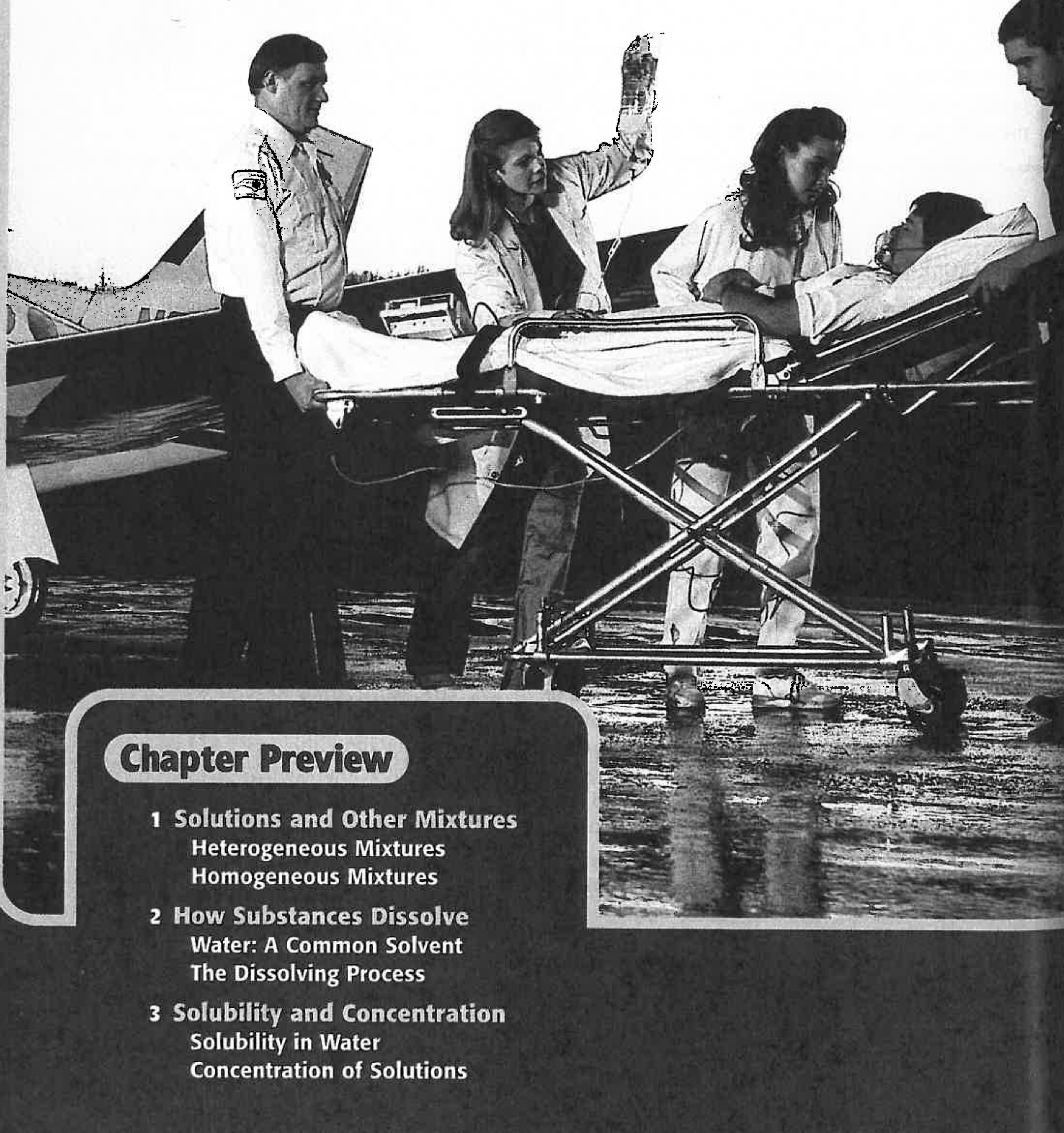
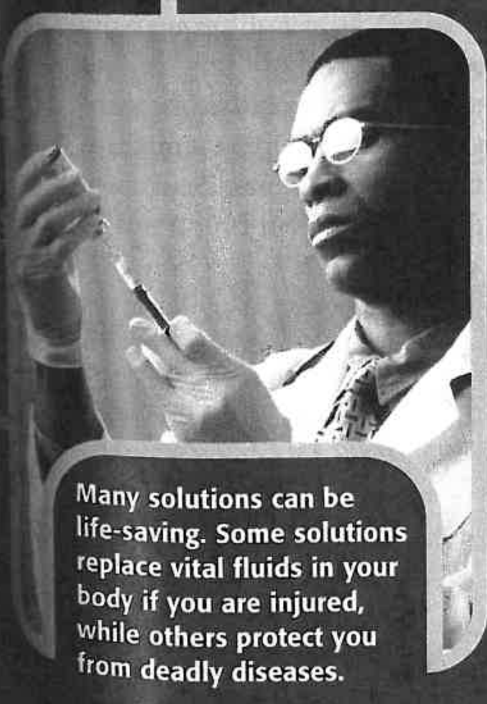
# CHAPTER 7

# Solutions



## Chapter Preview



Many solutions can be life-saving. Some solutions replace vital fluids in your body if you are injured, while others protect you from deadly diseases.

# *Focus* ACTIVITY

**Background** Paramedics rush to the scene of an accident. Someone has been injured, and the person's blood pressure has become dangerously low. Paramedics pump a *saline solution*, a mixture of water and sodium chloride that is similar to blood, into the person's veins. This mixture maintains the blood pressure that is needed to keep the person alive on the way to the hospital.

Shots called *vaccines* are mixtures that help protect you from many diseases. Vaccines have a tiny amount of the disease-causing organism you are trying to protect yourself from. The shot you get is harmless because the organism contained in it is dead, or inactivated. But the shot keeps you from getting the disease because your body can now recognize this harmful bacterium or virus again and fight it.

**Activity 1** Look up the word *saline* in the dictionary. Which group of elements in the periodic table form ionic compounds that can be described by the word *saline*? Explain how the word *saline* applies to sodium chloride.

**Activity 2** Fill a clear plastic cup with water. After the water settles, add table salt one teaspoon at a time to the water. Stir after you add each spoonful until all of the salt dissolves. How much salt are you able to dissolve before it stops dissolving and settles to the bottom of the cup? Perform this activity again, but this time use sugar instead of salt. Does the same amount of sugar dissolve? If not, what might explain the difference?

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**Topic: Vaccines SciLinks code: HK4144**

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## Pre-Reading Questions

1. The label on some drinks reads "Shake well before serving." Why do you need to shake these drinks? Why don't you need to shake all drinks?
2. Frozen orange juice must be mixed with water before you can drink it. Why is frozen orange juice referred to as *orange juice concentrate*?

SECTION 1

# Solutions and Other Mixtures

**KEY TERMS**
**suspension**
**colloid**
**emulsion**
**solution**
**solute**
**solvent**
**alloy**

**OBJECTIVES**

- **Distinguish** between heterogeneous mixtures and homogeneous mixtures.
- **Compare** the properties of suspensions, colloids, and solutions.
- **Give** examples of solutions that contain solids or gases.

Any sample of matter is either a pure substance or a mixture of pure substances. You can easily tell that fruit salad is a mixture because it is a blend of different kinds of fruit. But some mixtures look like they are pure substances. For example, a mixture of salt and water looks the same as pure water. Air is a mixture of several gases, but you cannot see different gases in the air.

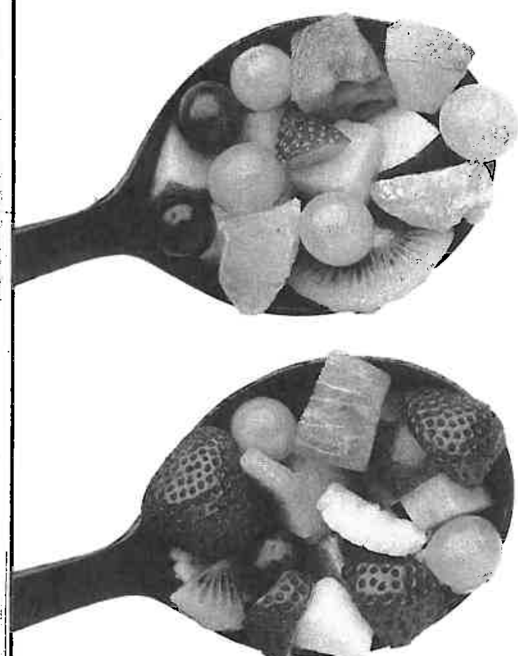


**Figure 1**
Fruit salad is a heterogeneous mixture. Each spoonful has a different composition of fruit because the fruits are not distributed evenly throughout the salad.

## Heterogeneous Mixtures

The amount of each substance in different samples of a *heterogeneous mixture* varies, just as the amount of each kind of fruit varies in each spoonful of fruit salad, as shown in ***Figure 1.*** If you compared two shovels full of dirt from a garden, they would not be exactly the same. Each shovelful would have a different mixture of rock, sand, clay, and decayed matter.

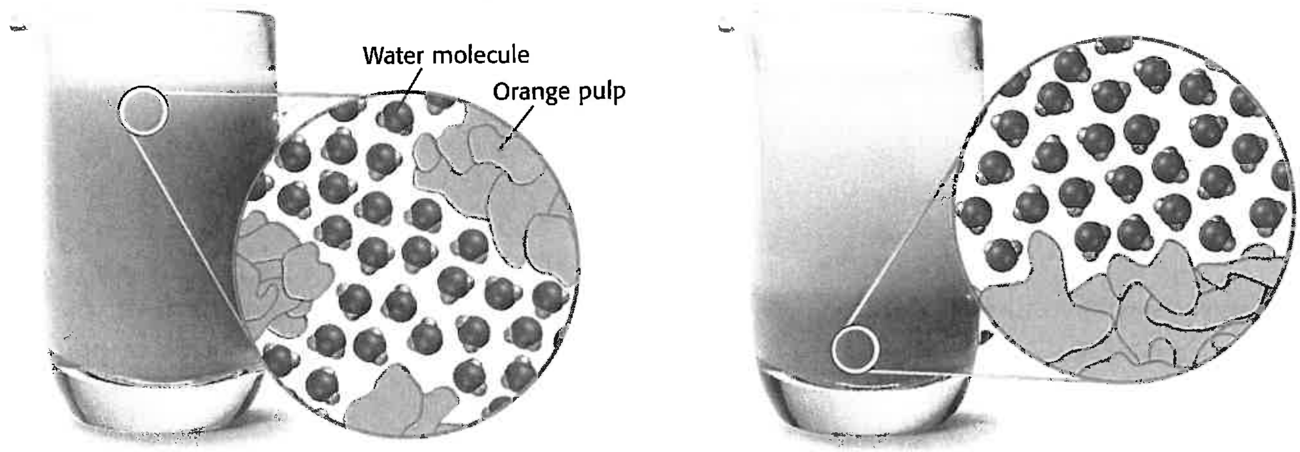
Another naturally occurring heterogeneous mixture is granite, a type of igneous rock shown in ***Figure 2.*** Granite is a mixture of crystals of the minerals quartz, mica, and feldspar. Because a mixture has no fixed composition, samples of granite from different locations can vary greatly in appearance because the samples have different proportions of minerals.

