## Solubility

In our earlier experiment, we dissolved 100 g of sugar in flask $W_S$, containing 200 mL of water at 20°C. If we continue to add sugar to the solution, we will eventually reach saturation. If an additional 350 g of sugar (450 g total) are placed in the flask, about 45 g of solute will not dissolve. Now, heat the solution and gently agitate it as you heat. As the temperature rises, what happens to the undissolved solute? Why does more sugar dissolve in hot water than in cold water? We will answer that question a bit later.

The components of a solution do not combine in definite proportions. Therefore, the amount of solute that dissolves in a solvent is not a definite amount—it depends on certain conditions. Even a saturated solution will dissolve more solute if the conditions are correct. ***Solubility*** *is a measure of how much solute will dissolve in a solvent.* Solubility depends on two major factors—the identity of the solution components and the existing conditions of pressure and temperature.

**Identity of solution components.** Recall that sugar dissolves easily in water but does not dissolve in gasoline. Oil does not dissolve in water but does dissolve in gasoline. Some substances easily dissolve in other substances. Other substances will not dissolve, no matter how vigorously they are agitated. Solubility is, in part, determined by the identity of the solution components. The identity of a component is based on its chemical bonds—both molecular bonds (within formula units) and intermolecular bonds (between formula units). The cardinal rule for determining solubility is "like dissolves like." A substance with a particular bond type, polarity, and intermolecular force will generally dissolve a substance with similar properties. Water molecules have highly polar covalent bonds within each molecule and hydrogen bonds between each molecule. A substance, such as sugar, that is also highly polar covalent, will probably dissolve in water. Sugar molecules also contain O—H bonds, which bond with water's hydrogen bonds. Oil molecules are nonpolar with little or no hydrogen bonding, so they are insoluble in water. Gasoline molecules are nonpolar, so they will dissolve oil molecules.

*Two substances in the same phase that are highly soluble in each other are called* ***miscible.*** Molecules of ethanol (ethyl alcohol) and water are both polar and bond with hydrogen bonds. These two liquids will combine in almost any proportion. *Two substances in the same phase that are not soluble in each other are called* ***immiscible.*** Oil and water are immiscible. Ethyl ether molecules are polar but not as polar as water molecules, so ethyl ether and water are partially miscible.

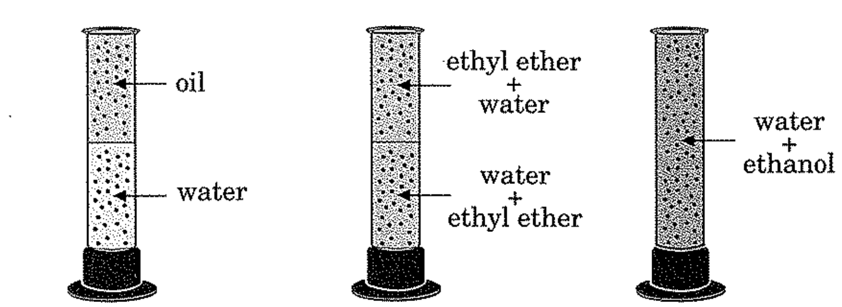
### YOU DO SAY!

The term *miscible* is pronounced: mĭs′ə·bəl.

### YOU DON'T SAY!

If you throw a rock but fail to hit a male bovine, you miscible.

(miss a bull)




Water and oil are immiscible. | Water and ethyl ether are partially miscible. | Water and ethanol are completely miscible.

**Fig. 6.15 Miscibility:** Two liquids such as oil and water are immiscible—they are nearly insoluble. Liquids such as water and ethyl ether are partially miscible. If mixed, they separate into layers—a layer of ether saturated with water and a layer of water saturated with ether. Two liquids such as water and ethanol are completely miscible. When mixed, they form a homogeneous solution.



