PROBLEM 9.19

Hydrochloric acid is normally purchased at a concentration of 12.0 M. What is the final concentration if 100.0 mL of 12.0 M HCl is diluted to 500.0 mL?

PROBLEM 9.20

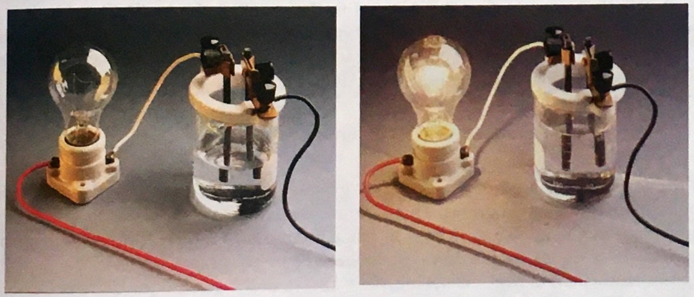
Aqueous ammonia is commercially available at a concentration of 16.0 M. How much of the concentrated solution would you use to prepare 500.0 mL of a 1.25 M solution?

PROBLEM 9.21

The Environmental Protection Agency has set the limit for arsenic in drinking water at 0.010 ppm. To what volume would you need to dilute 1.5 L of water containing 5.0 ppm arsenic to reach the acceptable limit?

9.9 Ions in Solution: Electrolytes

Look at Figure 9.8, which shows a light bulb connected to a power source through a circuit that is interrupted by two metal strips dipped into a beaker of liquid. When the strips are dipped into pure water, the bulb remains dark, but when they are dipped into an aqueous NaCl solution, the circuit is closed and the bulb lights. As mentioned previously in Section 4.1, this simple demonstration shows that ionic compounds in aqueous solution can conduct electricity. (⊂⊃, p. 79)


(a) (b)

▲ **FIGURE 9.8 A simple demonstration shows that electricity can flow through a solution of ions.** (a) With pure water in the beaker, the circuit is incomplete, no electricity flows, and the bulb does not light. (b) With a concentrated NaCl solution in the beaker, the circuit is complete, electricity flows, and the light bulb glows.

Substances like NaCl that conduct an electric current when dissolved in water are called **electrolytes**. Conduction occurs because negatively charged Cl^- anions migrate through the solution toward the metal strip connected to the positive terminal of the power source, whereas positively charged Na^+ cations migrate toward the strip connected to the negative terminal. As you might expect, the ability of a solution to conduct electricity depends on the concentration of ions in solution. Distilled water contains virtually no ions and is nonconducting; ordinary tap water contains low concentrations of dissolved ions (mostly Na^+, K^+, Mg^{2+}, Ca^{2+}, and Cl^-) and is weakly conducting; and a concentrated solution of NaCl is strongly conducting.

Electrolyte A substance that produces ions and therefore conducts electricity when dissolved in water.

Strong electrolyte A substance that ionizes completely when dissolved in water.

Weak electrolyte A substance that is only partly ionized in water.

Nonelectrolyte A substance that does not produce ions when dissolved in water.

Ionic substances like NaCl that ionize completely when dissolved in water are called **strong electrolytes**, and molecular substances like acetic acid (CH_3CO_2H) that are only partially ionized are **weak electrolytes**. Molecular substances like glucose that do not produce ions when dissolved in water are **nonelectrolytes**.

Strong electrolyte; completely ionized	$NaCl(s) \xrightarrow[\text{in water}]{\text{Dissolve}} Na^+(aq) + Cl^-(aq)$
Weak electrolyte; partly ionized	$CH_3CO_2H(l) \underset{\text{in water}}{\overset{\text{Dissolve}}{\rightleftharpoons}} CH_3CO_2^-(aq) + H^+(aq)$
Nonelectrolyte; not ionized	$Glucose(s) \underset{\text{in water}}{\overset{\text{Dissolve}}{\rightleftharpoons}} Glucose(aq)$

9.10 Electrolytes in Body Fluids: Equivalents and Milliequivalents

What happens if NaCl and KBr are dissolved in the same solution? Because the cations (K^+ and Na^+) and anions (Cl^- and Br^-) are all mixed together and no reactions occur between them, an identical solution could just as well be made from KCl and NaBr. Thus, we can no longer speak of having an NaCl + KBr solution; we can only speak of having a solution with four different ions in it.

A similar situation exists for blood and other body fluids, which contain many different anions and cations. Since they are all mixed together, it is difficult to "assign" specific cations to specific anions or to talk about specific ionic compounds. Instead, we are interested only in individual ions and in the total numbers of positive and negative charges. To discuss such mixtures, we use a new term—*equivalents* of ions.

For ions, one **equivalent (Eq)** is equal to the number of ions that carry 1 mol of charge. Of more practical use is the unit **gram-equivalent**, which is the mass of the ion that contains one mole of charge. It can be calculated simply as the molar mass of the ion divided by the absolute value of its charge.

Equivalent For ions, the amount equal to 1 mol of charge.