**Figure 2**
These paperweights are made from granite, which is a mixture of quartz, black mica, and feldspar.

***Figure 3*** **Orange Juice: A Heterogeneous Mixture**

**A** The pulp in the orange juice is spread throughout the mixture right after the orange juice is shaken.

**B** Over time, the pulp does not stay mixed with the water molecules. The pulp settles to the bottom, and two layers form.




## Particles in a suspension are large and eventually settle out

Have you ever forgotten to shake the orange juice carton before pouring yourself a glass of juice? The juice probably tasted thin and watery. Natural orange juice is a **suspension** of orange pulp in a clear liquid that is mostly water, as shown in ***Figure 3A.*** A property of a suspension is that the particles settle out when the mixture is allowed to stand. So if the orange juice carton is not shaken, the top layer of the juice in the carton is mostly water because all the pulp has settled to the bottom.

Orange juice is clearly a heterogeneous mixture because after it settles, the liquid near the top of the container is not the same as the liquid near the bottom. Shaking the container mixes the pulp and water, but the pulp pieces are big enough that they will eventually settle out again, as shown in ***Figure 3B.***

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

## Particles in a suspension may be filtered out

Particles in suspensions are usually the size of or larger than the tip of an extremely sharp pencil, which has a diameter of about 1000 nm, or about the size of a bacterial cell. Particles of this size are large enough that they can be filtered out of the mixture. For example, a filter made of porous paper can be used to catch the suspended pulp in orange juice. That is, the pulp stays in the filter, while water molecules pass through the filter easily. You can classify a mixture as a suspension if the particles settle out or can be filtered out.

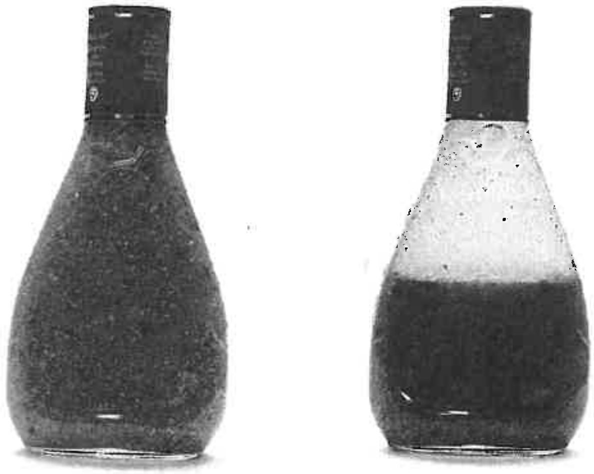


*Figure 4*

Some salad dressings are made with oil and vinegar, which form a suspension when shaken. Oil and vinegar mixtures separate after standing for a few minutes.

## Some mixtures of two liquids will separate

Oil, vinegar, and flavorings can be shaken together to make salad dressing. But when the dressing stands for a few minutes, two layers form, as shown in ***Figure 4.*** The two liquids separate because they are *immiscible,* which means they do not mix. Eventually, the oil, which is less dense, rises and floats on the vinegar, which is denser.

One way to separate two immiscible liquids is to carefully pour the less dense liquid off the top. Some cooks use this technique to separate melted fat from meat juices. The cook removes the fat by pouring or spooning it off the meat juices, which are denser than the fat. The process of pouring a lighter liquid off of a heavier liquid is called *decanting.*

**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

## Particles in a colloid are too small to settle out

Latex paint is an example of another kind of heterogeneous mixture called a **colloid.** Latex paint is a thick combination of solid particles of pigment dispersed in water and other substances that make the pigment stick to a surface. The difference between colloids and suspensions is that the particles in colloids are smaller than those in suspensions—ranging from only 1 to 100 nm in diameter. Because the particles in colloids are so small, they pass through ordinary filters and stay dispersed throughout the mixture. However, the particles are large enough to scatter light that passes through the colloid, even though the colloid may look like clear water, as shown in ***Figure 5.*** This scattering of light is called the *Tyndall effect.*

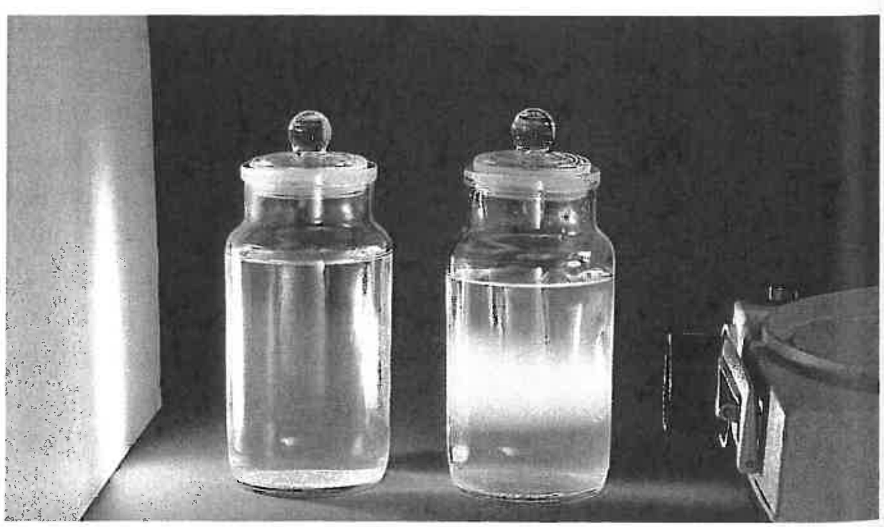


*Figure 5*

The liquid in the jar on the right is a colloid. Colloids exhibit the Tyndall effect, in which light is scattered by the invisible particles.

## Quick ACTIVITY

### Making Butter

Cream is a lipid-in-water emulsion. Churning or shaking cream causes lipid droplets to stick to one another, which forms butter. You can make your own butter in the following way.

1. Pour 250 mL (about 1/2 pint) of heavy cream into an empty 500 mL container.
2. Add a clean marble, and seal the container tightly so that it does not leak.
3. Take turns shaking the container. When the cream becomes very thick, you will no longer hear the marble moving.
4. Open the container to look at the substance that formed. Record your observations.
5. If the butter is made of joined lipid droplets, what must make up most of the liquid that is left behind?

### Other familiar materials are also colloids

The particles in most colloids are composed of many atoms, ions, or molecules, but individual protein molecules are also large enough to form colloids. Examples of protein colloids include gelatin desserts, egg whites, and blood plasma. These colloids consist of protein molecules dispersed in a liquid.

Other examples of colloids include whipped cream, which is made by dispersing a gas in a liquid, and marshmallows, which are made by dispersing a gas in a solid. Fog consists of small droplets of water dispersed in air, and smoke contains small solid particles dispersed in air.

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

### Some immiscible liquids can form colloids

Mayonnaise is a colloid consisting of very small droplets of oil suspended in vinegar. Vinegar-and-oil salad dressings separate into two layers, but the vinegar and oil in mayonnaise do not separate. Mayonnaise has another ingredient that keeps the oil and vinegar together—egg yolk. Egg yolk coats the oil droplets, which keeps them from joining to form a separate layer. Mayonnaise is an **emulsion,** a colloid in which liquids that normally do not mix are spread throughout each other. Another example of an emulsion is found in your body. Bile salts cause fats to form an emulsion in the small intestine. Then digestive enzymes are able to break down the smaller fat particles.

Like other colloids, an emulsion has particles so small that it may appear to be uniform. But a closer look shows that it is not. For example, cream does not form separate layers, so it looks like a pure substance. Cream is really a mixture of oily fats, proteins, and carbohydrates dispersed in water. The lipid droplets are coated with a protein. The protein is an *emulsifier* that keeps the lipid droplets dispersed in the water so that they can spread throughout the entire mixture, as shown in ***Figure 6.***

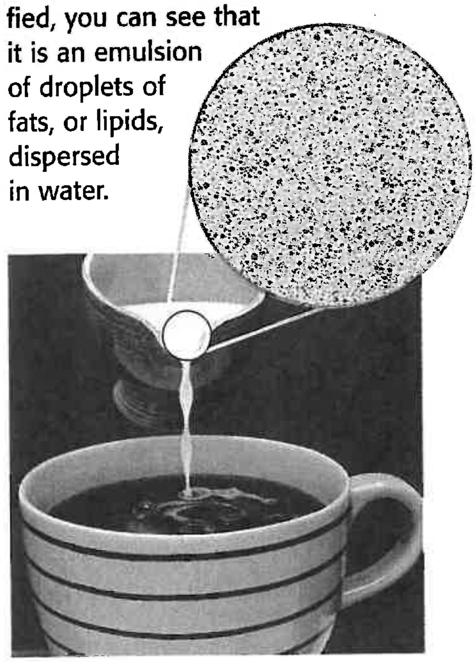


***Figure 6***
Cream may look like a pure substance, but when cream is magnified, you can see that it is an emulsion of droplets of fats, or lipids, dispersed in water.

## Connection to ENGINEERING

Ink is a complicated mixture of substances. Some inks, such as those used in printing books and magazines, contain *pigments* that give the ink most of its color. Pigments are added to a liquid as finely ground, solid particles to form a suspension. The ink is then applied to the paper on a printing press and allowed to dry.

The ink used in some ballpoint pens is different from ink used in printing. The ink in pens contains a dissolved iron salt, such as ferrous sulfate, and an organic substance called *tannic acid.* Tannic acid and the iron salt mix to form a dark blue solution that gives the ink its blue-black color.

### Making the Connection

1. In a library or on the Internet, research two different types of inks used in printing. How do the properties of the inks differ? What properties make them good choices for printing different materials?