**Pressure and temperature.** Conditions of pressure and temperature also affect solubility. Pressure has little effect on the solubility of solids or liquids but does affect the solubility of a gas in liquid or solid. When gas at a particular pressure comes in contact with a liquid, gas molecules strike the surface of a liquid, and some molecules change phase and go into solution with the liquid. As more gas molecules enter the solution, some of them leave the solution and return to the gas phase. Equilibrium is achieved when the rate of molecules entering the solution equals the rate of molecules leaving the solution. At equilibrium, the amount of gas dissolved in the liquid is constant.

If the pressure of the gas above a liquid is increased, gas molecules enter the solution at a faster rate. At the same time, more molecules leave the solution and return to the gas phase. Equilibrium is again achieved, but this time a greater amount of solute is dissolved in the liquid. The amount of gas dissolved in a liquid depends on the pressure of the gas above the surface of the liquid. This relationship was first recognized in 1803 by William Henry, an English physician and chemist. ***Henry's law*** states that *the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas on the surface of the liquid.*

Only a limited amount of $CO_2$ dissolves in water at atmospheric pressure. When carbonated drinks are bottled, $CO_2$ is forced into the bottle at a pressure of 5–10 atm, thus greatly increasing the amount of $CO_2$ that goes into solution. When the cap or top is removed, the excess pressure is released, and some of the $CO_2$ leaves the solution. *This rapid escape of gas from a liquid solution is called* ***effervescence.***

Earlier, we mentioned that temperature affects the rate of solution of some substances. Temperature also affects solubility. An increase in temperature of a liquid produces a decrease in the amount of gas that will dissolve in the liquid. Temperature also affects the solubility of liquids and solids, but the effects are not so direct. For most solids, an increase in temperature produces a corresponding increase in the amount of solute that dissolves in a liquid. For some solids, an increase in temperature produces only a small increase in the amount of dissolved solute. Yet, for a few solids, solubility decreases if temperature increases. Why solubility *increases* with an increase in temperature for some substances and *decreases* with an increase of temperature for other substances can be explained using Le Châtelier's principle and heats of solution, which will be discussed in the next section.

Notice from **Table 6.4** on the next page that an increase of temperature from 20°C to 40°C results in an increase in solubility of 34 g for sugar ($C_{12}H_{22}O_{11}$). The same increase in temperature results in an increase of only 0.5 g for salt (NaCl), and a decrease of 0.16 g for lithium carbonate ($Li_2CO_3$).

In the solubility experiment performed earlier, an increase in water temperature increased the amount of sugar that dissolved. In most cases, when a saturated solution dissolves more solute at a higher temperature, the excess solute precipitates when the temperature returns to normal. In a few cases, however, a solution may become supersaturated. *A* ***supersaturated solution*** *is a solution that contains more solute than a normal saturated solution under the same conditions.* Sodium acetate ($NaC_2H_3O_2$) forms a supersaturated solution if a saturated solution is gently heated and agitated while adding more solute. Even after the solution cools, it remains supersaturated. If a crystal of sodium acetate is suspended in the solution, rapid crystallization occurs as the excess solute leaves the solution.

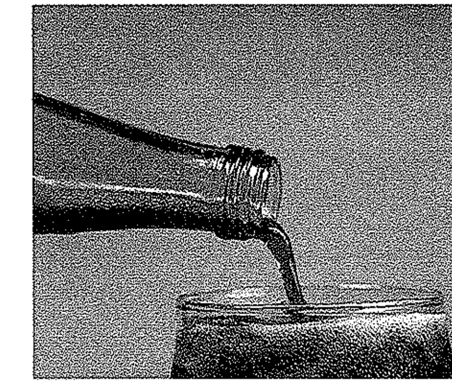


**Fig. 6.16 Effervescence:** When a carbonated drink is opened, $CO_2$ is rapidly released from the solution.

### FOR YOUR INFORMATION

Le Châtelier's (lə · shä′ tə · lyāz) principle was named for Henri Louis Le Châtelier, a French chemist. In 1884 he discovered the thermodynamic principle that a system in equilibrium tends to remain in equilibrium. If something happens to upset the equilibrium, the system will tend to offset the change and return to equilibrium. Le Châtelier's principle will be discussed in more detail in the next section.

