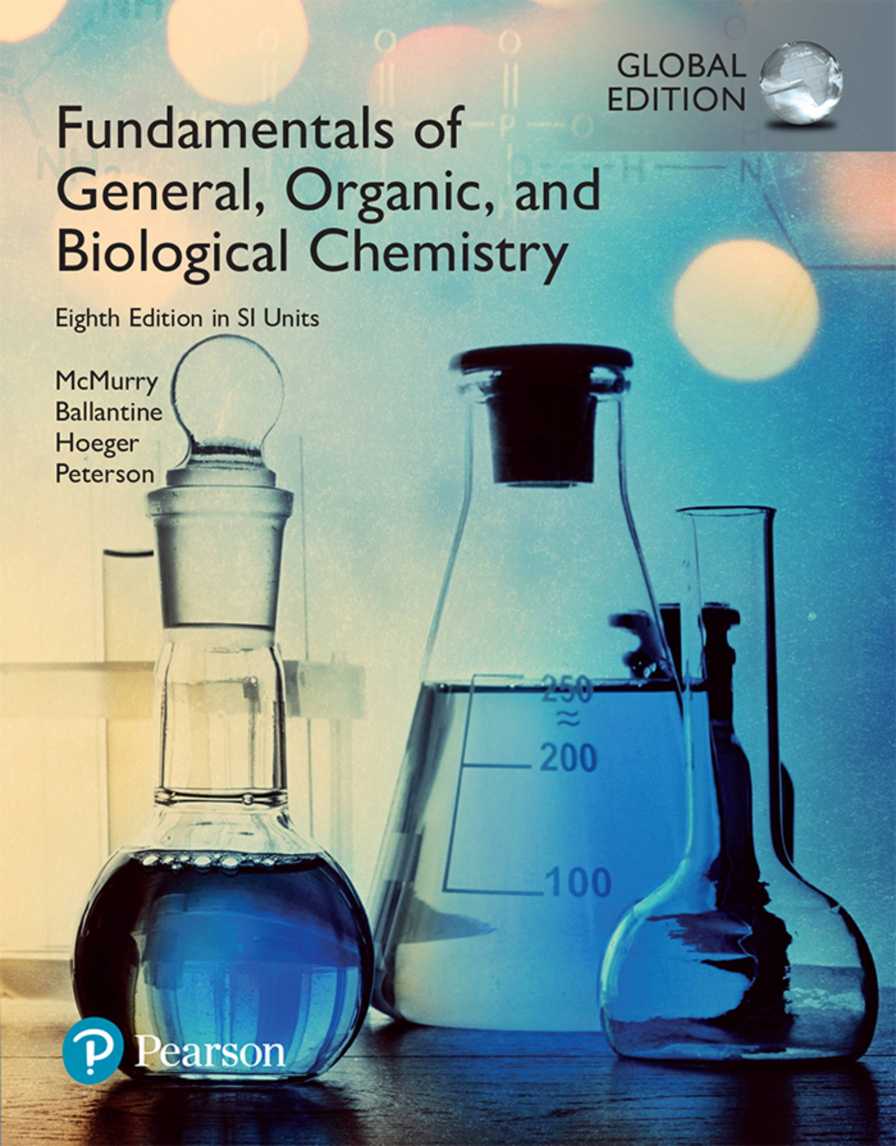


GLOBAL
EDITION

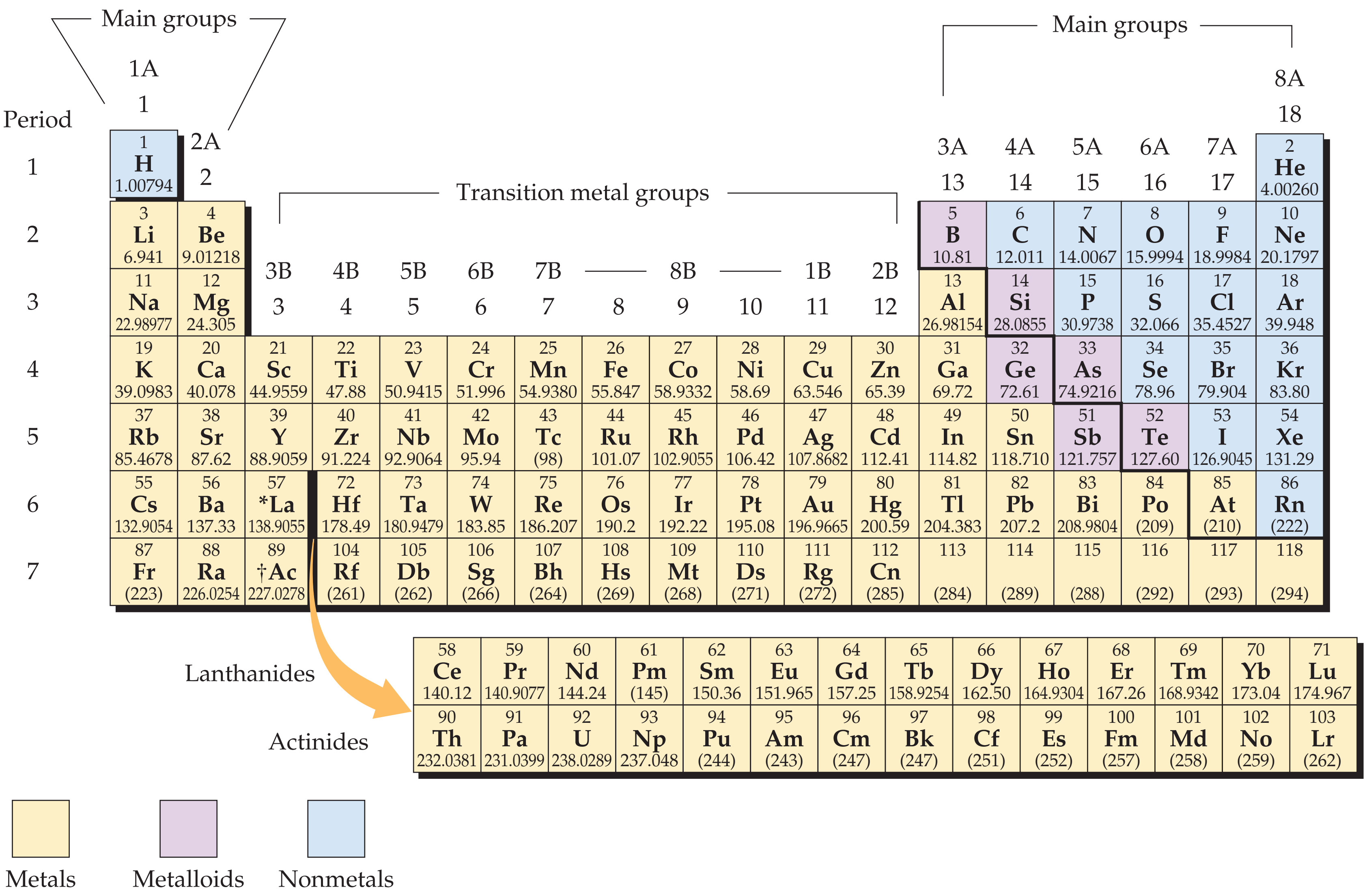
Fundamentals of General, Organic, and Biological Chemistry

