent is connected to the salt of a sulfonic acid, –SOOOH, instead of the organic acid, –COOH. The difference between the structures of a synthetic detergent and soap is shown in **Figure 28.** The hexagon represents a six-carbon benzene ring, which is also nonpolar. The sulfonate group is negatively charged, just as the anionic group on a typical soap is. However, the sulfonate group does not form precipitates with magnesium, calcium, and iron(II) ions, which are found in hard water. The carboxylate anion, $–COO^-$, reacts with hard water ions to form salts that come out of solution, but the sulfonate anion, $–SOOO^-$, does not.

4 Section Review

UNDERSTANDING KEY IDEAS

1. What carries an electric current through a solution?

2. Is sugar an electrolyte? Why or why not?

3. How is a *weak electrolyte* different from a *strong electrolyte*?

4. Why does tap water conduct electricity, whereas distilled water does not?

5. What effect does a solute have on the boiling point of a solvent?

6. Why does spreading salt on an icy sidewalk cause the ice to melt?

7. What is the difference between the meaning of the terms *detergent* and *soap*?

8. What is hard water?

9. What is an emulsion?

10. What is an emulsifying agent?

CRITICAL THINKING

11. Will 1 mol of sugar have the same effect as 1 mol of table salt in lowering the freezing point of water? Explain.

12. Suppose you were taking a bath in distilled water but were using soap. Should you still worry about electric shock?

13. Are soap and synthetic detergents equally good as emulsifiers? Explain your answer.

14. A water softener removes calcium and magnesium ions from water. Why does a softening agent improve the cleansing ability of soap?

15. Why is soap described as a detergent? Why is it described as a surfactant?

CHAPTER HIGHLIGHTS 13

KEY TERMS / KEY IDEAS

SECTION ONE What Is a Solution?

Key terms: solution, suspension, solvent, solute, colloid

- A solution is a homogeneous mixture of a solute dissolved in a solvent.
- Several methods can be used to separate the components in a mixture.

SECTION TWO Concentration and Molarity

Key terms: concentration, molarity

- Units of concentration express the ratio of solute to solution, or solute to solvent, that is present throughout a solution.
- Molarity is moles of solute per liter of solution.

SECTION THREE Solubility and the Dissolving Process

Key terms: solubility, miscible, immiscible, dissociation, hydration, saturated solution, unsaturated solution, supersaturated solution, solubility equilibrium, Henry's law

- Whether substances dissolve in each other depends on their chemical nature, on temperature, and on their ability to form hydrogen bonds.
- In general, polar dissolves in polar, and nonpolar dissolves in nonpolar.
- Ionic solubility can be roughly predicted using a table of ionic solubilities.
- A saturated solution has the solute in equilibrium with excess solute. A supersaturated solution has more dissolved solute than the equilibrium amount.
- Pressure and temperature affect the solubility of gases.

SECTION FOUR Physical Properties of Solutions

Key terms: conductivity, electrolyte, nonelectrolyte, hydronium ion, colligative property, surfactant, detergent, soap, emulsion

- Ions are mobile in solution, so ionic solutions conduct electricity.
- Colligative properties involve the number of solute particles in solution.
- Surfactants make oil and water miscible.