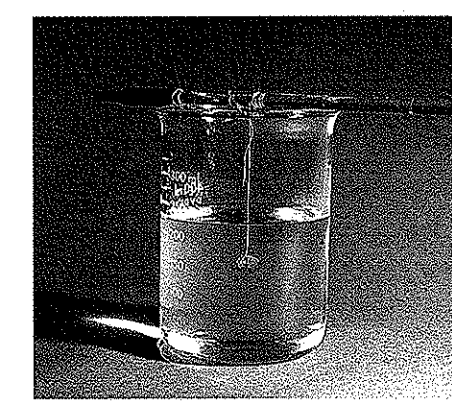


**Fig. 6.17 Supersaturated solution:** A crystal forms rapidly from a supersaturated solution of sodium acetate ($NaC_2H_3O_2$).

Table 6.4

**Solubility of Selected Solids**
(grams of solute/100 g of water)

| Substance | 0°C | 20°C | 40°C | 60°C | 80°C | 100°C |
|---|---|---|---|---|---|---|
| $C_{12}H_{22}O_{11}$ | 179 | 204 | 238 | 287 | 362 | 487 |
| KI | 128 | 144 | 162 | 176 | 192 | 206 |
| $AgNO_3$ | 122 | 216 | 311 | 440 | 585 | 733 |
| $NaNO_3$ | 73.0 | 86.7 | 102 | 122 | 148 | 180 |
| NaCl | 35.7 | 35.9 | 36.4 | 37.1 | 38.0 | 39.2 |
| $KNO_3$ | 13.9 | 31.6 | 61.3 | 106 | 167 | 245 |
| $Li_2CO_3$ | 1.54 | 1.33 | 1.17 | 1.01 | 0.85 | 0.72 |
| $Ca(OH)_2$ | 0.189 | 0.173 | 0.141 | 0.121 | 0.094 | 0.070 |

## Heat of Solution

Here is an experiment that demonstrates the concept of heat of solution. Prepare three beakers with 200 g of water at 20°C. Also, prepare 80 g of calcium chloride ($CaCl_2$) and 60 g of ammonium nitrate ($NH_4NO_3$). Set up the beakers as illustrated in Fig. 6.18 so that the temperature of each beaker can be monitored with a thermometer. Add the $CaCl_2$ to beaker $H_1$ and the $NH_4NO_3$ to beaker $H_2$. Beaker $H_3$ is the control. Stir each beaker until the solutes dissolve. Notice the temperature of each solution. See if your records match those in Table 6.5.

When a solute dissolves in a solvent, an energy change occurs. The change may be exothermic (energy released) or endothermic (energy absorbed). *The amount of heat absorbed or released when a solute dissolves in a specific amount of solvent is called the* ***heat of solution*** ($\Delta H_{soln}$).

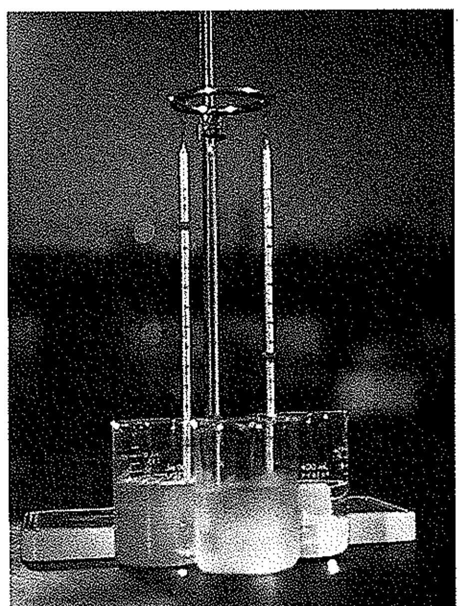


**Fig. 6.18 Heat of solution setup:** This experiment demonstrates the heats of solution of two substances.

Table 6.5

**Heat of Solution Experiment**

| Beaker | Solute | Beginning Temperature | Ending Temperature |
|---|---|---|---|
| $H_1$ | $CaCl_2$ | 20°C | 90°C |
| $H_2$ | $NH_4NO_3$ | 20°C | 0°C |
| $H_3$ | none | 20°C | 20°C |