Gram-equivalent For ions, the molar mass of the ion divided by the ionic charge.

$$\text{One gram-equivalent of ion} = \frac{\text{Molar mass of ion (g)}}{\text{Charge on ion}}$$

If the ion has a charge of +1 or −1, 1 gram-equivalent of the ion is simply the molar mass of the ion in grams. Thus, 1 gram-equivalent of Na^+ is 23 g, and 1 gram-equivalent of Cl^- is 35.5 g. If the ion has a charge of +2 or −2, however, 1 gram-equivalent is equal to the ion's formula weight in grams divided by 2. Thus, 1 gram-equivalent of Mg^{2+} is $(24.3\ g)/2 = 12.2\ g$, and 1 gram-equivalent of CO_3^{2-} is $[12.0\ g + (3 \times 16.0\ g)]/2 = 30.0\ g$. The gram-equivalent is a useful conversion factor when converting from volume of solution to mass of ions, as seen in Worked Example 9.14.

The number of equivalents of a given ion per liter of solution can be found by multiplying the molarity of the ion (moles per liter) by the charge on the ion. Because ion concentrations in body fluids are often low, clinical chemists find it more convenient to talk about *milliequivalents* of ions rather than equivalents. One milliequivalent (mEq) of an ion is 1/1000 of an equivalent. For example, the normal concentration of Na^+ in blood is 0.14 Eq/L, or 140 mEq/L.

$$1\ mEq = 0.001\ Eq \qquad 1\ Eq = 1000\ mEq$$

Note that the gram-equivalent for an ion can now be expressed as grams per equivalent or as mg per mEq.

Average concentrations of the major electrolytes in blood plasma are given in Table 9.6. As you might expect, the total milliequivalents of positively and negatively charged electrolytes must be equal to maintain electrical neutrality. Adding

TABLE 9.6 Concentrations of Major Electrolytes in Blood Plasma

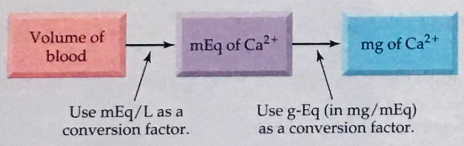
CATION	CONCENTRATION (mEq/L)
Na^+	136–145
Ca^{2+}	4.5–6.0
K^+	3.6–5.0
Mg^{2+}	3
ANION	**CONCENTRATION (mEq/L)**
Cl^-	98–106
HCO_3^-	25–29
SO_4^{2-} and HPO_4^{2-}	2

the milliequivalents of positive and negative ions in Table 9.6, however, shows a higher concentration of positive ions than negative ions. The difference, called the *anion gap*, is made up by the presence of negatively charged proteins and the anions of organic acids.

WORKED EXAMPLE **9.14** Equivalents as Conversion Factors: Volume to Mass

The normal concentration of Ca^{2+} in blood is 5.0 mEq/L. How many milligrams of Ca^{2+} are in 1.00 L of blood?

ANALYSIS We are given a volume and a concentration in milliequivalents per liter, and we need to find an amount in milligrams. Thus, we need to calculate the gram-equivalent for Ca^{2+} and then use concentration as a conversion factor between volume and mass, as indicated in the following flow diagram:



BALLPARK ESTIMATE The molar mass of calcium is 40.08 g/mol, and the calcium ion carries a charge of 2+. Thus, 1 g-Eq of Ca^{2+} equals about 20 g/Eq or 20 mg/mEq. This means that the 5.0 mEq of Ca^{2+} ions in 1.00 L of blood corresponds to a mass of 5.0 mEq Ca^{2+} × 20 mg/mEq = 100 mg Ca^{2+}.

SOLUTION

$$(1.00\ \cancel{\text{L blood}})\left(\frac{5.0\ \cancel{\text{mEq Ca}^{2+}}}{1.0\ \cancel{\text{L blood}}}\right)\left(\frac{20.04\ \text{mg Ca}^{2+}}{1\ \cancel{\text{mEq Ca}^{2+}}}\right) = 100\ \text{mg Ca}^{2+}$$

BALLPARK CHECK: The calculated answer (100 mg of Ca^{2+} in 1.00 L of blood) matches our estimate.

PROBLEM 9.22

How many grams are in 1 Eq of the following ions? How many grams in 1 mEq?

(a) K^+ (b) Br^- (c) Mg^{2+} (d) SO_4^{2-} (e) Al^{3+} (f) PO_4^{3-}

PROBLEM 9.23

Look at the data in Table 9.6, and calculate how many milligrams of Mg^{2+} are in 250 mL of blood.

9.11 Properties of Solutions

The properties of solutions are similar in many respects to those of pure solvents, but there are also some interesting and important differences. One such difference is that solutions have higher boiling points than the pure solvents; another is that solutions have lower freezing points. Pure water boils at 100.0 °C and freezes at 0.0 °C, for example, but a 1.0 M solution of NaCl in water boils at 101.0 °C and freezes at −3.7 °C.

Colligative property A property of a solution that depends only on the number of dissolved particles, not on their chemical identity.