KEY SKILLS

Calculating Parts per Million
Sample Problem A p. 461

Preparing 1.000 L of a 0.5000 M Solution
Skills Toolkit 1 p. 463

Calculating Molarity
Skills Toolkit 2 p. 464
Sample Problem B p. 465

Solution Stoichiometry
Sample Problem C p. 466

13 CHAPTER REVIEW

USING KEY TERMS

1. What happens to a suspension when it is allowed to stand over a period of time?
2. If sugar is dissolved in water, which component is the solute, and which component is the solvent?
3. Explain why milk is a colloid.
4. What ratio does molarity express?
5. Of the following three substances, which two are miscible with one another: oil, water, and ethanol?
6. One solution is made by dissolving sucrose in water. Another solution is made by dissolving NaCl in water. Which of these dissolving processes involves dissociation?
7. What mass of ammonium chloride can be added to 100 g of water at 20°C before the solution becomes saturated? (See **Figure 12.**)
8. If 20 g KCl is dissolved in 100 g of water at 20°C, is the solution unsaturated, saturated, or supersaturated? (See **Figure 12.**)
9. WRITING SKILLS Write a paragraph explaining what happens to an ionic salt in the following steps: it is dissolved in water, more of it than its solubility amount is added to the solution, the solution is heated, the solution is cooled to room temperature, and the solution is disturbed by adding more solute.
10. State Henry's law in your own words.
11. Give an example of a nonelectrolyte.
12. What kind of mixture is soap able to form in order to make oil and water soluble?
13. Name two colligative properties.
14. What is the formula of a hydronium ion?

UNDERSTANDING KEY IDEAS

Solutions, Suspensions, and Colloids

15. What is a solution? How does it differ from a colloid?
16. What are the two components of a solution, and how do they relate to each other?
17. Explain how distillation can be used to obtain drinking water from sea water.
18. Explain how paper chromatography separates the components in a solution.
19. List these mixtures in order of increasing particle size: muddy water, sugar water, sand in water, and milk.
20. A few drops of milk are added to a glass of water, producing a cloudy mixture. The water is still cloudy after standing in the refrigerator for a week. What is this mixture called?

Concentration and Molarity

21. Name a unit of concentration commonly used to express small concentrations.
22. State the following expression in words: $[K_3PO_4]$
23. A solution of NaCl is 1 M. Why is the concentration of particles 2 M?

24. Describe how you would prepare 250.0 mL of a 0.500 M solution of NaCl by using apparatus found in a chemistry lab.

Solubility and the Dissolving Process

25. Explain why vitamin C is soluble in water.

26. Explain why gasoline is insoluble in water.

27. Why do small solid crystals dissolve in liquid more quickly than large crystals?

28. Would ammonium chloride be considered soluble in water?

29. Would the compound $BaSO_4$ be considered soluble in water?

30. Would the compound K_2O be considered soluble in water?

31. Why does warmer liquid dissolve less gas than colder liquid?

Physical Properties of Solutions

32. A solution of salt in water conducts electricity, but a solution of sugar does not. Explain why.

33. A 1 M solution of NaCl in water has a freezing point that is 3.7°C lower than pure water. Estimate what the freezing point would be for a 1 M solution of $CaCl_2$.

34. Explain why soap is a surfactant, a detergent, and an emulsifying agent.

35. Explain why acetic acid is considered a weak electrolyte and why HCl is considered a strong electrolyte.

36. Draw a diagram of an oil droplet suspended in soapy water.

PRACTICE PROBLEMS

PROBLEM SOLVING SKILL

Sample Problem A **Calculating Parts per Million**

37. A saturated solution of $PbCO_3$ contains 0.00011 g $PbCO_3$ in 100 g of water. What is this concentration in parts per million?

38. Community water supplies usually contain 1.0 ppm of sodium fluoride. A particular water supply contains 0.0016 g of NaF in 1.60 L of water. Does it have enough NaF?

39. Most community water supplies have 0.5 ppm of chlorine added for purification. What mass of chlorine must be added to 100.0 L of water to achieve this level?

40. A 12.5 kg sample of shark meat contained 22 mg of methyl mercury, CH_3Hg^+. Is this amount within the legal limit of 1.00 ppm of methyl mercury in meat?