Although solvation is not a chemical reaction, it does involve several changes in the intermolecular forces of attraction affecting particles of the solute and particles of the solvent. Whenever a change is made in a chemical bond, energy is either absorbed or released. Heat of solution is caused by these changes. Although heat of solution is a single value, it is actually the total heat of three different energy changes caused by three different types of interactions between the particles of the components of a solution. Those interactions are:

- Solute-solute interactions.
- Solvent-solvent interactions.
- Solute-solvent interactions.



Solute particles are held together by intermolecular forces (dipole-dipole forces, hydrogen bonds, or London dispersion forces). The same is true for solvent particles. Before a solvent particle in a solid-liquid solution can attract a solute particle away from its solute crystal, the solvent particle must be free from the intermolecular forces that bind it to other solvent particles (solvent-solvent bond). At the same time, the solute particle must be free from intermolecular forces that bind it to other solute particles (solute-solute bond). When a solvent particle comes near a solute particle, the solvent particle begins to exert an attraction for the solute particle. As this happens, both the solvent-solvent bond and the solute-solute bond are broken. The breaking of these bonds is endothermic—energy is absorbed by the particles. Now that the solute particle and the solvent particle are free to interact, a bond forms between them. The formation of this bond is exothermic—energy is released.

The energy required to break the solvent-solvent bonds and the solute-solute bonds comes from the solution. When these bonds are broken, the temperature of the solution decreases slightly. When solute-solvent bonds are formed, energy is released, and the temperature of the solution increases. If the total energy released by the formation of solute-solvent bonds is greater than the total energy absorbed by the breaking of solvent-solvent and solute-solute bonds, the net heat of solution is negative ($-\Delta H_{soln}$), and the temperature of the solution rises. If the total energy released by the formation of solute-solvent bonds is less than the total energy absorbed by the breaking of solvent-solvent and solute-solute bonds, the net heat of solution is positive ($+\Delta H_{soln}$), and the temperature of the solution drops.

**IN CASE YOU'RE CONFUSED**

In a general discussion about solutions, we often use the generic term *particle* to refer to a formula unit of solute or solvent. The reason for this is that, in a solution, a solvated solute may consist of an ion, a molecule, or even a group of molecules. Rather than differentiate between each type of formula unit, we categorically refer to them all as particles.

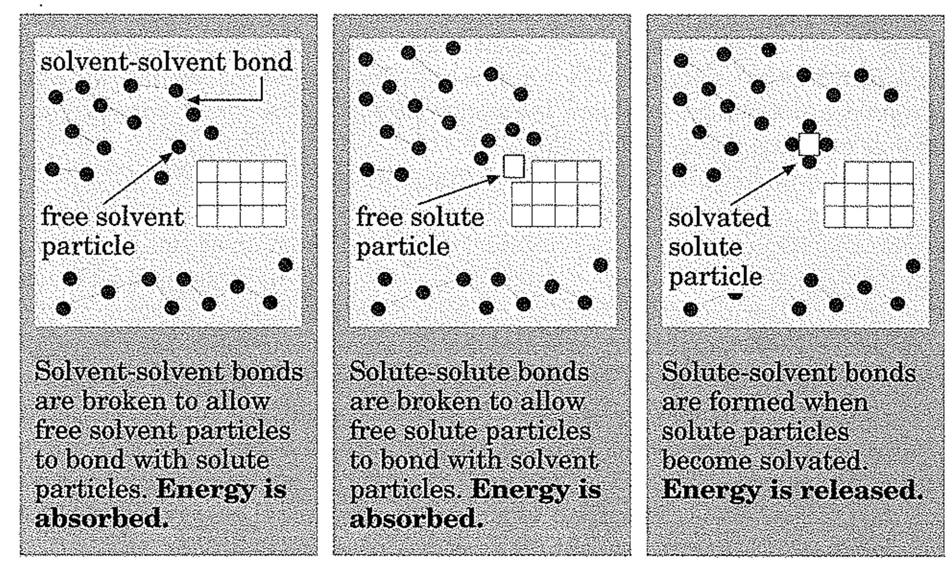



**Fig. 6.19 Heat of solution:** Heat of solution involves energy from solvent-solvent interactions, solute-solute interactions, and solute-solvent interactions.