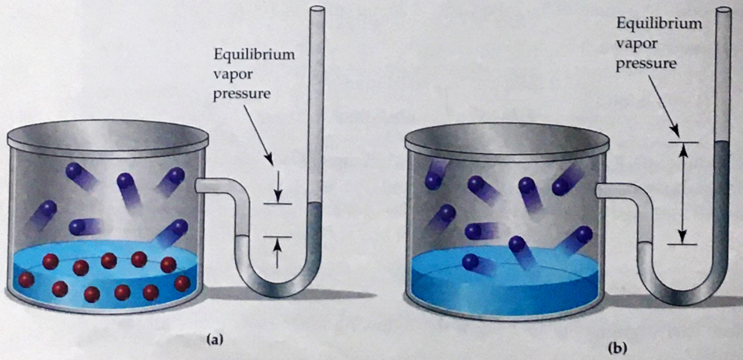
The elevation of boiling point and the lowering of freezing point for a solution as compared with a pure solvent are examples of **colligative properties**—properties that depend on the *concentration* of a dissolved solute but not on its chemical identity. Other colligative properties are a lower vapor pressure for a solution compared with the pure solvent and *osmosis*, the migration of solvent molecules through a semipermeable membrane.

Colligative properties

- Vapor pressure is lower for a solution than for a pure solvent.
- Boiling point is higher for a solution than for a pure solvent.
- Freezing point is lower for a solution than for a pure solvent.
- Osmosis occurs when a solution is separated from a pure solvent by a semipermeable membrane.

Vapor Pressure Lowering in Solutions

We said in Section 8.12 that the vapor pressure of a liquid depends on the equilibrium between molecules entering and leaving the liquid surface. (🔗, p. 241) Only those molecules at the surface of the liquid that are sufficiently energetic will evaporate. If, however, some of the liquid (solvent) molecules at the surface are replaced by other (solute) particles that do not evaporate, then the rate of evaporation of solvent molecules decreases and the vapor pressure of a solution is lower than that of the pure solvent (Figure 9.9). Note that the *identity* of the solute particles is irrelevant; only their concentration matters.



▲ **FIGURE 9.9 Vapor pressure lowering of solution.** (a) The vapor pressure of a solution is lower than (b) the vapor pressure of the pure solvent because fewer solvent molecules are able to escape from the surface of the solution.

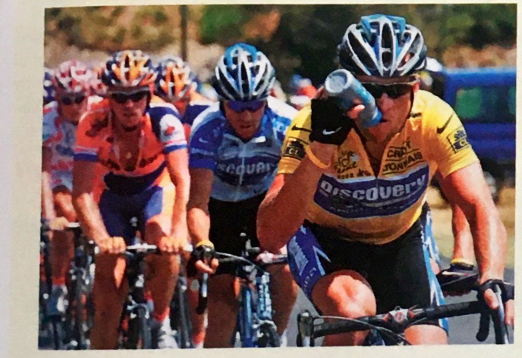
APPLICATION ▶ Electrolytes, Fluid Replacement, and Sports Drinks

Athletes sweat. And the hotter the day, the more intense and longer lasting the activity, the more they sweat. Sweat loss during strenuous exercise on a hot day can amount to as much as 2 L/h, and the total sweat loss during a 24 h endurance run can exceed 16 L, or approximately 35 lb.

The composition of sweat is highly variable, not only with the individual but also with the time during the exercise and the athlete's overall conditioning. Typically, however, the Na^+ ion concentration in sweat is about 30–40 mEq/L, and that of K^+ ion is about 5–10 mEq/L. In addition, there are small amounts of other metal ions, such as Mg^{2+}, and there are sufficient Cl^- ions (35–50 mEq/L) to balance the positive charge of all these cations.

Obviously, all the water and dissolved electrolytes lost by an athlete through sweating must be replaced. Otherwise, dehydration, hyperthermia and heat stroke, dizziness, nausea, muscle cramps, impaired kidney function, and other difficulties ensue. As a rule of thumb, a sweat loss equal to 5% of body weight—about 3.5 L for a 150 lb person—is the maximum amount that can be safely allowed for a well-conditioned athlete.

Plain water works perfectly well to replace sweat lost during short bouts of activity up to a few hours in length, but a carbohydrate–electrolyte beverage, or "sports drink," is much superior for rehydrating during and after longer activity in which substantial amounts of electrolytes have been lost. Some of the better known sports drinks are little more than overpriced sugar–water solutions, but others are carefully formulated and highly effective for fluid replacement. Nutritional research has shown that a serious sports drink should meet the following criteria. There are several dry-powder mixes on the market to choose from.



▲ Endurance athletes avoid serious medical problems by drinking sports drinks to replace water and electrolytes.

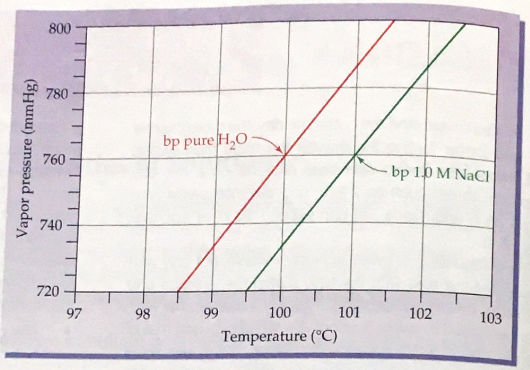
- The drink should contain 6–8% of soluble complex carbohydrates (about 15 g per 8 oz serving) and only a small amount of simple sugar for taste. The complex carbohydrates, which usually go by the name "maltodextrin," provide a slow release of glucose into the bloodstream. Not only does the glucose provide a steady source of energy, it also enhances the absorption of water from the stomach.
- The drink should contain electrolytes to replenish those lost in sweat. Concentrations of approximately 20 mEq/L for Na^+ ions, 10 mEq/L for K^+ ion, and 4 mEq/L for Mg^{2+} ions are recommended. These amounts correspond to about 100 mg sodium, 100 mg potassium, and 25 mg magnesium per 8 oz serving.
- The drink should be noncarbonated because carbonation can cause gastrointestinal upset during exercise, and it should not contain caffeine, which acts as a diuretic.
- The drink should taste good so the athlete will want to drink it. Thirst is a poor indicator of fluid requirements, and most people will drink less than needed unless a beverage is flavored.