Eighth Edition in SI Units

McMurry
Ballantine
Hoeger
Peterson

Pearson

Periodic Table of the Elements



	Main groups		Transition metal groups										Main groups					
Period	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8	8B 9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H** 1.00794																	2 **He** 4.00260
2	3 **Li** 6.941	4 **Be** 9.01218											5 **B** 10.81	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984	10 **Ne** 20.1797
3	11 **Na** 22.98977	12 **Mg** 24.305											13 **Al** 26.98154	14 **Si** 28.0855	15 **P** 30.9738	16 **S** 32.066	17 **Cl** 35.4527	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.996	25 **Mn** 54.9380	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.69	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.61	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.9055	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.710	51 **Sb** 121.757	52 **Te** 127.60	53 **I** 126.9045	54 **Xe** 131.29
6	55 **Cs** 132.9054	56 **Ba** 137.33	57 ***La** 138.9055	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.85	75 **Re** 186.207	76 **Os** 190.2	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.9665	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.9804	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** 226.0254	89 **†Ac** 227.0278	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (269)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)	112 **Cn** (285)	113 (284)	114 (289)	115 (288)	116 (292)	117 (293)	118 (294)

Lanthanides	58 **Ce** 140.12	59 **Pr** 140.9077	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.965	64 **Gd** 157.25	65 **Tb** 158.9254	66 **Dy** 162.50	67 **Ho** 164.9304	68 **Er** 167.26	69 **Tm** 168.9342	70 **Yb** 173.04	71 **Lu** 174.967
Actinides	90 **Th** 232.0381	91 **Pa** 231.0399	92 **U** 238.0289	93 **Np** 237.048	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

🟨 Metals

🟪 Metalloids

🟦 Nonmetals

CHEMISTRY IN ACTION

Solid Hydrates—Salt + Water

If you add salt to water, you would expect it to dissolve and form a solution. But some ionic compounds attract water strongly enough to hold on to water molecules even when crystalline, forming what are called *solid hydrates.* For example, the plaster of Paris used to make decorative objects and casts for broken limbs is calcium sulfate hemihydrate, $CaSO_4 \cdot \frac{1}{2}H_2O$. The dot between $CaSO_4$ and $\frac{1}{2}H_2O$ in the formula indicates that for every two $CaSO_4$ formula units in the crystal there is also one water molecule present.

$$CaSO_4 \cdot \tfrac{1}{2}H_2O \quad \text{A solid hydrate}$$

After being ground up and mixed with water to make plaster, $CaSO_4 \cdot \frac{1}{2}H_2O$ gradually changes into the crystalline dihydrate $CaSO_4 \cdot 2H_2O$, known as *gypsum.*

During the change, the plaster hardens and expands in volume, causing it to fill a mold or shape itself closely around a broken limb. Still other ionic compounds attract water so strongly that they pull water vapor from humid air to become hydrated. Compounds that show this behavior, such as calcium chloride $(CaCl_2)$, are called hygroscopic and are often used as drying agents. You might have noticed a small bag of a hygroscopic compound (probably silica gel, SiO_2) included in the packing material of a new MP3 player, camera, or other electronic device to keep humidity low during shipping. These and other ionic compounds that are handled primarily as hydrates are listed in the following table.

Some Common Solid Hydrates

Formula	Name	Uses
$AlCl_3 \cdot 6H_2O$	Aluminum chloride hexahydrate	Antiperspirant
$CaSO_4 \cdot 2H_2O$	Calcium sulfate dihydrate (gypsum)	Cements, wallboard molds
$CaSO_4 \cdot \frac{1}{2}H_2O$	Calcium sulfate hemihydrate (plaster of Paris)	Casts, molds
$CuSO_4 \cdot 5H_2O$	Copper(II) sulfate pentahydrate (blue vitriol)	Pesticide, germicide, topical fungicide
$MgSO_4 \cdot 7H_2O$	Magnesium sulfate heptahydrate (epsom salts)	Laxative, anticonvulsant
$NaB_4O_7 \cdot 10H_2O$	Sodium tetraborate decahydrate (borax)	Cleaning compounds, fireproofing agent
$Na_2S_2O_3 \cdot 5H_2O$	Sodium thiosulfate pentahydrate (hypo)	Photographic fixer

CIA Problem 9.1 Write the formula of sodium sulfate decahydrate, known as Glauber's salt and used as a laxative.

CIA Problem 9.2 What mass of Glauber's salt must be used to provide 1.00 mol of sodium sulfate?