Sample Problem B **Calculating Molarity**

41. If 15.55 g NaOH are dissolved in enough water to make a 500.0 mL solution, what is the molarity of the solution?

42. A solution contains 32.7 g H_3PO_4 in 455 mL of solution. Calculate its molarity.

43. How many moles of $AgNO_3$ are needed to prepare 0.50 L of a 4.0 M solution?

44. What is the molarity of a solution that contains 20.0 g NaOH in 2.00 L of solution?

45. Calculate the molarity of a H_3PO_4 solution of 6.66 g in 555 mL of solution.

46. Calculate the mass of NaOH in 65.0 mL of 2.25 M solution.

47. What mass of HCl is contained in 645 mL of 0.266 M solution?

48. What is the molarity of a hydrochloric acid solution that contains 18.3 g HCl in 100.0 mL of solution?

49. A saturated solution of NaCl contains 36 g NaCl in 114 mL of solution. What is the molarity of the solution?

50. Calculate the mass of LiF in 100.0 mL of 0.100 M solution.

51. How many grams of glucose, $C_6H_{12}O_6$, are in 255 mL of a 3.55 M solution?

Sample Problem C Solution Stoichiometry

52. You mix 1.00 L of 2.00 M $BaCl_2$ with 1.00 L of 2.00 M $AgNO_3$. What compounds remain in solution, and what are their concentrations?

$$BaCl_2(aq) + 2AgNO_3(aq) \longrightarrow 2AgCl(s) + Ba(NO_3)_2(aq)$$

53. How many milliliters of 18.0 M H_2SO_4 are required to react with 250 mL of 2.50 M $Al(OH)_3$ if the products are aluminum sulfate and water?

54. If 75.0 mL of an $AgNO_3$ solution reacts with enough Cu to produce 0.250 g Ag by single displacement, what is the molarity of the initial $AgNO_3$ solution if $Cu(NO_3)_2$ is the other product?

MIXED REVIEW

55. How many milliliters of 1.0 M $AgNO_3$ are needed to provide 168.88 g of pure $AgNO_3$?

56. What is the mass of potassium chromate, K_2CrO_4, in 20.0 mL of 6.0 M solution?

57. Sodium ions in blood serum normally are 0.145 M. How many grams of sodium ions are in 10.0 mL of serum?

58. A package of compounds used to achieve rehydration in sick patients contains 20.0 g of glucose, $C_6H_{12}O_6$. When this material is diluted to 1.00 L, what is the molarity of glucose?

CRITICAL THINKING

59. Calcium phosphate, $Ca_3(PO_4)_2$, is quite cheap and causes few pollution problems. Why is it not used to de-ice sidewalks? (Hint: See **Table 2** in Section 3.)

60. A calculation shows that a salt will have a negative ΔH and a positive ΔS when it dissolves. Is it actually soluble?

61. Imagine you are a sailor who must wash in sea water. Which is better to use, soap or synthetic detergent? Why?

62. Air pressure in an airplane cabin while in flight is significantly lower than at sea level. Explain in terms of Henry's law how this affects the speed at which a carbonated beverage, after opening, loses its fizz.

63. Why would a substance that contains only ionic bonds not work as an emulsifying agent?

ALTERNATIVE ASSESSMENT

64. Design a solubility experiment that would identify an unknown substance that is either CsCl, RbCl, LiCl, NH_4Cl, KCl, or NaCl. (Hint: You will need a solubility versus temperature graph for each of the salts.) If your instructor approves your design, get a sample from the instructor, and perform your experiment.

65. Many reagent chemicals used in the lab are sold in the form of concentrated aqueous solutions, as shown in the table below. Different volumes are diluted to 1.00 L to make less-concentrated solutions. Create a computer spreadsheet that will calculate the volume of concentrated reagent needed to make 1.00 L solutions of any molar concentration that you enter.

Reagent	Concentration (M)
H_2SO_4	18.0
HCl	12.1
HNO_3	16.0
H_3PO_4	14.8
CH_3COOH	17.4
NH_3	15.0

CONCEPT MAPPING

66. Use the following terms to create a concept map: *concentration, dissociates, electrical conductivity, solute,* and *solvent.*

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."