Heat of solution is measured in kilojoules per mole (kJ/mol) of solute dissolved in a specified amount of solvent. A solute with a high negative heat of solution will greatly raise the temperature of the solution. A solute with a high positive heat of solution will greatly lower the temperature of the solution. See Table 6.6 for a list of common heats of solution.

Notice from Table 6.6, that $CaCl_2$ has a $-\Delta H_{soln}$ of −82.8 kJ/mol and that $NH_4NO_3$ has a $+\Delta H_{soln}$ of +25.5 kJ/mol. From this information, we can predict that dissolving $CaCl_2$ in water will raise the temperature of the water and that dissolving $NH_4NO_3$ in water will lower the temperature. Did the substances in our heat of solution experiment behave according to their heats of solution?

$-\Delta H_{soln}$ = exothermic
solute + solvent $\rightarrow$ solution + heat

$+\Delta H_{soln}$ = endothermic
solute + solvent + heat $\rightarrow$ solution

On pages 1–9 in the PACE, study carefully:
I. Solutions
Solution Chemistry
Solution Process
Solution Rate

**NOTE: Some activities ask you to define a word found in the PACE text. You may need to use a dictionary to find the best meaning and complete the activity.**

**Match these items.**

_______ (1) dissolving medium of a solution
_______ (2) substance dissolved in a solution
_______ (3) general term for the process of dissolving
_______ (4) combination of a base metal and at least one other substance, which blend to form a solution with metallic properties
_______ (5) solution formed with liquid mercury blended with another metal or metals
_______ (6) condition in which a solution contains the maximum amount of solute that may be dissolved
_______ (7) physical state in which solvation and crystallization occur at the same rate
_______ (8) formation of ions when a molecular compound dissolves
_______ (9) formation of ions when an ionic compound dissolves
_______ (10) part of an experiment that is conducted without the same variables or conditions being tested
_______ (11) process of dissolving a solute in water

a. alloy
b. amalgam
c. control
d. dissociation
e. hydration
f. ionization
g. saturation
h. solute
i. solution
j. solution equilibrium
k. solvation
l. solvent

**Fill in the blanks.**

(12) A spontaneous chemical reaction begins as soon as the _______________ _______________.

(13) Reactants in the same _______________ usually do not react unless _______________ is applied or a _______________ is added.

(14) Some chemical reactions occur only if the reactants are in _______________, which means they are dissolved in water.

(15) Many single and double _______________ reactions involve reactants and/or products in aqueous solution.

Cover photograph: Laboratory (Courtesy of Digital Equipment Corporation, Corporate Photo Library)





(16) The properties of a solution include:

(a) A _______________ combination with a _______________ consistency.

(b) Components that do not _______________ combine in definite _______________.

(c) Components in a single _______________.

(d) A _______________ dissolved in a _______________.

(17) The type of solution is determined by the phase of the _______________; the form of the solution is determined by the phase of the _______________.

(18) An _______________ is a combination of a base metal and at least one other substance, which blend to form a solution with metallic properties.

(19) Why are alloys useful? _______________

_______________

_______________

(20) An alloy of iron blended with carbon is called _______________.

(21) List the reasons why heating iron over charcoal produced better iron.

(a) _______________

(b) _______________

(22) An _______________ is a solution of liquid mercury blended with another metal or metals.

(23) Sugar dissolves in water because sugar and water are both _______________ compounds consisting of _______________ molecules.

(24) The _______________ force of attraction between polar solvent molecules and polar solute molecules pulls solute molecules away from the surface of the solid.

(25) A solute molecule surrounded by solvent molecules is said to be _______________; if the solvent is water, the solute is said to be _______________.