In addition to complex carbohydrates, electrolytes, and flavorings, some sports drinks also contain vitamin A (as beta-carotene), vitamin C (ascorbic acid), and selenium, which act as antioxidants to protect cells from damage. Some drinks also contain the amino acid glutamine, which appears to lessen lactic acid buildup in muscles and thus helps muscles bounce back more quickly after an intense workout.

See Additional Problems 9.93 and 9.94 at the end of the chapter.

Boiling Point Elevation of Solutions

One consequence of the vapor pressure lowering for a solution is that the boiling point of the solution is higher than that of the pure solvent. Recall from Section 8.12 that boiling occurs when the vapor pressure of a liquid reaches atmospheric pressure. (, p. 241) But because the vapor pressure of a solution is lower than that of the pure solvent at a given temperature, the solution must be heated to a higher temperature for its vapor pressure to reach atmospheric pressure. Figure 9.10 shows a close-up plot of vapor pressure versus temperature for pure water and for a 1.0 M NaCl solution. The vapor pressure of pure water reaches atmospheric pressure (760 mmHg) at 100.0 °C, but the vapor pressure of the NaCl solution does not reach the same point until 101.0 °C.



▲ **FIGURE 9.10 Vapor pressure and temperature.** A close-up plot of vapor pressure versus temperature for pure water (red curve) and for a 1.0 M NaCl solution (green curve). Pure water boils at 100.0 °C, but the solution does not boil until 101.0 °C.

For each mole of solute particles added, regardless of chemical identity, the boiling point of 1 kg of water is raised by 0.51 °C, or

$$\Delta T_{\text{boiling}} = \left(0.51\ ^\circ\text{C}\,\frac{\text{kg water}}{\text{mol particles}}\right)\left(\frac{\text{mol particles}}{\text{kg water}}\right)$$

The addition of 1 mol of a molecular substance like glucose to 1 kg of water therefore raises the boiling point from 100.0 °C to 100.51 °C. The addition of 1 mol of NaCl per kilogram of water, however, raises the boiling point by 2 × 0.51 °C = 1.02 °C because the solution contains 2 mol of solute particles—Na^+ and Cl^- ions.

WORKED EXAMPLE 9.15 Properties of Solutions: Boiling Point Elevation

What is the boiling point of a solution of 0.75 mol of KBr in 1.0 kg of water?

ANALYSIS The boiling point increases 0.51 °C for each mole of solute per kilogram of water. Since KBr is a strong electrolyte, there are 2 mol of ions (K^+ and Br^-) for every 1 mol of KBr that dissolves.

BALLPARK ESTIMATE The boiling point will increase about 0.5 °C for every 1 mol of ions in 1 kg of water. Since 0.75 mol of KBr produce 1.5 mol of ions, the boiling point should increase by (1.5 mol ions) × 0.5 °C/mol ions) = 0.75 °C

SOLUTION

$$\Delta T_{\text{boiling}} = \left(0.51\ ^\circ\text{C}\,\frac{\cancel{\text{kg water}}}{\cancel{\text{mol ions}}}\right)\left(\frac{2\ \cancel{\text{mol ions}}}{1\ \cancel{\text{mol KBr}}}\right)\left(\frac{0.75\ \cancel{\text{mol KBr}}}{1.0\ \cancel{\text{kg water}}}\right) = 0.77\ ^\circ\text{C}$$

The normal boiling point of pure water is 100 °C, so the boiling point of the solution increases to 100.77 °C.

BALLPARK CHECK: The 0.77 °C increase is consistent with our estimate of 0.75 °C.

PROBLEM 9.24

What is the boiling point of a solution of 0.67 mol of $MgCl_2$ in 0.50 kg of water?

PROBLEM 9.25

When 1.0 mol of HF is dissolved in 1.0 kg of water, the boiling point of the resulting solution is 100.5 °C. Is HF a strong or weak electrolyte? Explain.

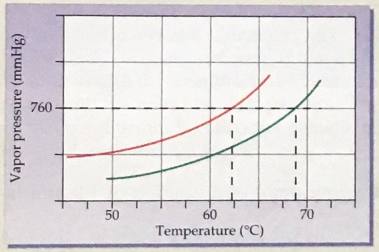
KEY CONCEPT PROBLEM 9.26

The following diagram shows plots of vapor pressure versus temperature for a solvent and a solution.

(a) Which curve represents the pure solvent, and which the solution?

(b) What are the approximate boiling points of the pure solvent and the solution?

(c) What is the approximate concentration of the solution in mol/kg, if 1 mol of solute particles raises the boiling point of 1 kg of solvent by 3.63 °C?