67. What do the numbers on the *y*-axis represent?

68. What does each curve on the graph represent?

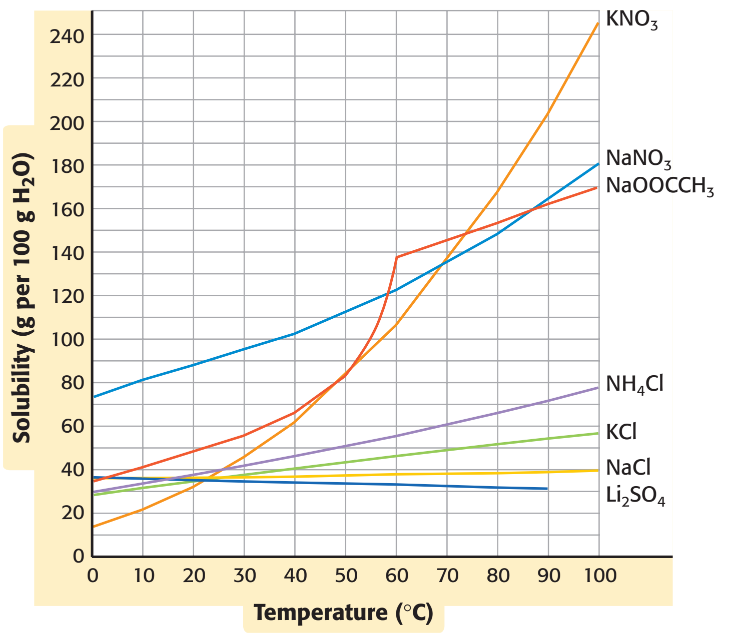
69. Are most of the substances represented on the graph more or less soluble at higher temperatures?

70. Which salt is most soluble at 10°C? at 60°C? at 80°C?

71. If you heat water to 80°C, what amount of NaCl could you dissolve in it as compared to water that is at 20°C?

72. Which salt's solubility is most strongly affected by changes in temperature?

Solubility Vs. Temperature for Some Solid Solutes



TECHNOLOGY AND LEARNING

73. Graphing Calculator

Predicting Solubility from Tabular Data

The graphing calculator can run a program that graphs solubility data. Given solubility measurements for KCl, you will use the data to predict its solubility at various temperatures.

Go to Appendix C. If you are using a TI-83 Plus, you can download the program and data sets and run the application as directed. Press the APPS key on your calculator, then choose the application CHEMAPPS. Press 3, then highlight ALL on the screen, press 1, then highlight LOAD and press 2 to load the data into your calculator. Press the keys 2nd and then QUIT, and then run the program SOLUBIL. For L_1, press 2nd and LIST, and choose TMP21. For L_2, press 2nd and LIST and choose SOL21.

If you are using another calculator, your teacher will provide you with keystrokes and data sets to use.

a. At what temperature would you expect the solubility to be 48.9 g per 100 g H_2O?

b. At what temperature would you expect the solubility to be 35 g per 100 g H_2O?

c. What would you expect the solubility to be at a temperature of 100°C?

13 STANDARDIZED TEST PREP

UNDERSTANDING CONCEPTS

Directions (1–3): **For *each* question, write on a separate sheet of paper the letter of the correct answer.**

1. Which of the following types of compounds are **most** likely to be strong electrolytes?
   A. covalent
   B. ionic
   C. nonpolar
   D. polar

2. Why is acetic acid a weak electrolyte?
   F. It is miscible with water.
   G. It lowers the freezing point of water.
   H. It ionizes only slightly in aqueous solution.
   I. It forms hydronium and hydroxide ions in aqueous solution.

3. Of the following solutions, which will have the lowest freezing point?
   A. 1 M $CaCl_2$
   B. 1 M $MgSO_4$
   C. 1 M NaCl
   D. 1 M sugar

Directions (4–6): **For *each* question, write a short response.**

4. In some coastal areas where fresh water is not available, potable water is produced by distillation of seawater. How does this process work?