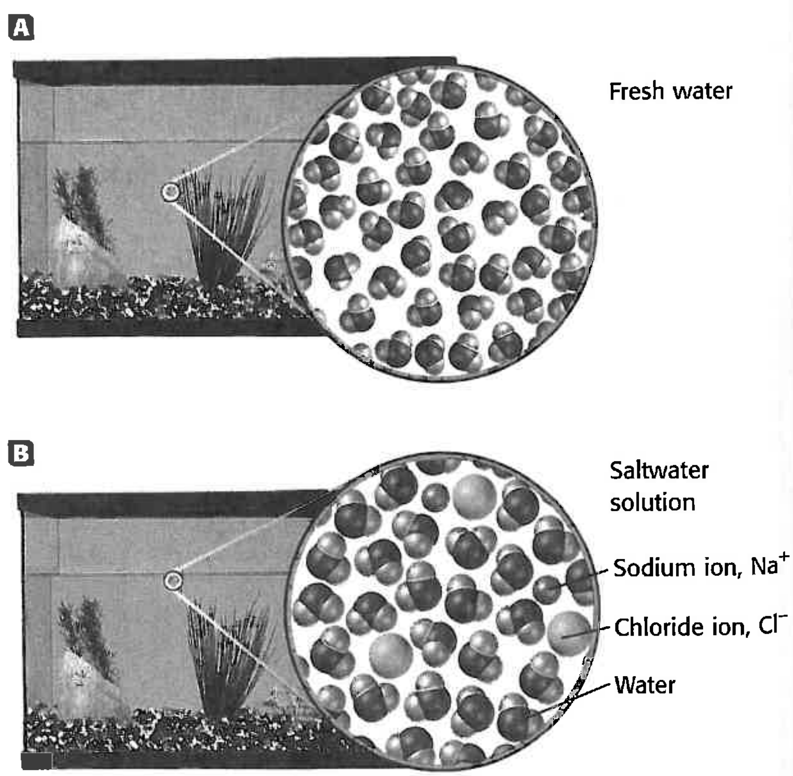
## Homogeneous Mixtures

Homogeneous mixtures not only look uniform but *are* uniform. Salt water is an example of a homogeneous mixture. If you add pure salt to a glass of water and stir, the mixture soon looks like pure water. The mixture looks uniform even when you examine it under a microscope. That's because the individual components of the mixture are too small to be seen. The mixture is made of sodium ions and chloride ions surrounded by water molecules, as shown in ***Figure 7.*** Because the number of ions will be the same throughout the salt water, the mixture is homogeneous.

When salt and water are mixed, no chemical reaction occurs. For this reason, the two substances can be easily separated by evaporating the water. As the water evaporates, the sodium and chloride ions of the salt begin to rejoin to form salt crystals like those that originally dissolved. When all the water evaporates, only salt is left.

***Figure 7***

**A** Plain water is homogeneous because it is a single substance. **B** Salt water is also homogeneous because it is a uniform mixture of water molecules, sodium ions, and chloride ions.



## Homogeneous mixtures are solutions

When you add aquarium salt to water and stir, the solid seems to disappear. What really happens is that salt has *dissolved* in water to form a **solution.** In this particular solution, aquarium salt is the **solute,** the substance that dissolves. Water is the **solvent,** the substance in which the solute dissolves. When a solute has dissolved in a solvent, the dissolved particles are so small that you can't see them, even with a powerful microscope. They are invisible because they are the smallest particles of the substance—atoms, ions, or molecules. Mixtures are homogeneous when the smallest particles of one substance are uniformly spread among similar particles of another substance. This description of a homogeneous mixture is also a description of a solution, so all homogeneous mixtures are also solutions.

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

**solute** in a solution, the substance that dissolves in the solvent

**solvent** in a solution, the substance in which the solute dissolves

## Miscible liquids mix to form solutions

Two or more liquids that form a single layer when mixed are said to be miscible. Some solutions of two liquids are useful. For example, water mixed with isopropanol makes a solution called *rubbing alcohol,* and some skin lotions contain a solution of glycerol and water.

Chemists often have to separate miscible liquids when purifying substances in the laboratory. Because miscible liquids do not separate into layers, they are not as easy to separate as immiscible liquids are. One way to separate miscible liquids is by a process called *distillation,* which involves an apparatus such as the one shown in ***Figure 8.*** Distillation works only when the two miscible liquids have different boiling points. For example, a mixture of methanol and water can be separated by distillation because methanol boils at 64.5°C and water boils at 100.0°C. When this mixture is heated in a distillation apparatus, the methanol boils away first, leaving most of the water behind.

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**Solutions**
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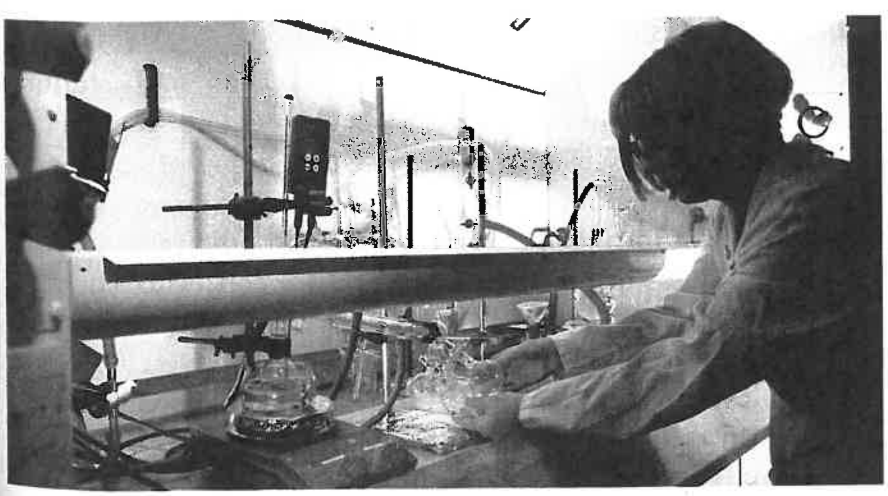


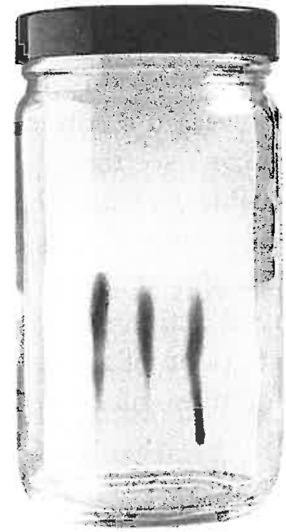
***Figure 8***
A solution of miscible liquids that have different boiling points may be separated by distillation in an apparatus such as this one.

REAL WORLD APPLICATIONS

**Chromatography** Chromatography is often used to separate mixtures that can't be separated by simpler methods. The figure at right shows how paper chromatography can be used to separate colored dyes in three different samples of black ink.

First, ink marks are made on absorbent paper. Then the paper is put in a jar holding a small volume of solvent. The solvent travels upward through the paper, carrying the ink with it. The finished *chromatogram* reveals which dyes make up each of the inks.

Each dye has a different chemical structure. Dyes that have structures more like that of the paper than that of the solvent stick to the paper and travel slower. Dyes that have structures more like that of the solvent move upward with the solvent and therefore travel farther.



**Applying Information**

1. Does the blue dye in each sample have a structure more like that of the paper or the solvent? Explain your answer.
2. How would the result differ if the inks were made from a single dye instead of a mixture of several dyes?

## Liquid solutions sometimes contain no water

Many examples of solutions of liquids in other liquids do not contain water. For example, some kinds of fingernail-polish removers and paint strippers are mixtures of liquid substances that contain no water. Fuels such as gasoline, diesel fuel, and kerosene are homogeneous mixtures of several different liquid substances. These fuels are made from a liquid solution called petroleum, which is also called *crude oil*. Petroleum is a solution of many different carbon compounds. Fuels and other useful materials are made from petroleum by distillation. For example, gasoline is a solution containing several liquid substances distilled from petroleum.

## Other states of matter can also form solutions

Like the water in a saltwater aquarium, many common solutions are solids dissolved in liquids. However, solutes and solvents can be in any state. For instance, vinegar is a solution of acetic acid dissolved in water, both of which are liquids. The air you breathe is a solution of nitrogen, oxygen, argon, and other gases. Gases can also dissolve in liquids. For example, a soft drink contains carbon dioxide gas dissolved in liquid water. Mothballs, which are made of a solid substance called *naphthalene*, slowly give off vapor that forms a solution with air. The liquid element mercury dissolves in solid silver to form a solution called an *amalgam*, which can be used to fill cavities in teeth.

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## Solids can dissolve in other solids

The musical instrument shown in ***Figure 9*** is made of brass, a solution of zinc metal dissolved in copper metal. Brass is an example of an **alloy,** a homogeneous mixture that is usually composed of two or more metals. Of course, the metals must be melted to liquids and then mixed, but when the mixture cools, the result is a solid solution of one metal in another metal.

Alloys are important because they have properties that the individual metals do not have. For example, pure copper is too soft and bends too easily to be used to make a sturdy musical instrument. When zinc is dissolved in copper, the resulting brass is harder and tougher than copper, but the brass is still easy to form into complicated shapes. Bronze, an alloy of tin in copper, is harder than tin or copper alone. Bronze resists corrosion, so bronze has been used since ancient times to make sculptures and other objects meant to last a long time. Not all alloys contain only metals. Some types of steel are alloys containing the nonmetal element carbon.

**alloy** a solid or liquid mixture of two or more metals

***Figure 9***

This cornet is made of the alloy brass, which is a solid solution.

## SECTION 1 REVIEW

### SUMMARY

- A heterogeneous mixture is a nonuniform blend of two or more substances.
- The particles in a suspension settle out of the mixture or may be filtered out.
- The dispersed particles in a colloid are too small to settle out or to be filtered out.
- A homogeneous mixture, or solution, is a uniform blend of substances.
- In a solution, the solute is dissolved in the solvent.
- Solutions may be formed from solids, liquids, or gases.

1. **Classify** the following mixtures as heterogeneous or homogeneous:
   **a.** orange juice without pulp
   **b.** sweat
   **c.** cinnamon sugar
   **d.** concrete
2. **Explain** the difference between a suspension and a colloid.
3. **List** three examples of solutions that are not liquids.
4. **Identify** the solvent and solute in a solution made by dissolving a small quantity of baking soda in water.
5. **Arrange** the following mixtures in order of increasing particle size: muddy water, sugar water, and egg white.
6. **Explain** why distillation would not be an effective way to separate a mixture of the miscible liquids formic acid, which boils at 100.7°C, and water, which boils at 100.0°C.
7. **Critical Thinking** A small child watches you as you stir a spoonful of sugar into a glass of clear lime-flavored drink. The child says she believes that the sugar went away because it seemed to disappear. How would you explain to the child what happened to the sugar, and how could you show her that you can get the sugar back?