PROBLEM 9.2

Which of the following pairs of substances would you expect to form solutions?

(a) CCl_4 and water
(b) Benzene (C_6H_6) and $MgSO_4$
(c) Hexane (C_6H_{14}) and heptane (C_7H_{16})
(d) Ethanol (C_2H_5OH) and heptanol $(C_7H_{15}OH)$

9.3 Solubility

Learning Objective:

- Define the properties of a solution, including miscibility, saturation, and solubility.

We learned in Section 9.2 that ethanol is soluble in water because hydrogen bonding is nearly as strong between water and ethanol molecules as it is between water molecules alone or ethanol molecules alone. So similar are the forces in this particular case, in fact, that the two liquids are **miscible** or mutually soluble in all proportions. Ethanol will continue to dissolve in water no matter how much is added.

Miscible Mutually soluble in all proportions.

Most substances, however, reach a solubility limit beyond which no more will dissolve in solution. Imagine, for instance, that you are asked to prepare a saline solution (aqueous NaCl). You might measure out some water, add solid NaCl, and stir the mixture. Dissolution occurs rapidly at first but then slows down as more and more NaCl is added. Eventually the dissolution stops because an equilibrium is reached when the numbers of Na^+ and Cl^- ions leaving a crystal and going into solution are equal to the numbers of ions returning from solution to the crystal. At this point, the solution is said to be **saturated.** A maximum of 35.8 g of NaCl will dissolve in 100 mL of water at 20 °C (293 K). Any amount above this limit simply sinks to the bottom of the container and sits there.

Saturated solution A solution that contains the maximum amount of dissolved solute at equilibrium.

The equilibrium reached by a saturated solution is like the equilibrium reached by a reversible reaction (Section 7.7). Both are dynamic situations in which no *apparent* change occurs because the rates of forward and backward processes are equal. Solute particles leave the solid surface and reenter the solid from solution at the same rate.

$$\text{Solid solute} \underset{\text{Crystallize}}{\overset{\text{Dissolve}}{\rightleftharpoons}} \text{Solution}$$

The maximum amount of a substance that will dissolve in a given amount of a solvent at a given temperature, usually expressed in grams per 100 mL (g/100 mL), is called the substance's **solubility.** Solubility is a characteristic property of a specific solute–solvent combination, and different substances have greatly differing solubilities. Only 9.6 g of sodium hydrogen carbonate will dissolve in 100 mL of water at 20 °C (293 K), for instance, but 204 g of sucrose will dissolve under the same conditions.

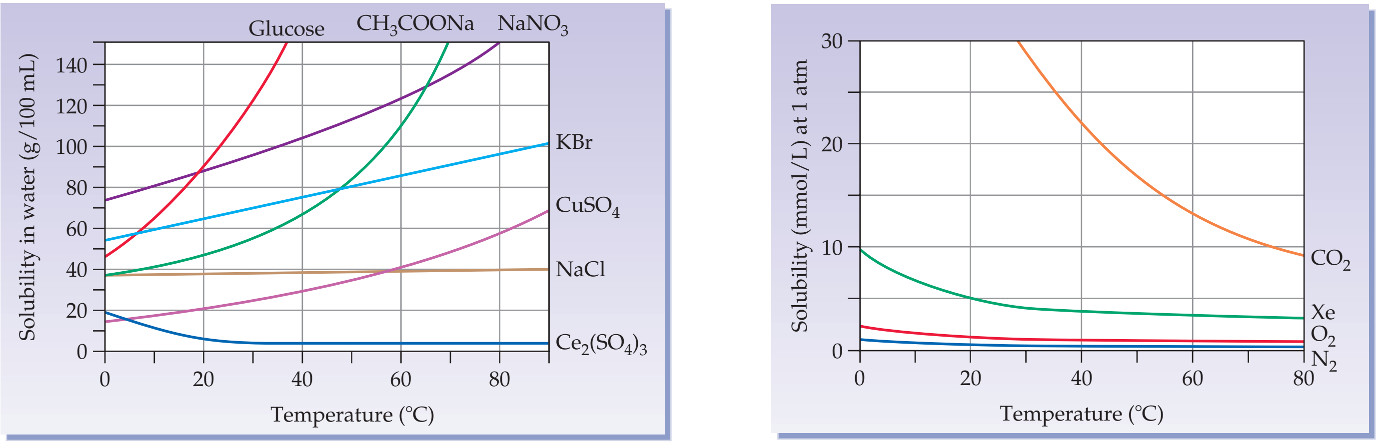
Solubility The maximum amount of a substance that will dissolve in a given amount of solvent at a specified temperature.