Freezing Point Depression of Solutions

Just as solutions have lower vapor pressure and consequently higher boiling points than pure solvents, they also have lower freezing points. Motorists in cold climates take advantage of this effect when they add “antifreeze” to the water in automobile cooling systems. Antifreeze is a nonvolatile solute, usually ethylene glycol ($HOCH_2CH_2OH$), that is added in sufficient concentration to lower the freezing point below the lowest expected outdoor temperature. In the same way, salt sprinkled on icy roads lowers the freezing point of ice below the road temperature and thus causes ice to melt.

Freezing point depression has much the same cause as vapor pressure lowering. Solute molecules are dispersed between solvent molecules throughout the solution, thereby making it more difficult for solvent molecules to come together and organize into ordered crystals.

For each mole of nonvolatile solute particles, the freezing point of 1 kg of water is lowered by 1.86 °C, or

$$\Delta T_{\text{freezing}} = \left(-1.86\ ^{\circ}\text{C}\,\frac{\text{kg water}}{\text{mol particles}}\right)\left(\frac{\text{mol particles}}{\text{kg water}}\right)$$

Thus, addition of 1 mol of antifreeze to 1 kg of water lowers the freezing point from 0.00 °C to −1.86 °C, and addition of 1 mol of NaCl (2 mol of particles) to 1 kg of water lowers the freezing point from 0.00 °C to −3.72 °C.

▲ A mixture of salt and ice is used to provide the low temperatures needed to make old-fashioned hand-cranked ice cream.

WORKED EXAMPLE **9.16** Properties of Solutions: Freezing Point Depression

The cells of a tomato contain mostly an aqueous solution of sugar and other substances. If a typical tomato freezes at −2.5 °C, what is the concentration of dissolved particles in the tomato cells (in moles of particles per kg of water)?

ANALYSIS The freezing point decreases by 1.86 °C for each mole of solute dissolved in 1 kg of water. We can use the decrease in freezing point (2.5 °C) to find the amount of solute per kg of water.

BALLPARK ESTIMATE The freezing point will decrease by about 1.9 °C for every 1 mol of solute particles in 1 kg of water. To lower the freezing point by 2.5 °C therefore requires about 1.5 mol of particles per kg of water.

SOLUTION

$$\begin{aligned}\Delta T_{\text{freezing}} &= -2.5\ ^{\circ}\text{C}\\ &= \left(-1.86\ ^{\circ}\text{C}\,\frac{\cancel{\text{kg water}}}{\cancel{\text{mol solute particles}}}\right)\left(\frac{??\ \cancel{\text{mol solute particles}}}{1.0\ \cancel{\text{kg water}}}\right)\end{aligned}$$

We can rearrange this expression to

$$(-2.5\ \cancel{^{\circ}\text{C}})\left(\frac{1}{-1.86\ \cancel{^{\circ}\text{C}}}\ \frac{\text{mol solute particles}}{\text{kg water}}\right) = 1.3\ \frac{\text{mol solute particles}}{\text{kg water}}$$

BALLPARK CHECK: The calculated answer is relatively close to our estimate (1.5 mol/kg).

PROBLEM 9.27

What is the freezing point of a solution of 1.0 mol of glucose in 1.0 kg of water?

PROBLEM 9.28

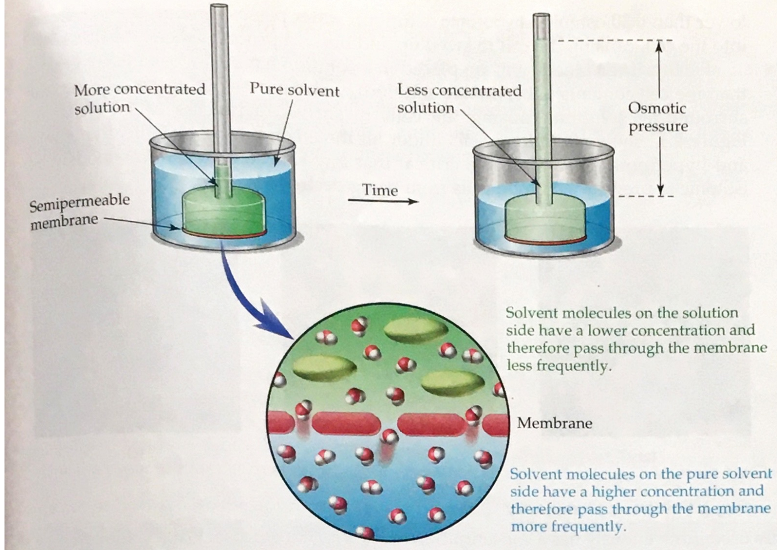
When 0.5 mol of a certain ionic substance is dissolved in 1.0 kg of water, the freezing point of the resulting solution is −2.8 °C. How many ions does the substance give when it dissolves?

9.12 Osmosis and Osmotic Pressure

Osmosis The passage of solvent through a semipermeable membrane separating two solutions of different concentration.