SECTION 2

# How Substances Dissolve

**KEY TERMS**

**polar compound**
**hydrogen bonding**
**nonpolar compound**

**OBJECTIVES**

- **Explain** how the polarity of water enables it to dissolve many different substances.
- **Relate** the ability of a solvent to dissolve a solute to the relative strengths of forces between molecules.
- **Describe** three ways to increase the rate at which a solute dissolves in a solvent.
- **Explain** how a solute affects the freezing point and boiling point of a solution.

Suppose you and a friend are drinking iced tea. You add one spoonful of loose sugar to your glass of tea, stir, and all of the sugar dissolves quickly. Your friend adds a sugar cube to her tea and finds that she must stir longer than you did to dissolve all of the sugar cube. Why does the sugar cube take longer to dissolve? Why does sugar dissolve in water at all?

**polar compound** a molecule that has an uneven distribution of electrons

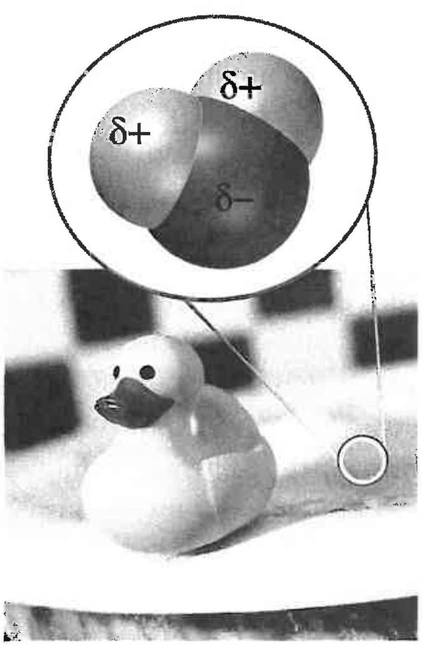
## Water: A Common Solvent

Two-thirds of Earth's surface is water. The liquids you drink are mostly water, and three-fourths of your body weight is water. Many different substances can dissolve in water. For this reason, water is sometimes called the *universal solvent*.

### Water can dissolve ionic compounds because of its structure

To understand what makes water such a good solvent, consider the structure of water. A water molecule is made up of two hydrogen atoms bonded to one oxygen atom by covalent bonds. But electrons are not evenly distributed throughout a water molecule because oxygen atoms strongly attract electrons. The oxygen atom pulls electrons away from the hydrogen atoms, giving them a partial positive charge. The electrons are then bunched around the oxygen atom, giving it a partial negative charge. This uneven distribution of electrons, combined with a water molecule's bent shape, causes it to be a **polar compound,** as shown in ***Figure 10.*** A polar molecule has distinct positively and negatively charged sides, which are indicated by $\delta+$ and $\delta-$ in ***Figure 10.*** Water molecules attract both the positive and negative ions of an ionic compound.

***Figure 10***
Water is a polar molecule because the oxygen atom strongly attracts electrons, which leaves the hydrogen atoms with a positive charge.



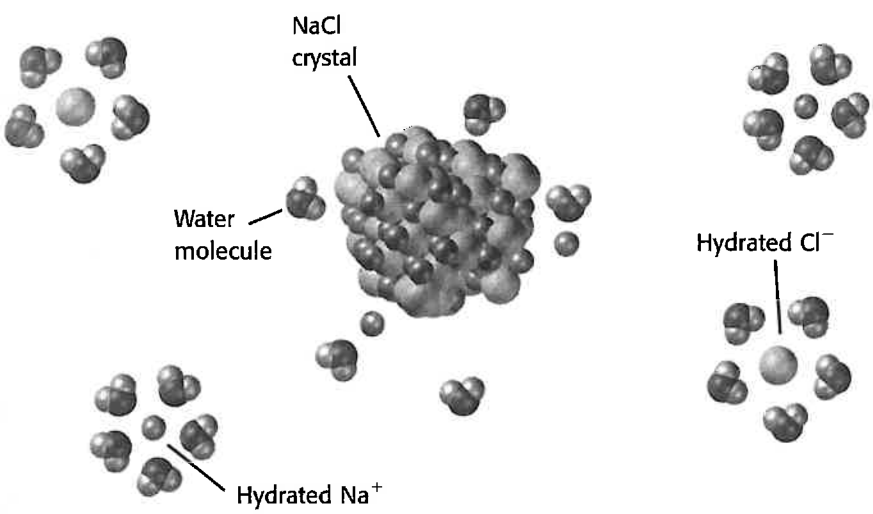



***Figure 11***
Polar water molecules attract the positive sodium ions and negative chloride ions, pulling the ions away from the crystal of NaCl. Notice that the negative oxygen atoms of the water molecules are attracted to the positive sodium ion and that the positive hydrogen atoms are attracted to the negative chloride ions.

### Polar water molecules pull ionic crystals apart

***Figure 11*** shows how a sodium chloride, NaCl, crystal dissolves in water. First, the partially negative oxygen atoms of water molecules attract the positively charged sodium ions at the surface of the sodium chloride crystal. The partially positive hydrogen atoms of water molecules also attract the negatively charged chloride ions. As more water molecules attract the ions, the force of attraction between the ions and the water molecules increases. Finally, this force becomes stronger than the force of attraction between the sodium and chloride atoms in the crystal. Then, the ions are pulled away from the crystal and surrounded by water molecules. Eventually, all of the ions in the crystal are pulled into solution, and the substance is completely dissolved.

### Dissolving depends on forces between particles

Water dissolves baking soda and many other ionic compounds in exactly the same way that water dissolves sodium chloride. The attraction of water molecules pulls the crystals apart into individual ions. Many other ionic compounds, though, do not dissolve in water. Silver chloride, unlike sodium chloride, is an ionic compound that does not dissolve in water.

Why does one ionic compound dissolve in water, but another ionic compound does not? The answer is related to forces of attraction. To dissolve an ionic substance, water molecules must exert a force on the ions that is more attractive than the force between the ions in the crystal. This principle applies not only to water molecules and ionic compounds but also to any solvent and any substance. To dissolve a substance, solvent molecules must exert more force on the particles of the substance than the particles exert on one another.

**INTEGRATING**

**BIOLOGY**

When a solute dissolves in water, the random movement of particles called *diffusion* ensures that the solute spreads out evenly through the solution. Cells rely on diffusion to transport molecules. When the concentration of a solute is greater inside a cell than it is outside, the solute moves out of the cell through the cell membrane.

Not all substances can diffuse across a cell membrane. Sodium ions and potassium ions, for example, are transported into and out of cells through structures called *sodium-potassium pumps.* These ions give many cells an electrical charge that makes it possible for the cells to send electrical signals to different parts of the body.

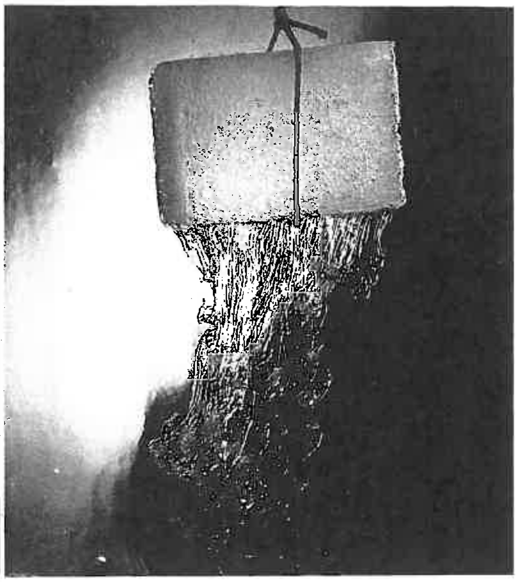


*Figure 12*
As water slowly dissolves the sugar cube, streams of denser sugar solution move downward. Table sugar, or sucrose, is a molecular compound.

## Water dissolves many molecular compounds

Water has a low molecular mass, but it is a fairly dense liquid that has a high boiling point. Water has these properties because **hydrogen bonding** occurs between water molecules. Recall that electrons are pulled away from the hydrogen atoms of water by the oxygen atom. As a result, the hydrogen atom of one water molecule attracts electrons from the oxygen atom of another water molecule to form a partial covalent bond. This hydrogen bond pulls water molecules close together.

Besides ionic compounds, water also dissolves many molecular compounds, such as ethanol, ascorbic acid (vitamin C), and table sugar, as shown in ***Figure 12.*** These molecular compounds are polar because they have hydrogen atoms bonded to oxygen, which attract electrons strongly. Ethanol, for example, has a polar $-OH$ group in its structure, $CH_3CH_2OH$. As a result, the negative oxygen atom of a water molecule attracts the positive hydrogen atom of an ethanol molecule. The positive hydrogen atom of a water molecule also attracts the negative oxygen atom of an ethanol molecule. This attraction is one force that pulls an ethanol molecule into water solution.

Hydrogen bonding plays a large part in the dissolving of other molecular compounds such as sucrose, $C_{12}H_{22}O_{11}$ as shown in ***Figure 13.*** Water molecules form hydrogen bonds with the $-OH$ groups in the sucrose molecule and easily pull the sucrose molecules away from the sugar crystal and into solution.

**hydrogen bonding** the intermolecular force occurring when a hydrogen atom that is bonded to a highly electronegative atom of one molecule is attracted to two unshared electrons of another molecule

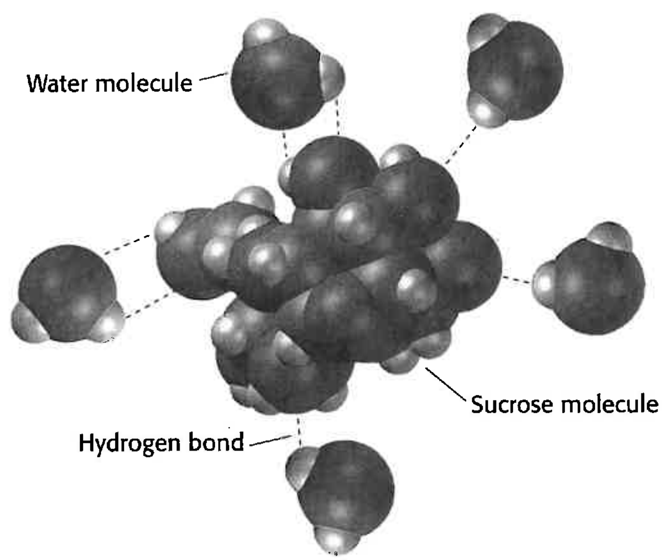



*Figure 13*
Water molecules form hydrogen bonds with the $-OH$ groups of a sucrose molecule. These forces pull the sucrose molecule away from the sugar crystal and into solution.