9.4 The Effect of Temperature on Solubility

Learning Objective:

- Determine the effect of temperature changes on the solubility of a solute in a solution.

As anyone who has ever made tea or coffee knows, temperature often has a dramatic effect on solubility. The compounds in tea leaves or coffee beans, for instance, dissolve easily in hot water but not in cold water. The effect of temperature is different for every substance, however, and is usually unpredictable. As shown in Figure 9.3a, the solubilities of most molecular and ionic solids increase with increasing temperature, but the



(a) (b)

▲ **Figure 9.3**

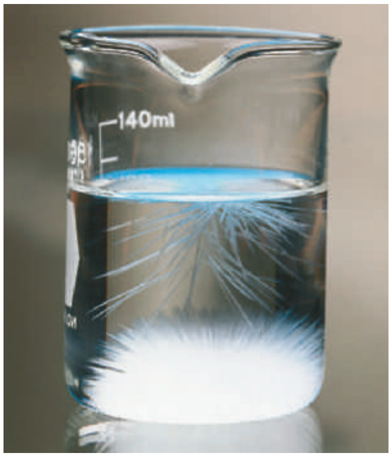
Solubilities of some (a) solids and (b) gases in water as a function of temperature.

Most solid substances become more soluble as temperature rises (although the exact relationship is usually complex), whereas the solubility of gases decreases.

solubilities of others (NaCl) are almost unchanged, and the solubilities of still others $[Ce_2(SO_4)_3]$ decrease with increasing temperature.

Solids that are more soluble at high temperature than at low temperature can sometimes form **supersaturated solutions,** which contain even more solute than a saturated solution. Suppose, for instance, that a large amount of a substance is dissolved at a high temperature. As the solution cools, the solubility decreases and the excess solute should precipitate to maintain equilibrium. But if the cooling is done very slowly, and if the container stands quietly, crystallization might not occur immediately and a supersaturated solution might result. Such a solution is unstable, however, and precipitation can occur dramatically when a tiny seed crystal is added to initiate crystal growth or when the container is disturbed (Figure 9.4).

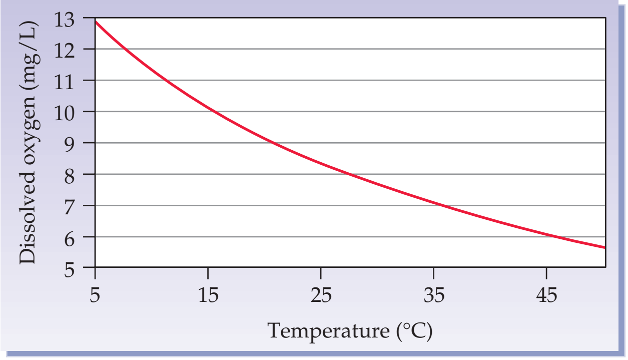
Unlike solids, the influence of temperature on the solubility of gases *is* predictable: Addition of heat decreases the solubility of most gases, as seen in Figure 9.3b (helium is the only common exception). One result of this temperature-dependent decrease in gas solubility can sometimes be noted in a stream or lake near the outflow of warm water from an industrial operation. As water temperature increases, the concentration of dissolved oxygen in the water decreases, killing fish that cannot tolerate the lower oxygen levels.