Certain materials, including those that make up the membranes around living cells, are *semipermeable*. They allow water and other small molecules to pass through, but they block the passage of large solute molecules or ions. When a solution and a pure solvent, or two solutions of different concentration, are separated by a semipermeable membrane, solvent molecules pass through the membrane in a process called **osmosis**. Although the passage of solvent through the membrane takes place in both directions, passage from the pure solvent side to the solution side is favored and occurs more often. As a result, the amount of liquid on the pure solvent side decreases, the amount of liquid on the solution side increases, and the concentration of the solution decreases.

For the simplest explanation of osmosis, let us look at what happens on the molecular level. As shown in Figure 9.11, a solution inside a bulb is separated by a semipermeable membrane from pure solvent in the outer container. Solvent molecules in the outer container, because of their somewhat higher concentration, approach the membrane more frequently than do molecules in the bulb, thereby passing through more often and causing the liquid level in the attached tube to rise.



◀ **FIGURE 9.11 The phenomenon of osmosis.** A solution inside the bulb is separated from pure solvent in the outer container by a semipermeable membrane. Solvent molecules in the container have a slightly higher concentration than molecules in the bulb and therefore pass through the membrane more frequently. The liquid in the tube therefore rises until an equilibrium is reached. At equilibrium, the osmotic pressure exerted by the column of liquid in the tube is sufficient to prevent further net passage of solvent.

As the liquid in the tube rises, its increased weight creates an increased pressure that pushes solvent back through the membrane until the rates of forward and reverse passage become equal and the liquid level stops rising. The amount of pressure necessary to achieve this equilibrium is called the **osmotic pressure** (π) of the solution and can be determined from the expression

$$\pi = \left(\frac{n}{V}\right)RT$$

Osmotic pressure The amount of external pressure applied to the more concentrated solution to halt the passage of solvent molecules across a semipermeable membrane.

where n is the number of moles of particles in the solution, V is the solution volume, R is the gas constant (🔗, p. 231), and T is the absolute temperature of the solution. Note the similarity between this equation for the osmotic pressure of a solution and the equation for the pressure of an ideal gas, $P = (n/V)RT$. In both cases, the pressure has units of atmospheres.

Osmotic pressures can be extremely high, even for relatively dilute solutions. The osmotic pressure of a 0.15 M NaCl solution at 25 °C, for example, is 7.3 atm, a value that supports a difference in water level of approximately 250 ft!

As with other colligative properties, the amount of osmotic pressure depends only on the concentration of solute particles, not on their identity. Thus, it is convenient to use a new unit, *osmolarity* (osmol), to describe the concentration of particles in solution. The **osmolarity** of a solution is equal to the number of moles of dissolved particles (ions or molecules) per liter of solution. A 0.2 M glucose solution, for instance, has an osmolarity of 0.2 osmol, but a 0.2 M solution of NaCl has an osmolarity of 0.4 osmol because it contains 0.2 mol of Na^+ ions and 0.2 mol of Cl^- ions.

Osmolarity (osmol) The sum of the molarities of all dissolved particles in a solution.

Osmosis is particularly important in living organisms because the membranes around cells are semipermeable. The fluids both inside and outside cells must therefore have the same osmolarity to prevent buildup of osmotic pressure and consequent rupture of the cell membrane.

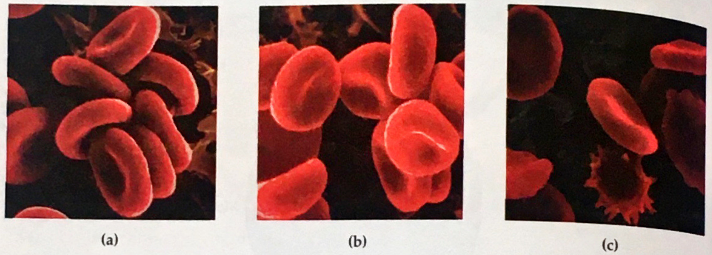
In blood, the plasma surrounding red blood cells has an osmolarity of approximately 0.30 osmol and is said to be **isotonic** with (that is, has the same osmolarity as) the cell contents. If the cells are removed from plasma and placed in 0.15 M NaCl (called *physiological saline solution*), they are unharmed because the osmolarity of the saline solution (0.30 osmol) is the same as that of plasma. If, however, red blood cells are placed in pure water or in any solution with an osmolarity much lower than 0.30 osmol (a **hypotonic** solution), water passes through the membrane into the cell, causing the cell to swell up and burst, a process called *hemolysis*.

Isotonic Having the same osmolarity.

Hypotonic Having an osmolarity *less than* the surrounding blood plasma or cells.

Hypertonic Having an osmolarity *greater than* the surrounding blood plasma or cells.

Finally, if red blood cells are placed in a solution having an osmolarity greater than the cell contents (a **hypertonic** solution), water passes out of the cells into the surrounding solution, causing the cells to shrivel, a process called *crenation*. Figure 9.12 shows red blood cells under all three conditions: isotonic, hypotonic, and hypertonic. Therefore, it is critical that any solution used intravenously be isotonic to prevent red blood cells from being destroyed.



▶ **FIGURE 9.12 Red blood cells.** In (a) an isotonic solution the blood cells are normal in appearance, but the cells in (b) a hypotonic solution are swollen because of water gain, and those in (c) a hypertonic solution are shriveled because of water loss.