**Quick ACTIVITY**

## What dissolves a nonpolar substance?

Some elements, such as oxygen, nitrogen, and chlorine exist as diatomic molecules. These molecules are completely nonpolar because neither atom in the molecule attracts electrons more strongly. Iodine, $I_2$, is a nonpolar molecule. In this activity, you will determine whether water or ethanol dissolves iodine better.

1. Dip a cotton swab in tincture of iodine, and make two small spots on the palm of your hand. Let the spots dry. The spots that remain are iodine.
2. Dip a second cotton swab in water, and wash one of the iodine spots with it. What happens to the iodine spot?
3. Dip a third cotton swab in ethanol, and wash the other iodine spot with it. What happens to the iodine spot?
4. Did water or ethanol dissolve the iodine spot better? Is water more polar or less polar than ethanol? Write a paragraph explaining your reasoning. **WRITING SKILL**

### Like dissolves like

A rule of thumb in chemistry is that like dissolves like. This rule means that a solvent will dissolve substances that are like the solvent in molecular structure. For example, water is a polar molecule because it has positive and negative ends. So, water dissolves ions and polar molecules, which also have charges.

**Nonpolar compounds** usually will not dissolve in water. Nonpolar molecules do not have charges on opposite sides, like water molecules have. For example, olive oil is not soluble in water because it is a mixture of nonpolar compounds. Water does not attract nonpolar molecules enough to overcome the attractive forces between the nonpolar molecules. So, nonpolar solvents must be used to dissolve nonpolar materials, as shown in ***Figure 14.*** Nonpolar solvents that are distilled from petroleum are often used to dissolve nonpolar materials.

**nonpolar compound** a compound whose electrons are equally distributed among its atoms


***Figure 14***

Nonpolar substances, such as oil-based paint, must be dissolved by a like solvent. This means that a nonpolar solvent must be used.

## The Dissolving Process

According to the kinetic theory of matter, water molecules in a glass of tea are always moving. When sugar is poured into the tea, water molecules collide and transfer energy to the sugar molecules at the surface of the sugar crystal. This energy, as well as the attractive forces between water and sugar molecules, causes molecules at the surface of the crystal to dissolve.

Every time a layer of sugar molecules leaves the crystal, another layer of sugar molecules is uncovered. Sugar molecules break away from the crystal layer by layer in this way until the crystal completely dissolves.

### Solutes with a larger surface area dissolve faster

When a solid is crushed into small pieces, it dissolves faster than the same substance in large pieces. Breaking the solid into smaller pieces exposes more surface area, uncovering more of the solute. More molecules exposed to a solvent leads to more solute-solvent collisions. Therefore, the solid dissolves faster.

***Figure 15*** shows how breaking a solid into many smaller pieces increases the total exposed surface area. The large crystal is a cube that is 1 cm on each edge, so each face of the crystal has an area of 1 $cm^2$. A cube has six surfaces, so the total surface area is 6 $cm^2$. Suppose the large cube is cut into 1000 cubes that are 0.1 cm on each edge. The face of each small cube has an area of $0.1 \text{ cm} \times 0.1 \text{ cm} = 0.01 \text{ cm}^2$, so each cube has a surface area of 0.06 $cm^2$. The total surface area of the small cubes is $1000 \times 0.06 \text{ cm}^2 = 60 \text{ cm}^2$, which is ten times the surface area of the large cube.

***Figure 15***

**A** This salt crystal has a small surface area compared to its total volume.

**B** When the salt crystal is crushed into small pieces, more of the salt can be exposed to a solute, which allows the salt to dissolve faster.

**A**

**B**

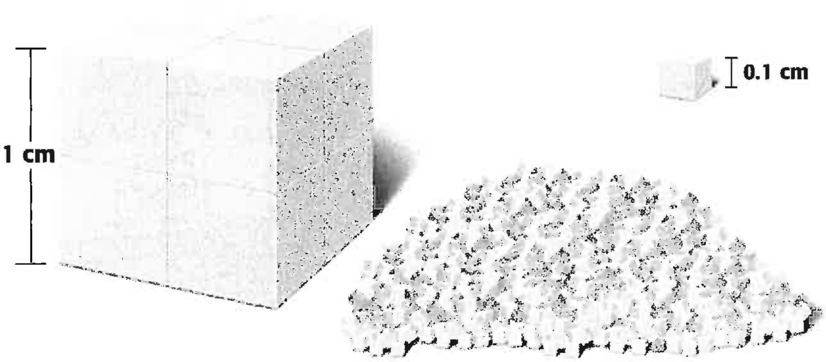
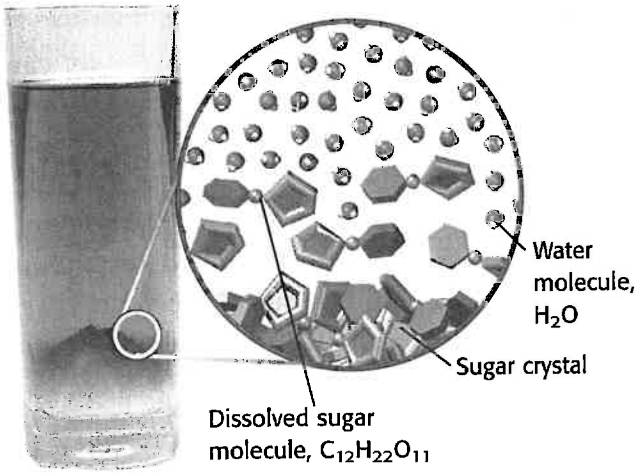
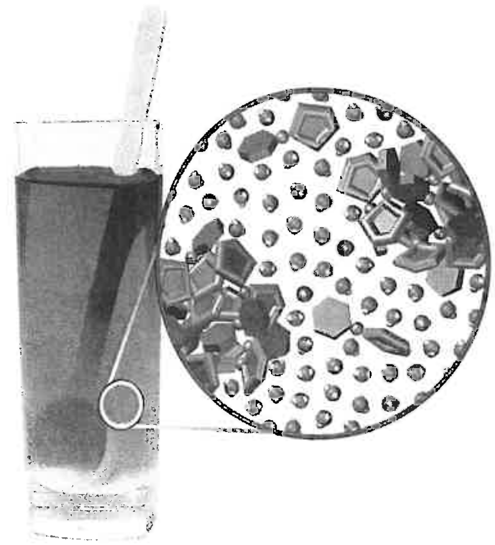


**Figure 16**
**How Stirring Affects the Dissolving Process**

**A** If the solution is not stirred, crystals of table sugar, or sucrose, become surrounded by dissolved sugar molecules, which prevent water molecules from reaching the undissolved sugar.




**B** Stirring moves the dissolved sugar molecules away from the undissolved sugar, and more water molecules can reach the undissolved sugar.



## Stirring or shaking a solution helps the solute dissolve faster

If you pour sugar in a glass of water and let it sit without stirring, the sugar will take a long time to dissolve completely. The sugar takes a long time to dissolve because it is at the bottom of the glass surrounded by dissolved sugar molecules, as shown in ***Figure 16A.*** Dissolved sugar molecules will slowly *diffuse,* or spread out, throughout the entire solution. But until that happens, the dissolved sugar molecules cluster near the surface of the crystal. These dissolved molecules keep water molecules from reaching the sugar that has not yet dissolved.

Stirring or shaking the solution moves the dissolved sugar away from the sugar crystals. So, more water molecules can interact with the solid, as shown in ***Figure 16B,*** and the sugar crystals dissolve faster.

## Solutes dissolve faster when the solvent is hot

Solutes dissolve faster in a hot solvent than in a cold solvent. The kinetic theory states that when matter is heated, its particles move faster. As a result of heating, particles of solvent collide with undissolved solute more often. These collisions also transfer more energy than collisions that occur when the solvent is cold. The greater frequency and energy of the collisions help "knock" undissolved solute particles away from each other and spread them throughout the solution.

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*Figure 17*

The coolant mixture of ethylene glycol and water keeps the radiator fluid from freezing in winter and boiling in summer.

## Solutes affect the physical properties of a solution

The boiling point of pure water is 100°C. But, if you dissolve 12 g of sodium chloride in 100 mL of water, you will find that the boiling point of the solution is increased to about 102°C. Also, the freezing point of the solution will be lowered to about −7°C. Many solutes increase the boiling point of a solution above that of the pure solvent. Likewise, the same solutes lower the freezing point of the solution below that of the pure solvent.

The effect of a solute on freezing point and boiling point can be useful, as shown in ***Figure 17.*** For example, a car's cooling system often contains a mixture that is 50% water and 50% ethylene glycol, a type of alcohol. This solution acts as antifreeze in cold weather because its freezing point is about −30°C. The solution also helps prevent boiling in hot weather because its boiling point is about 109°C. Adding sodium chloride to ice can lower the freezing point of the mixture to about −15°C, which is cold enough to freeze ice cream.

## SECTION 2 REVIEW

### SUMMARY

- Water dissolves ionic compounds and polar molecular compounds because water molecules are polar.
- To dissolve a solute, a solvent must attract solute particles more strongly than solute particles attract each other.
- Nonpolar compounds do not dissolve in water but may be dissolved by nonpolar solvents.
- Solids with a greater surface area dissolve faster.
- Stirring or heating a solvent dissolves solutes faster.
- Solutes can lower the freezing point of a solution and raise its boiling point.

1. **Explain** how water can dissolve some ionic compounds, such as ammonium chloride, $NH_4Cl$, as well as some molecular compounds, such as methanol.
2. **Describe** the relationship of attractive forces between molecules and the ability of a solvent to dissolve a substance.
3. **Explain** why large crystals of coarse sea salt take longer to dissolve in water than crystals of fine table salt.
4. **Use** the like dissolves like rule to predict whether glycerol, which is a polar molecular compound, is soluble in water.
5. **Describe** three methods you could use to make a spoonful of salt dissolve faster in water.
6. **Critical Thinking** You have made some strawberry-flavored drink from water, sugar, and drink mix. You decide to freeze the mixture into ice cubes, so you pour the liquid into an ice-cube tray and place it in the freezer along with another tray of plain water. Two hours later, you find the water has frozen, but the fruit drink has not. How can you explain this result?