▲ **Figure 9.4**
A supersaturated solution of sodium acetate in water.
When a tiny seed crystal is added, larger crystals rapidly grow and precipitate from the solution until equilibrium is reached.

Worked Example 9.2 Solubility of Gases: Effect of Temperature

From the following graph of solubility versus temperature for O_2, estimate the concentration of dissolved oxygen in water at 25 °C and at 35 °C. By what percentage does the concentration of O_2 change?



ANALYSIS The solubility of O_2 (on the *y*-axis) can be determined by finding the appropriate temperature (on the *x*-axis) and extrapolating. The percent change is calculated as

$$\frac{(\text{Solubility at } 25\ ^\circ\text{C}) - (\text{Solubility at } 35\ ^\circ\text{C})}{(\text{Solubility at } 25\ ^\circ\text{C})} \times 100$$

SOLUTION

From the graph, we estimate that the solubility of O_2 at 25 °C is approximately 8.3 mg/L and at 35 °C is 7.0 mg/L. The percent change in solubility is

$$\frac{8.3 - 7.0}{8.3} \times 100 = 16$$

Supersaturated solution A solution that contains more than the maximum amount of dissolved solute; a nonequilibrium situation.

PROBLEM 9.3

A solution is prepared by dissolving 12.5 g of KBr in 20 mL of water at 60 °C (see Figure 9.3). Is this solution saturated, unsaturated, or supersaturated? What will happen if the solution is cooled to 10 °C?

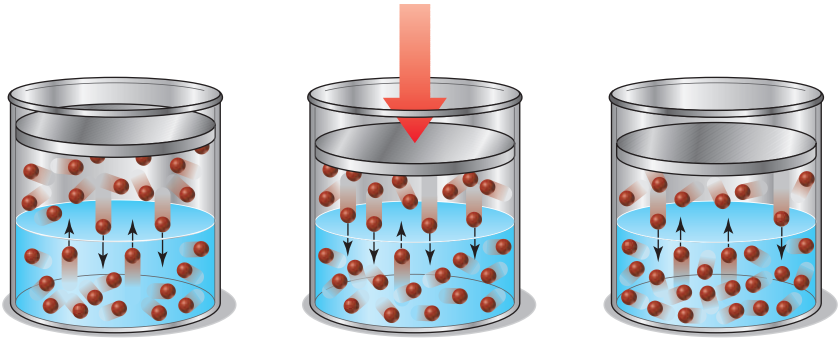
9.5 The Effect of Pressure on Solubility: Henry's Law

Learning Objective:

- Determine the effect of a change in pressure on the solubility of a gas in solution.

Pressure has virtually no effect on the solubility of a solid or liquid, but it has a strong effect on the solubility of a gas. According to **Henry's law,** the solubility (or concentration) of a gas in a liquid is directly proportional to the partial pressure of the gas over the liquid. If the partial pressure of the gas doubles, solubility doubles; if the gas pressure is halved, solubility is halved (Figure 9.5).

◀◀◀ **CONCEPTS TO REVIEW** Recall from Section 8.11 that each gas in a mixture exerts a partial pressure independent of other gases present (Dalton's law of partial pressures).



(a) Equilibrium **(b)** Pressure increase **(c)** Equilibrium restored

◀ **Figure 9.5**
Henry's law.
The solubility of a gas is directly proportional to its partial pressure. An increase in pressure causes more gas molecules to enter solution until equilibrium is restored between the dissolved and undissolved gas.

Henry's law The solubility (or concentration) of a gas is directly proportional to the partial pressure of the gas if the temperature is constant. That is, concentration *(C)* divided by pressure *(P)* is constant when *T* is constant, or

$$\frac{C}{P_{\text{gas}}} = k \quad \text{(At a constant temperature)}$$

◀◀◀ Le Châtelier's principle states that when a system at equilibrium is placed under stress, the equilibrium shifts to relieve that stress (Section 7.9).