WORKED EXAMPLE 9.17 Properties of Solutions: Osmolarity

The solution of glucose commonly used intravenously has a concentration of 5.0% (w/v) glucose. What is the osmolarity of this solution? The molecular weight of glucose is 180 amu.

ANALYSIS Since glucose is a molecular substance that does not give ions in solution, the osmolarity of the solution is the same as the molarity. Recall from Section 9.7 that a solution of 5.0% (w/v) glucose has a concentration of 5.0 g glucose per 100 mL of solution, which is equivalent to 50 g per liter of solution. Thus, finding the molar concentration of glucose requires a mass to mole conversion.

BALLPARK ESTIMATE One liter of solution contains 50 g of glucose (MW= 180 g/mol). Thus, 50 g of glucose is equal to a little more than 0.25 mol, so a solution concentration of 50 g/L is equal to about 0.25 osmol, or 0.25 M.

SOLUTION

STEP 1: Identify known information. We know the (w/v)% concentration of the glucose solution.

$$5.0\%\ (\text{w/v}) = \frac{5.0\text{ g glucose}}{100\text{ mL solution}} \times 100\%$$

STEP 2: Identify answer and units. We are looking for osmolarity, which in this case is equal to the molarity of the solution because glucose is a molecular substance and does not dissociate into ions.

$$\text{Osmolarity} = \text{Molarity} = ??\ \text{mol/liter}$$

STEP 3: Identify conversion factors. The (w/v)% concentration is defined as grams of solute per 100 mL of solution, and molarity is defined as moles of solute per liter of solution. We will need to convert from milliliters to liters and then use molar mass to convert grams of glucose to moles of glucose.

$$\frac{\text{g glucose}}{100\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{\text{L}} = \frac{\text{g glucose}}{\text{L}}$$

$$\frac{\cancel{\text{g glucose}}}{\text{L}} \times \frac{1\text{ mol glucose}}{180\ \cancel{\text{g glucose}}} = \frac{\text{moles glucose}}{\text{L}}$$

STEP 4: Solve. Starting with the (w/v)% glucose concentration, we first find the number of grams of glucose in 1 L of solution and then convert to moles of glucose per liter.

$$\left(\frac{5.0\text{ g glucose}}{100\ \cancel{\text{mL}}\text{ solution}}\right)\left(\frac{1000\ \cancel{\text{mL}}}{1\text{ L}}\right) = 50\,\frac{\text{g glucose}}{\text{L solution}}$$

$$\left(\frac{50\ \cancel{\text{g}}\text{ glucose}}{1\text{ L}}\right)\left(\frac{1\text{ mol}}{180\ \cancel{\text{g}}}\right) = 0.28\text{ M glucose} = 0.28\text{ osmol}$$

BALLPARK CHECK: The calculated osmolarity is reasonably close to our estimate of 0.25 osmol

9.58 If you had only 23 g of KOH remaining in a bottle, how many milliliters of 10.0% (w/v) solution could you prepare? How many milliliters of 0.25 M solution?

9.59 Over-the-counter hydrogen peroxide solutions are 3% (w/v). What is this concentration in moles per liter?

9.60 The lethal dosage of potassium cyanide in rats is 10 mg KCN per kilogram of body weight. What is this concentration in parts per million?

9.61 The maximum concentration set by the U.S. Environmental Protection Agency for lead in drinking water is 15 ppb. What is this concentration in milligrams per liter? How many liters of water contaminated at this maximum level must you drink to consume 1.0 µg of lead?

9.62 What is the molarity of the following solutions?

(a) 12.5 g $NaHCO_3$ in 350.0 mL solution
(b) 45.0 g H_2SO_4 in 300.0 mL solution
(c) 30.0 g NaCl dissolved to make 500.0 mL solution

9.63 How many moles of solute are in the following solutions?

(a) 200 mL of 0.30 M acetic acid, CH_3CO_2H
(b) 1.50 L of 0.25 M NaOH
(c) 750 mL of 2.5 M nitric acid, HNO_3

9.64 How many milliliters of a 0.75 M HCl solution do you need to obtain 0.0040 mol of HCl?

9.65 Nalorphine, a relative of morphine, is used to combat withdrawal symptoms in heroin users. How many milliliters of a 0.40% (w/v) solution of nalorphine must be injected to obtain a dose of 1.5 mg?

9.66 A flask containing 450 mL of 0.50 M H_2SO_4 was accidentally knocked to the floor. How many grams of $NaHCO_3$ do you need to put on the spill to neutralize the acid according to the following equation?

$$\begin{gathered} H_2SO_4(aq) + 2\,NaHCO_3(aq) \longrightarrow \\ Na_2SO_4(aq) + 2\,H_2O(l) + 2\,CO_2(g) \end{gathered}$$

9.67 How many milliliters of 0.0200 M $Na_2S_2O_3$ solution are needed to dissolve 0.450 g of AgBr?

$$\begin{gathered} AgBr(s) + 2\,Na_2S_2O_3(aq) \longrightarrow \\ Na_3Ag(S_2O_3)_2(aq) + NaBr(aq) \end{gathered}$$

9.68 How much water must you add to 100.0 mL of orange juice concentrate if you want the final juice to be 20.0% of the strength of the original?