SECTION 3

# Solubility and Concentration

OBJECTIVES

- **Explain** the meaning of solubility and compare the solubilities of various substances.
- **Describe** dilute, concentrated, saturated, unsaturated, and supersaturated solutions.
- **Relate** changes in temperature and pressure to changes in solubility of solid and gaseous solutes.
- **Express** the concentration of a solution as molarity, and calculate the molarity of a solution given the amount of solute and the volume of the solution.

KEY TERMS

**solubility**
**concentration**
**unsaturated solution**
**saturated solution**
**supersaturated solution**
**molarity**

How much would you have to shake, stir, or heat the olive oil and water mixture shown in ***Figure 18*** to dissolve the oil in the water? The answer is that the oil would never dissolve in the water no matter what you did. Some substances are *insoluble* in water, meaning they do not dissolve. Other substances such as sugar and baking soda are said to be *soluble* in water because they dissolve easily in water. However, there is often a limit to how much of a substance will dissolve.

**solubility** the maximum amount of a solute that will dissolve in a given quantity of solvent at a given temperature and pressure

## Solubility in Water

Have you ever tried to dissolve several spoonfuls of salt in a glass of water? If you have tried, you may have observed that some of the salt did not dissolve, no matter how much you stirred. Unlike olive oil, salt is soluble in water, but the amount of salt that will dissolve is limited. The maximum amount of salt that can be dissolved in 100 g of water at room temperature is 36 g, or about two tablespoonfuls. The maximum mass of a solute that will dissolve in 100 g of water at 20°C and standard atmospheric pressure is known as the **solubility** of the substance in water.

Some substances such as acetic acid, methanol, ethanol, glycerol, and ethylene glycol are completely soluble in water. This means that any amount of the substance will mix with water to form a solution. Some ionic compounds, such as silver chloride, AgCl, are almost completely insoluble in water. Only 0.000 19 g of AgCl will dissolve in 100 g of water.

***Figure 18***
Olive oil and water form two layers when they are mixed because olive oil is *insoluble* in water.

***Table 1*** **Solubilities of Some Ionic Compounds in Water**

| Substance | Formula | Solubility in g/100 g $H_2O$ at 20°C |
|---|---|---|
| Silver nitrate | $AgNO_3$ | 216 |
| Silver chloride | AgCl | 0.000 19 |
| Sodium fluoride | NaF | 4.06 |
| Sodium chloride | NaCl | 35.9 |
| Sodium iodide | NaI | 178 |
| Calcium chloride | $CaCl_2$ | 75 |
| Calcium sulfate | $CaSO_4$ | 0.32 |
| Calcium fluoride | $CaF_2$ | 0.0015 |
| Sodium sulfide | $Na_2S$ | 26.3 |
| Iron(II) sulfide | FeS | 0.0006 |

## Different substances have different solubilities

Even closely related compounds can have very different solubilities. Compare the solubilities of some of the ionic compounds in ***Table 1***. All compounds listed that contain sodium also contain one other element, but sodium iodide is much more soluble than either sodium chloride or sodium fluoride.

The solubility of any substance in water depends on the strength of the forces acting between water molecules and solute particles compared to the forces acting between the solute particles. Sodium iodide is more soluble than many other compounds that also contain sodium because the forces between water molecules and particles of sodium iodide are much greater than the forces between sodium iodide particles.

## How much of a substance can dissolve in a solvent?

Because substances vary greatly in solubility, you need a way to specify how much solute is dissolved in a given solution. One way is to refer to a solution as *weak* if only a small amount of solute is dissolved or *strong* if a large amount of solute is dissolved. However, *weak* and *strong* mean different things to different people. For example, most people would describe the sulfuric acid solution found in automobile batteries as strong. The solution can injure the skin and react with the fiber in textiles to create holes in clothing. A chemist, though, might describe battery acid as a weak solution of sulfuric acid because the chemist knows that much stronger solutions of sulfuric acid can be prepared.

The terms *weak* and *strong* do not specify the **concentration** of a solution. Concentration is the quantity of solute dissolved in a given volume of solution. Scientists refer to a solution that contains a large amount of solute as *concentrated*. A solution that contains only a small amount of solute is *dilute*.

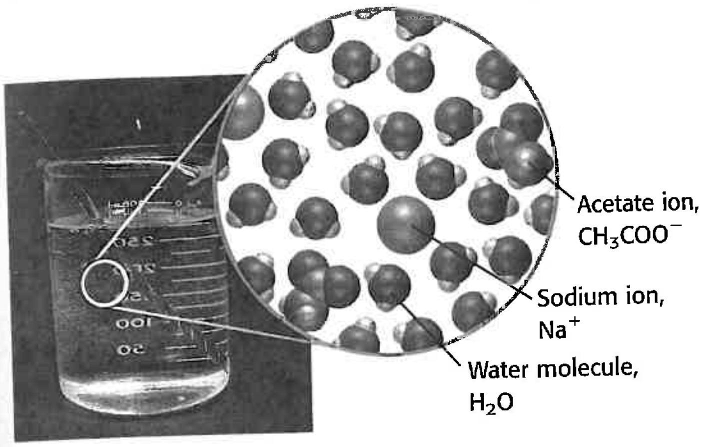
**concentration** the amount of a particular substance in a given quantity of a mixture, solution, or ore

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

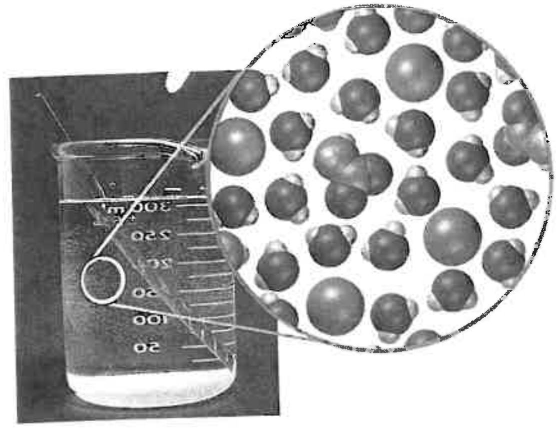
## Unsaturated solutions can dissolve more solute

The terms *concentrated* and *dilute* give no information about the actual quantity of solute dissolved in a solution. An **unsaturated solution** contains less than the maximum amount of solute that will dissolve in the solvent under the same conditions. The solution of sodium acetate in ***Figure 19A*** is unsaturated. A solution is unsaturated as long as it is able to dissolve more solute.

**Figure 19**
**How Concentration Affects the Dissolving Process**




**A** Additional sodium acetate can dissolve when added to this unsaturated solution.



**B** No more sodium acetate will dissolve in this saturated solution. Any additional sodium acetate that is dissolved causes an equal amount to settle out of the solution.

## At some point, most solutions become saturated with solutes

If you keep adding sodium acetate to the solution in ***Figure 19A,*** the added sodium acetate dissolves until the solution becomes saturated, as shown in ***Figure 19B.*** A **saturated solution** can dissolve no more solute. The dissolved solute is in equilibrium with undissolved solute. To be in equilibrium means that the dissolved solute re-forms crystals at the same rate that the undissolved solute dissolves. So if you add more solute, it just settles to the bottom of the container. No matter how much you stir, no more sodium acetate will dissolve in a saturated solution.

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

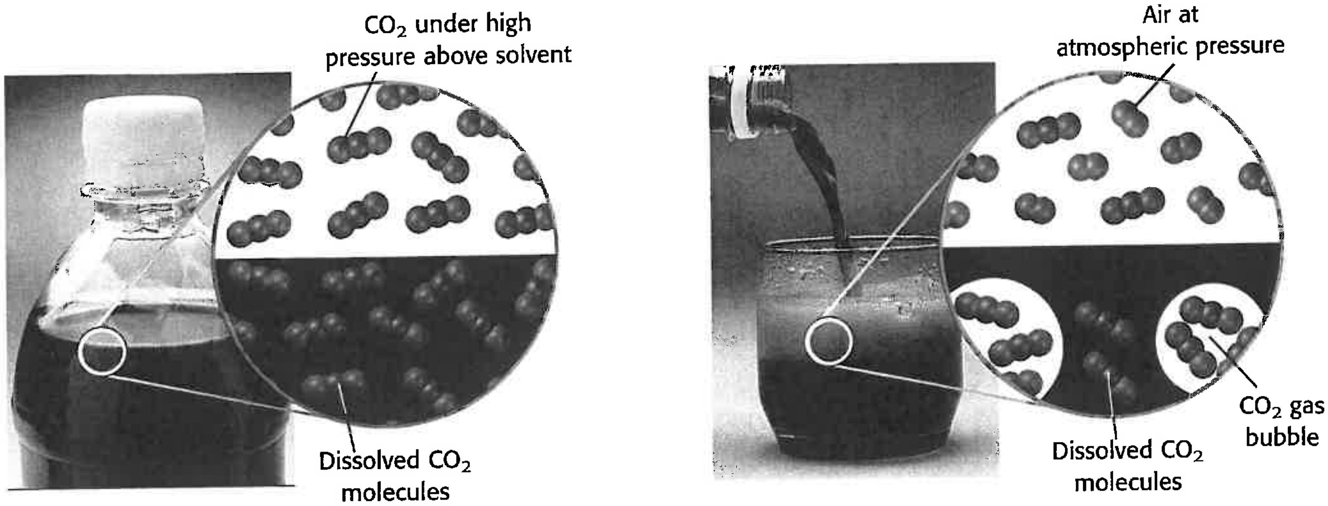
**supersaturated solution** a solution that holds more dissolved solute than is required to reach equilibrium at a given temperature

## Heating a saturated solution usually dissolves more solute

The solubility of most solutes increases as the temperature of the solution increases. If you heat a saturated solution of sodium acetate, more sodium acetate can dissolve until the solution becomes saturated at the higher temperature.

But something interesting happens when the temperature of the solution decreases again. The extra sodium acetate does not re-form crystals, but remains in solution. At the cooler temperature, this unstable **supersaturated solution** holds more solute than it normally can. Adding a small crystal of sodium acetate to the solution provides the surface that the excess solute needs to crystallize, as shown in ***Figure 20.*** The solute crystallizes out of the solution until the solution is saturated at the cooler temperature.