**Essay: Write a few good sentences about the following topic.**

(26) Explain why a sugar cube disappears when it dissolves in water. _______________

_______________

_______________

_______________

_______________

_______________

**Fill in the blanks.**

(27) The condition called ______________ occurs when a solution contains the maximum amount of solute that may be dissolved.

(28) When solvation and crystallization occur at the same rate, ______________ is achieved.

(29) ______________ is the formation of ions when a molecular compound dissolves.

(30) ______________ is the formation of ions when an ionic compound dissolves.

(31) The three factors that affect the rate at which a solution dissolves are ______________ the solvent, increasing the ______________ of the solute, and increasing the ______________ of the solvent.

**Complete these statements that explain the following processes regarding a sugar-water solution.**

(32) Explain the solution process leading to solution equilibrium.

(a) A sugar molecule is ______________ away from the crystal.

(b) Water molecules ______________ the sugar molecule as soon as it is freed from the crystal.

(c) The ______________ sugar molecule is carried away from the crystal.

(d) Other sugar molecules become ______________ and are carried away from the crystal.

(e) Since hydrated sugar molecules are surrounded by ______________, the force of attraction is slight between sugar molecules, and sugar molecules stay in solution unless ______________.

(f) As more and more sugar molecules dissolve, the ______________ of hydrated sugar molecules increases, and hydrated sugar molecules bump into ______________.

(g) The force of attraction between hydrated sugar molecules and undissolved sugar molecules is sufficient to attract ______________ back into the crystal.

(h) When the rate of solvation equals the rate of crystallization, ______________ is reached.

(i) Since two opposing forces occur at the same rate, ______________ is established.

(33) Explain why agitating water affects the solvation rate of sugar.

(a) When sugar dissolves, ______________ carry ______________ away from the crystal.

(b) As hydrated sugar molecules are carried away, ______________ move in to hydrate other sugar molecules.

(c) Water molecules move by ______________, which is not very fast.



(d) Hydrated sugar molecules congest around the crystal, and ______________ cannot move in.

(e) When the water is ______________, water currents move the hydrated molecules away from the crystal and bring in fresh water molecules.

(f) Because water molecules get to the crystal quicker, they ______________ faster, and the sugar dissolves faster.

**Place the letter of the correct answer in the blank.**

(34) 40 grams of potassium nitrate and 100 grams of water form a __________ solution.

a. liquid-liquid b. solid-liquid c. liquid-solid d. liquid-gaseous

(35) 1.5 liters of ethylene glycol and 4.5 liters of water form a __________ solution.

a. liquid-liquid b. solid-liquid c. liquid-solid d. liquid-gaseous

(36) Water vapor and air form a __________ solution.

a. liquid-liquid b. solid-liquid c. liquid-solid d. liquid-gaseous

(37) 25 grams of copper and 75 grams of gold form a __________ solution.

a. solid-solid b. solid-liquid c. solid-gaseous d. liquid-gaseous

(38) A combination of 17% oxygen gas and 83% nitrogen gas is a __________ solution.

a. solid-solid b. solid-gaseous c. liquid-gaseous d. gaseous-gaseous

(39) Your favorite cola contains two solutions—a sugar and water solution, which is a __________ solution, and a carbon dioxide and water solution, which is a __________ solution.

a. solid-liquid b. liquid-liquid c. gaseous-liquid d. liquid-gaseous

(40) Iodine crystals sublimate and mix with air to form a __________ solution.

a. solid-gaseous b. liquid-gaseous c. gaseous-gaseous d. gaseous-solid

(41) Oxygen and aquarium water form a __________ solution.

a. solid-liquid b. liquid-liquid c. gaseous-liquid d. liquid-gaseous

(42) Hydrogen gas and palladium metal form a __________ solution.

a. solid-gaseous b. gaseous-solid c. gaseous-gaseous d. solid-solid

(43) When water and oil mix, __________ forms.

a. a solid-liquid solution b. a liquid-liquid solution c. a gaseous-liquid solution d. no solution

*Keep your face to the sunshine and you will never see the shadows.*