Henry's law can be explained using Le Châtelier's principle. In the case of a saturated solution of a gas in a liquid, an equilibrium exists whereby gas molecules enter and leave the solution at the same rate. When the system is stressed by increasing the pressure of the gas, more gas molecules go into solution to relieve that increase. Conversely, when the pressure of the gas is decreased, more gas molecules come out of solution to relieve the decrease.

$$\overset{[\text{Pressure increases} \longrightarrow]}{\text{Gas} + \text{Solvent} \rightleftharpoons \text{Solution}}$$

As an example of Henry's law in action, think about the fizzing that occurs when you open a bottle of soft drink or champagne. The bottle is sealed under greater than 1 atm of CO_2 pressure, causing some of the CO_2 to dissolve. When the bottle is opened, however, CO_2 pressure drops and gas comes fizzing out of solution.

Writing Henry's law in the form $P_{\text{gas}} = C/k$ shows that partial pressure can be used to express the concentration of a gas in a solution, a practice especially common in health-related sciences. Table 9.3 gives some typical values and illustrates the convenience of having the same unit for concentration of a gas in both air and blood. Compare the oxygen partial pressures in saturated alveolar air (air in the lungs) and in arterial blood, for instance. The values are almost the same because the gases dissolved in blood come to equilibrium with the same gases in the lungs.

If the partial pressure of a gas over a solution changes while the temperature is constant, the new solubility of the gas can be found easily. Because C/P is a constant value at constant temperature, Henry's law can be restated to show how one variable changes if the other changes.

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} = k \quad \text{(Where } k \text{ is constant at a fixed temperature)}$$

Worked Example 9.3 gives an illustration of how to use this equation.

Table 9.3 Partial Pressures and Normal Gas Concentrations in Body Fluids

	Partial Pressure (Pa)			
Sample	P_{N_2}	P_{O_2}	P_{CO_2}	P_{H_2O}
Inspired air (dry)	79.6×10^3	21.2×10^3	40.0	493
Alveolar air (saturated)	76.3×10^3	13.3×10^3	5.33×10^3	6.27×10^3
Expired air (saturated)	75.9×10^3	15.5×10^3	3.7×10^3	6.27×10^3
Arterial blood	76.4×10^3	12.7×10^3	5.33×10^3	
Venous blood	76.4×10^3	5.33×10^3	6.0×10^3	
Peripheral tissues	76.4×10^3	5.33×10^3	6.0×10^3	

Worked Example 9.3 Solubility of Gases: Henry's Law

At a partial pressure of oxygen in the atmosphere of 21.2×10^3 Pa, the solubility of oxygen in blood is 0.44 g/100 mL. What is the solubility of oxygen in blood at 7925 m, where the partial pressure of O_2 is 7.5×10^3 Pa?

ANALYSIS According to Henry's law, the solubility of the gas divided by its pressure is constant.

$$\frac{C_1}{P_1} = \frac{C_2}{P_2}$$

Of the four variables in this equation, we know P_1, C_1, and P_2, and we need to find C_2.

BALLPARK ESTIMATE The pressure drops by a factor of about 3 (from 21.2×10^3 Pa to 7.5×10^3 Pa). Since the ratio of solubility to pressure is constant, the solubility must also drop by a factor of 3 (from 0.44 g/100 mL to about 0.15 g/100 mL).

SOLUTION

STEP 1: Identify known information. We have values for P_1, C_1, and P_2.

$$P_1 = 21.2 \times 10^3\ \text{Pa}$$
$$C_1 = 0.44\ \text{g}/100\ \text{mL}$$
$$P_2 = 7.5 \times 10^3\ \text{Pa}$$

STEP 2: Identify answer and units. We are looking for the solubility of O_2 (C_2) at a partial pressure P_2.

Solubility of O_2, $C_2 = ??$ g/100 mL

STEP 3: Identify conversion factors or equations. In this case, we restate Henry's law to solve for C_2.