5. How does the rate of dissolution of a solute compare to the rate of its recrystallization in a saturated solution?

6. Explain the effect of pressure on the solubility of a gas in a liquid.

READING SKILLS

Directions (7–9): **Read the passage below. Then answer the questions.**

Many cold water streams are habitats for abundant bacteria and animal life which relies on the oxygen dissolved in the water. As water temperature is raised from just above freezing to 30°C, the solubility of oxygen decreases by about half. At the same time chemical reactions, including those in living cells, proceed faster. Many industrial processes, such as power plants producing waste heat in the form of large volumes of hot water, now use cooling towers to cool water before it is discharged or reused.

7. Why is the discharge of hot water into a stream called thermal pollution even if the water is pure?
   F. Other pollutants become more soluble in warm water.
   G. The increased temperature of the stream is harmful to the environment.
   H. Heating the water increases the amount of oxygen in the stream.
   I. Anything that does not occur naturally, including hot water, is a pollutant.

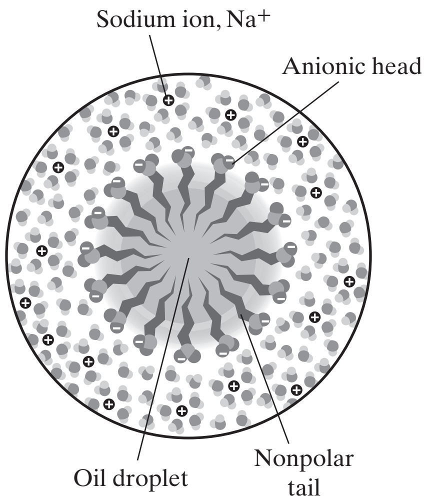
8. Why would an increase in chemical reaction rates worsen thermal pollution?
   A. It decreases the ability of organisms to use oxygen.
   B. It increases the rate of production of chemical pollutants.
   C. It increases the need for oxygen by speeding up reactions inside organisms.
   D. It decreases water flow rate so organisms are unable to use dissolved oxygen.

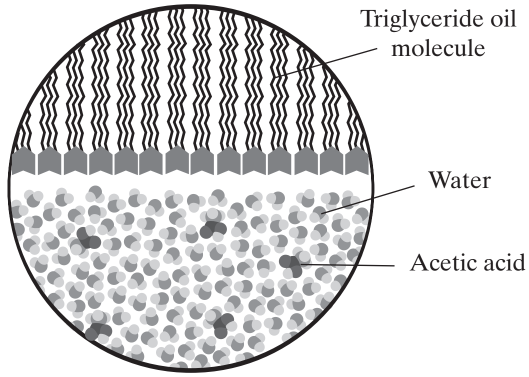
9. In a hot summer with little rainfall, the volume of water flowing in streams is drastically reduced and many fish die. What is the cause of these fish kills?

INTERPRETING GRAPHICS

***Directions (10–13):* For *each* question below, record the correct answer on a separate sheet of paper.**

The diagram below on the left shows what happens when you wash your hands with soap. The diagram below on the right shows what happens when you mix oil and vinegar. Use these diagrams to answer questions 10 through 13.





10 What do you call a mixture of oil, water, and soap as represented by the illustration on the left?

F. a colligative mixture
G. an emulsion
H. a homogenous mixture
I. a solution

11 Why are oil and water immiscible liquids?

A. Both are polar molecules.
B. Both are nonpolar molecules.
C. Oil is nonpolar and water is polar.
D. Water is nonpolar and oil is polar.

12 Why does acetic acid remain in the water layer of the salad dressing?

F. Acetic acid molecules are too large to dissolve in oil.
G. Acetic acid ionizes completely and ions remain in the water layer.
H. Acetic acid molecules are polar and interact with polar water molecules.
I. Acetic acid molecules react with the water molecules to form a new compound.

13 After a vinegar and oil salad dressing is shaken (right illustration), it quickly separates into two layers, unlike the mixture of oil and water with soap. How does soap cause the oil to form droplets instead of layers?

Test TIP

When using an illustration that has labels to answer a question, read the labels carefully, and then check that the answer you choose matches your interpretation of the labels.

CHAPTER

14

Chemical Equilibrium