***Figure 20***
Adding a single crystal of sodium acetate to a supersaturated solution causes the excess sodium acetate to quickly crystallize out of the solution.




**A** Carbon dioxide gas is dissolved in this unopened bottle of soda.

**B** When the bottle is opened, the pressure inside the bottle decreases. Carbon dioxide gas then forms bubbles as it comes out of solution.

***Figure 21***

The solubility of a gas in water depends on the pressure of the gas.

### Temperature and pressure affect the solubility of gases

Gases can also dissolve. For instance, some household cleaners contain ammonia gas dissolved in water. Soda is a solution of carbon dioxide gas, sweetener, and flavorings in water. Unlike solid solutes, gaseous solutes are less soluble in warmer water. For example, soda goes flat quickly at room temperature but stays fizzy for a longer period of time when it is cold.

The solubility of gases also depends on pressure, as shown in ***Figure 21.*** Carbon dioxide is dissolved in the soda under high pressure, and the bottle is sealed. When the bottle is opened, the gas pressure decreases to atmospheric pressure and the soda fizzes as the carbon dioxide comes out of solution.

Divers must understand the solubility of gases. Increased pressure underwater causes nitrogen gas to dissolve in the blood. If the diver returns to the surface too quickly, nitrogen comes out of solution and forms bubbles in blood vessels. This condition, called the *bends,* is extremely painful and dangerous.

## Concentration of Solutions

Describing a solution as concentrated, dilute, saturated, or unsaturated still does not reveal the quantity of dissolved solute. For example, only 0.173 g of calcium hydroxide will dissolve in 100 g of water. This solution is saturated but is still dilute because so little solute is present. Scientists express the quantity of solute in a solution in several ways, but one of the most useful forms of expressing this measurement is molarity.

## Molarity is a precise way of measuring concentration

You have already seen that the solubility of a substance can be expressed as grams of solute per 100 grams of solvent. Concentration can also be expressed as **molarity.**

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}, \text{ or } M = \frac{\text{mol}}{\text{L}}$$

Note that molarity is expressed as moles per liter of solution, not per liter of solvent.

A 1.0 M, which is read as "one molar," solution of NaCl contains 1.0 mol of dissolved NaCl in every 1.0 L of solution. To find the molar concentration of a solute, divide the number of moles of solute by the volume of the solution in liters.

**molarity** an expression of the concentration of a solution in moles of dissolved solute per liter of solution

### Math Skills

**Molarity** Calculate the molarity of sodium carbonate, $Na_2CO_3$, in a solution of 38.6 g of solute in 0.500 L of solution.

**1 List the given and unknown values.**

**Given:** mass of sodium carbonate = 38.6 g
volume of solution = 0.500 L

**Unknown:** molarity, amount of $Na_2CO_3$ in 1 L of solution

**2 Write the equation for moles $Na_2CO_3$ and molarity.**

$$\textit{moles } Na_2CO_3 = \frac{\textit{mass } Na_2CO_3}{\textit{molar mass } Na_2CO_3}$$

$$\textit{molarity} = \frac{\textit{moles } Na_2CO_3}{\textit{volume of solution}}$$

**3 Find the number of moles of $Na_2CO_3$ and calculate molarity.**

$$\textit{molar mass } Na_2CO_3 = 106 \text{ g}$$

$$\textit{moles } Na_2CO_3 = \frac{38.6 \text{ g}}{106 \text{ g}} = 0.364 \text{ mol } Na_2CO_3$$

$$\textit{molarity of solution} = \frac{0.364 \text{ mol } Na_2CO_3}{0.500 \text{ L solution}} = 0.728 \text{ M}$$

### Practice HINT

- When calculating molarity, remember that molarity is moles per liter of solution, not moles per liter of solvent.
- If volume is given in milliliters, you must multiply by 1 L/1000 mL to change milliliters to liters. You will need to do this in Practice problems 1c and 1d.

### Practice

**Molarity**

**1.** Determine the molarity of each of the following solutions:

**a.** 2 mol of calcium chloride, $CaCl_2$, dissolved in 1 L of solution

**b.** 0.75 mol of copper(II) sulfate, $CuSO_4$, dissolved in 1.5 L of solution

**c.** 2.25 mol of sulfuric acid, $H_2SO_4$, dissolved in 725 mL of solution

**d.** 525 g of lead(II) nitrate, $Pb(NO_3)_2$, dissolved in 1250 mL of solution

## Other measures of solution concentration can be used

Concentration is sometimes expressed as mass percent, which is grams of solute per 100 g of solution. To make a 5.0% solution of sodium chloride, you would dissolve 5 g of sodium chloride in 95 g of water. The concentration of ordinary rubbing alcohol solution is usually 70%, as shown in ***Figure 22.***

The concentrations of solutions that contain extremely small amounts of solutes are sometimes given in parts per million, ppm, or parts per billion, ppb. Sea water contains slightly more than 1 ppm of fluoride ions, which means that every 1 000 000 g of seawater (about 1000 L) contains 1 g of dissolved fluoride ions. One million grams of water is about the mass of water in a child's plastic wading pool. Sea water also contains about 4 ppb of arsenic, or 4 g arsenic in every 1 000 000 000 g of sea water. The U.S. Environmental Protection Agency has set a maximum safe level of lead in drinking water at 15 ppb.


***Figure 22***

There are 70 grams of alcohol in every 100 g of this solution.

## SECTION 3 REVIEW

### SUMMARY

- The solubility of a substance is the maximum amount that can dissolve in a solvent at a certain temperature and pressure.
- The concentration of a solution is the quantity of solute dissolved per unit volume of solution.
- An unsaturated solution can dissolve more solute.
- A saturated solution cannot dissolve more solute.
- A solute's solubility is exceeded in a supersaturated solution.
- The solubility of a gas in water depends on both pressure and temperature.
- Molarity is a useful way of expressing the concentration of a solution. Molarity is moles of solute per liter of solution.

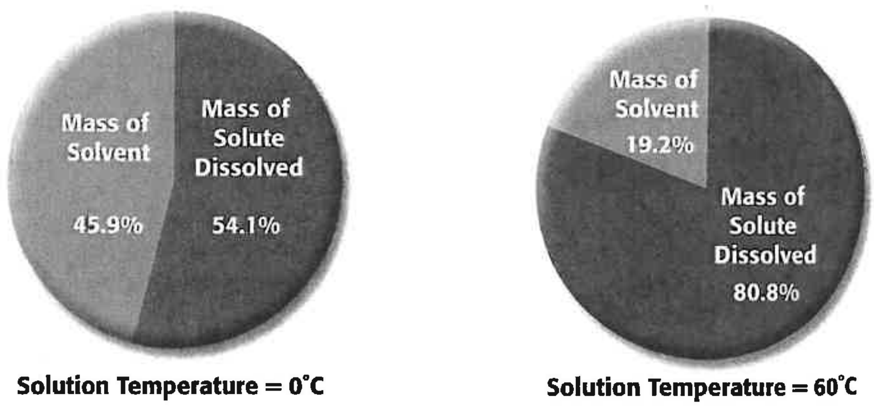
1. **Explain** how a solution can be both saturated and dilute at the same time. Use an example from ***Table 1.***
2. **Describe** how a saturated solution can become supersaturated.
3. **Propose** a way to determine whether a saltwater solution is unsaturated, saturated, or supersaturated.
4. **Compare** the solubility of olive oil and acetic acid in water. What is the solubility of olive oil? of acetic acid?
5. **Determine** whether your sweat would evaporate more quickly if the humidity were 92% or 37%. (**Hint:** When the humidity is 100%, the air is saturated with dissolved water vapor.)
6. **Express** the molarity of a solution that contains 0.5 mol of calcium acetate per 1.0 L of solution.
7. **Critical Thinking** When you fill a glass with cold water from a faucet and then let the glass sit undisturbed for two hours, you will see small bubbles sticking to the glass. What are the bubbles? Why did they form?

### Math Skills

8. **Math Skills** Calculate the molarity of a solution that contains 35.0 g of barium chloride, $BaCl_2$, dissolved in 450.0 mL of solution.

# Graphing Skills

**Maximum Percentage of Ammonium Nitrate Dissolved in Water**




Examine the graphs, and answer the following questions. (See Appendix A for help interpreting a graph.)

1. What type of graphs are these?
2. Identify the quantities that are given in each graph. What important quantities relate the two graphs to each other?
3. By examining the graphs, what can you tell about the solubility of ammonium nitrate?
4. Suppose 500.0 g of water is used to dissolve the ammonium nitrate. What is the maximum mass of ammonium nitrate that can be dissolved at 0°C? What is the maximum mass of ammonium nitrate that can be dissolved at 60°C?
5. From the two graphs, what might the percentage of solute dissolved be if the temperature of the solution is 30°C?
6. The information in the table below shows the solubility of calcium acetate dihydrate in water at various temperatures. Construct a graph best suited for the information listed. How does solubility vary with temperature?

| Temperature (°C) | Maximum mass of solute dissolved in 100 g of water (g) |
|---|---|
| 0 | 37.4 |
| 20 | 34.7 |
| 60 | 32.7 |
| 100 | 29.7 |

# CHAPTER 7 REVIEW

## Chapter Highlights

*Before you begin, review the summaries of key ideas of each section, found at the end of each section. The vocabulary terms are listed on the first page of each section.*

### UNDERSTANDING CONCEPTS

1. Which of the following is a homogeneous mixture?
   **a.** tossed salad  **c.** salt water
   **b.** soil  **d.** vegetable soup

2. If the label on a bottle of medicine reads "shake well before using," the medicine is probably a
   **a.** solution.  **c.** colloid.
   **b.** suspension.  **d.** gel.

3. Which of the following affects the solubility of a solute in a solvent?
   **a.** the surface area of the solute
   **b.** stirring the solution
   **c.** the temperature of the solvent
   **d.** All of the above

4. Suppose you add a teaspoon of table salt to a cool saltwater solution and stir until all of the salt dissolves. The solution you started with was
   **a.** unsaturated.  **c.** saturated.
   **b.** supersaturated.  **d.** concentrated.

5. Which of the following materials is an example of a heterogeneous mixture?
   **a.** air  **c.** water
   **b.** granite  **d.** aluminum

6. Which of the following materials is an example of a solid dissolved in another solid?
   **a.** smoke  **c.** mayonnaise
   **b.** bronze  **d.** ice

7. The dispersed particles of a suspension are ______ than the particles of a colloid.
   **a.** larger  **c.** lighter
   **b.** smaller  **d.** less dense

8. Water attracts the ions of ionic compounds because water molecules are
   **a.** polar.  **c.** magnetic.
   **b.** ions.  **d.** nonpolar.