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} \Rightarrow C_2 = \frac{C_1 P_2}{P_1}$$

STEP 4: Solve. Substitute the known values into the equation and calculate C_2.

$$C_2 = \frac{C_1 P_2}{P_1} = \frac{(0.44\ \text{g}/100\ \text{mL})(7.5 \times 10^3\ \cancel{\text{Pa}})}{21.2 \times 10^3\ \cancel{\text{Pa}}} = 0.16\ \text{g}/100\ \text{mL}$$

BALLPARK CHECK The calculated answer matches our estimate.

PROBLEM 9.4

At 20 °C (293 K) and a partial pressure of 10^5 Pa, the solubility of CO_2 in water is 0.169 g/100 mL. What is the solubility of CO_2 at 33.3×10^5 Pa?

PROBLEM 9.5

At a total atmospheric pressure of 10^5 Pa, the partial pressure of CO_2 in air is approximately 4.0×10^{29} Pa. Using the data in Problem 9.4, what is the solubility of CO_2 in an open bottle of seltzer water at 20 °C (293 K)?

9.6 Units of Concentration

Learning Objective:

- Define units of concentration, and calculate the concentration of a solute in solution.

Although we speak casually of a solution of, say, orange juice as either “dilute” or “concentrated,” laboratory work usually requires an exact knowledge of a solution’s concentration. As indicated in Table 9.4, there are several common methods for expressing concentration. The units differ, but all the methods describe how much solute is present in a given quantity of solution.

Table 9.4 Some Units for Expressing Concentration

Concentration Measure	Solute Measure	Solution Measure
Percent		
Mass/mass percent, (m/m)%	Mass (g)	Mass (g)
Volume/volume percent, (v/v)%	Volume*	Volume*
Mass/volume percent, (m/v)%	Mass (g)	Volume (mL)
Parts per million, ppm	Parts*	10^6 parts*
Parts per billion, ppb	Parts*	10^9 parts*
Molarity, M	Moles	Volume (L)

*Any units can be used as long as they are the same for both solute and solution.

Let us look at each of the concentration measures listed in Table 9.4 individually, beginning with *percent concentrations*.

Percent Concentrations

Percent concentrations express the amount of solute in 100 units of solution. The amount of solute and the amount of solution can be represented in units of mass or volume. For solid solutions, such as a metal alloy, concentrations are typically expressed as **mass/mass percent concentration, (m/m)%.**

mass/mass percent concentration, (m/m)% Concentration expressed as the number of grams of solute per 100 g of solution.

$$\textbf{(m/m)\% concentration} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 100$$

For example, the mass percent of copper in a red-gold ring that contains 19.20 g of gold and 4.80 g of copper would be calculated as

$$\textbf{(m/m)\% Cu} = \frac{\text{mass of Cu (g)}}{\text{mass of Cu (g)} + \text{mass of Au (g)}} \times 100$$
$$= \frac{4.80\text{ g}}{4.80\text{ g} + 19.20\text{ g}} \times 100 = 20.0$$

The concentration of a solution made by dissolving one liquid in another is often given by expressing the volume of solute as a percentage of the volume of final solution—the **volume/volume percent concentration, (v/v)%.**

volume/volume percent concentration, (v/v)% Concentration expressed as the number of milliliters of solute dissolved in 100 mL of solution.

$$\textbf{(v/v)\% concentration} = \frac{\text{Volume of solute (mL)}}{\text{Volume of solution (mL)}} \times 100$$

For example, if 10.0 mL of ethanol is dissolved in enough water to give 100.0 mL of solution, the ethanol concentration is $(10.0\text{ mL}/100.0\text{ mL}) \times 100 = 10.0\%$ (v/v).

A third common method for expressing percent concentration is to give the number of grams (mass) as a percentage of the number of milliliters (volume) of the final solution—called the **mass/volume percent concentration, (m/v)%.**

mass/volume percent concentration, (m/v)% Concentration expressed as the number of grams of solute per 100 mL of solution.