9. To dissolve a substance, a solvent must attract particles of the substance more strongly than the ______ attract each other.
   **a.** solvent particles  **c.** ions
   **b.** water molecules  **d.** solute particles

10. To cause a solute to dissolve faster, you could
    **a.** add more solute.  **c.** stir the solution.
    **b.** cool the solution.  **d.** saturate the solution.

11. To increase the solubility of a solid substance in a solvent, you could
    **a.** add more solute.  **c.** stir the solution.
    **b.** heat the solution.  **d.** lower the pressure.

12. A solution is ______ if it contains as much dissolved solute as it will hold under a given set of conditions.
    **a.** saturated  **c.** supersaturated
    **b.** dilute  **d.** concentrated

13. Gases are more soluble in liquids when the pressure is ______ and the temperature is ______.
    **a.** high, high  **c.** low, high
    **b.** high, low  **d.** low, low

14. A solution that contains 2.0 mol of dissolved magnesium sulfate in 1.0 L of solution has a concentration of
    **a.** 0.5 M.  **c.** 2.0 M.
    **b.** 1.0 M.  **d.** 4.0 M.

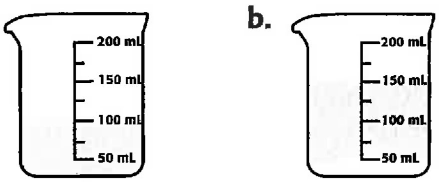
15. The maximum amount of a substance that dissolves in 100 g of water is the ______ of the substance.
    **a.** unsaturation  **c.** dilution
    **b.** molarity  **d.** solubility

16. The boiling point of a solution of sugar in water is ______ the boiling point of water.
    **a.** higher than  **c.** the same as
    **b.** lower than  **d.** not related to

## USING VOCABULARY

**17.** A small amount of *solute* is added to two different solutions. Based on the figures below, which solution was *unsaturated*? Which solution was *saturated*? Explain your answer.

**a.** **b.**




**18.** Explain why an *alloy* is a type of solution. Why are alloys used more often than pure metals?

**19.** You need to clean up a spill of oil-based liquid furniture polish and another spot where some lemon-lime soda has spilled and dried. Explain how you would apply the like dissolves like rule to clean up these two spills. Use the words *polar* and *nonpolar* in your explanation.

**20.** The chemical formula of ethylene glycol is $HOCH_2CH_2OH$, and the chemical formula of 1,2-dichloroethane is $ClCH_2CH_2Cl$. Ethylene glycol is *miscible* in water, but 1,2-dichloroethane is almost completely *immiscible* in water. Explain this difference using two properties of water.

**21.** Ethanol is completely *miscible* in water. Is it possible to prepare a *saturated solution* of ethanol in water? Explain your answer.

**22.** How does temperature affect the *solubility* of a gas solute in a liquid solvent? How does the solubility of a gas solute differ from the solubility of a solid solute in a liquid solvent?

**23.** Classify the following as either a *suspension*, a *colloid*, an *emulsion*, or a *solution*: muddy water, salt water, mayonnaise, vinegar, fog, dry air, and cream.

## BUILDING MATH SKILLS

**24. Molarity** How many moles of lithium chloride are dissolved in 3.00 L of a 0.200 M solution of lithium chloride?

**25. Molarity** What is the molarity of 250 mL of a solution that contains 12.5 g of dissolved zinc bromide, $ZnBr_2$?

**26. Solubility** The solubility of lead(II) chloride, $PbCl_2$, in water at 20°C is 1.00 g $PbCl_2$/100 g of water. If you stirred 7.50 g $PbCl_2$ in 400 g of water at 20°C, what mass of lead(II) chloride would remain undissolved?

**27. Solubility** The solubility of sodium fluoride, NaF, is 4.06 g NaF/100 g $H_2O$ at 20°C. What mass of NaF would you have to dissolve in 1000 g $H_2O$ to make a saturated solution?

## BUILDING GRAPHING SKILLS

**28. Graphing** Make a solubility graph for $AgNO_3$ from the data in the table below. Plot temperature on the *x*-axis, and plot solubility on the *y*-axis. Answer the following questions.

**a.** How does the solubility of $AgNO_3$ vary with the temperature of water?

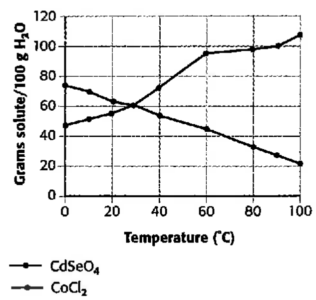
**b.** Estimate the solubility of $AgNO_3$ at 35°C, at 55°C, and at 75°C.

**c.** At what temperature would the solubility of $AgNO_3$ be 512 g per 100 g of $H_2O$?

**d.** If 100 g $AgNO_3$ were added to 100 g $H_2O$ at 10°C, would the solution be saturated or unsaturated?

| Temperature (°C) | Solubility of $AgNO_3$ in g $AgNO_3$/100 g $H_2O$ |
|---|---|
| 0 | 122 |
| 20 | 216 |
| 40 | 311 |
| 60 | 440 |
| 80 | 585 |

**29. Interpreting Graphs** The graph below shows the solubilities of two different substances, cadmium selenate, $CdSeO_4$, and cobalt(II) chloride, $CoCl_2$, over a range of temperatures. Use the graph to answer the following questions.




**a.** At what temperature do both substances have the same solubility?

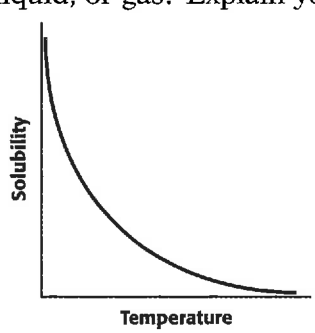
**b.** At what temperature is the solubility of $CdSeO_4$ equal to 40 g $CdSeO_4$/100 g $H_2O$?

**c.** Describe how the solubility of each substance changes with increasing temperature. Which substance has an unusual solubility trend? How is the trend unusual?

**d.** Propose a way to make a supersaturated solution of $CdSeO_4$ at 50°C from a saturated solution.

**e.** Suppose that you add 80 g $CoCl_2$ to 100 g of water at 20°C and stir until no more $CoCl_2$ dissolves. What mass of $CoCl_2$ remains undissolved at 20°C?

**30. Interpreting Graphs** The line in the graph below represents the change in solubility of a substance with the temperature of the solvent. Is the substance most likely to be a solid, liquid, or gas? Explain your reasoning.




## THINKING CRITICALLY

**31. Applying Knowledge** Substances that have similar molecular masses tend to have similar boiling points if no other factors affect the molecules of the substances. The molecular mass of methane, $CH_4$, is 16 amu, and the molecular mass of water is 18 amu. However, the boiling point of methane is −162°C, and the boiling point of water is 100°C. How can you account for the difference of 262°C in the boiling points of methane and water?

Methane Water

**32. Creative Thinking** Sea water is a solution that contains dissolved sodium chloride, dissolved magnesium chloride, and many other dissolved ions. Suppose you have a sample of seawater that contains mud and sand. What is the simplest way to get clear sea water from this mixture? Could you use distillation to separate the sea water from the mud? Explain your answer.

**33. Applying Knowledge** You have been investigating the nature of suspensions, colloids, and solutions and have made the observations on four unknown mixtures in the table below. From your data, decide whether each mixture is a solution, a suspension, or a colloid.

| Sample | Clarity | Settles out | Scatters light |
|---|---|---|---|
| 1 | clear | no | yes |
| 2 | clear | no | no |
| 3 | cloudy | yes | yes |
| 4 | cloudy | no | yes |

34. **Designing Systems** Use what you have learned about the polarity of water molecules and how ionic substances dissolve to make a three-step diagram or computer presentation showing how water molecules attract, surround, and dissolve an ionic substance. COMPUTER SKILL

35. **Problem Solving** Sometimes, powerful searchlights are pointed toward the sky to draw the public's attention to grand openings of stores and restaurants. Even when the air appears clear, you can see the searchlight beam shining into the sky. Explain why these beams are visible in air. WRITING SKILL

## BUILDING LIFE/WORK SKILLS

36. **Communicating Effectively** When there is an oil spill, emergency-response teams use the properties of oil and water along with solubility principles to clean spills and prevent them from spreading. Describe the research behind these techniques, and evaluate the impact this research has had on the environment.

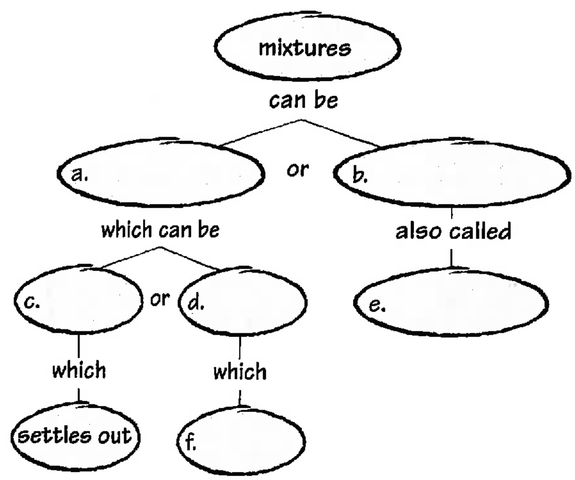
37. **Making Presentations** Research the technique of *reverse osmosis*. Develop a presentation explaining how reverse osmosis produces drinking water from sea water. Include information about the areas of the world where this technology is commonly used and about the energy sources that are used to power the process. Demonstrate a reverse osmosis device from a store that deals in outdoor equipment.

38. **Interpreting Data** Use a reference book to look up the solubilities of the fluoride compounds and chloride compounds made from alkali metals and alkaline earth metals. What conclusions can you draw about the solubilities of the compounds of these metals?

## INTEGRATING CONCEPTS

39. **Connection to Biology** Hydrogen bonding is extremely important to the function of substances within organisms. Use biology books to find three examples in which hydrogen bonds play a crucial role in the life processes of a cell.

40. **Concept Mapping** Copy the unfinished concept map below onto a sheet of paper. Complete the map by writing the correct word or phrase in the lettered boxes.




**internet**connect

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**Topic: Hydrogen Bonding SciLinks code: HK4070**

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