9.69 How much water would you add to 100.0 mL of 0.500 M NaOH if you wanted the final concentration to be 0.150 M?

9.70 An aqueous solution that contains 285 ppm of potassium nitrate (KNO_3) is being used to feed plants in a garden. What volume of this solution is needed to prepare 2.0 L of a solution that is 75 ppm in KNO_3?

9.71 What is the concentration of a NaCl solution, in (w/v)%, prepared by diluting 65 mL of a saturated solution, which has a concentration of 37 (w/v)%, to 480 mL?

9.72 Concentrated (12.0 M) hydrochloric acid is sold for household and industrial purposes under the name "muriatic acid." How many milliliters of 0.500 M HCl solution can be made from 25.0 mL of 12.0 M HCl solution?

9.73 Dilute solutions of $NaHCO_3$ are sometimes used in treating acid burns. How many milliliters of 0.100 M $NaHCO_3$ solution are needed to prepare 750.0 mL of 0.0500 M $NaHCO_3$ solution?

ELECTROLYTES

9.74 What is an electrolyte?

9.75 Give an example of a strong electrolyte and a nonelectrolyte.

9.76 What does it mean when we say that the concentration of Ca^{2+} in blood is 3.0 mEq/L?

9.77 What is the total anion concentration (in mEq/L) of a solution that contains 5.0 mEq/L Na^+, 12.0 mEq/L Ca^{2+}, and 2.0 mEq/L Li^+?

9.78 Kaochlor, a 10% (w/v) KCl solution, is an oral electrolyte supplement administered for potassium deficiency. How many milliequivalents of K^+ are in a 30 mL dose?

9.79 Calculate the gram-equivalent for each of the following ions:

(a) Ca^{2+}
(b) K^+
(c) $SO_4^{\,2-}$
(d) $PO_4^{\,3-}$

9.80 Look up the concentration of Cl^- ion in blood in Table 9.6. How many milliliters of blood would be needed to obtain 1.0 g of Cl^- ions?

9.81 Normal blood contains 3 mEq/L of Mg^{2+}. How many milligrams of Mg^{2+} are present in 150.0 mL of blood?

PROPERTIES OF SOLUTIONS

9.82 Which lowers the freezing point of 2.0 kg of water more, 0.20 mol NaOH or 0.20 mol $Ba(OH)_2$? Both compounds are strong electrolytes. Explain.

9.83 Which solution has the higher boiling point, 0.500 M glucose or 0.300 M KCl? Explain.

9.84 Methanol, CH_3OH, is sometimes used as an antifreeze for the water in automobile windshield washer fluids. How many grams of methanol must be added to 5.00 kg of water to lower its freezing point to −10.0 °C? For each mole of solute, the freezing point of 1 kg of water is lowered 1.86 °C.

9.85 Hard candy is prepared by dissolving pure sugar and flavoring in water and heating the solution to boiling. What is the boiling point of a solution produced by adding 650 g of cane sugar (molar mass 342.3 g/mol) to 1.5 kg of water? For each mole of nonvolatile solute, the boiling point of 1 kg of water is raised 0.51 °C.

OSMOSIS

9.86 Why do red blood cells swell up and burst when placed in pure water?

9.87 What does it mean when we say that a 0.15 M NaCl solution is isotonic with blood, whereas distilled water is hypotonic?

9.88 Which of the following solutions has the higher osmolarity?

(a) 0.25 M KBr or 0.20 M Na_2SO_4
(b) 0.30 M NaOH or 3.0% (w/v) NaOH

9.89 Which of the following solutions will give rise to a greater osmotic pressure at equilibrium: 5.00 g of NaCl in 350.0 mL water or 35.0 g of glucose in 400.0 mL water? For NaCl, MW = 58.5 amu; for glucose, MW = 180 amu.

9.90 A pickling solution is prepared by dissolving 270 g of NaCl in 3.8 L of water. Calculate the osmolarity of the solution.

9.91 An isotonic solution must be approximately 0.30 osmol. How much KCl is needed to prepare 175 mL of an isotonic solution?

Applications

9.92 How does the body increase oxygen availability at high altitude? [*Breathing and Oxygen Transport, p. 266*]

9.93 What are the major electrolytes in sweat, and what are their approximate concentrations in mEq/L? [*Electrolytes, Fluid Replacement, and Sports Drinks, p. 279*]

9.94 Why is a sports drink more effective than plain water for rehydration after extended exercise? [*Electrolytes, Fluid Replacement, and Sports Drinks, p. 279*]

9.95 How does an enteric coating on a medication work? [*Timed-Release Medications, p. 286*]

General Questions and Problems

9.96 Hyperbaric chambers, which provide high pressures (up to 6 atm) of either air or pure oxygen, are used to treat a variety of conditions, ranging from decompression sickness in deep-sea divers to carbon monoxide poisoning.

(a) What is the partial pressure of O_2 (in millimeters of Hg) in a hyperbaric chamber pressurized to 5 atm with air that is 18% in O_2?
(b) What is the solubility of O_2 (in grams per 100 mL) in the blood at this partial pressure? The solubility of O_2 is 2.1 g/100 mL for $P_{O_2} = 1$ atm.

9.97 Express the solubility of O_2 in Problem 9.96(b) in units of molarity.

9.98 Uric acid, the principal constituent of some kidney stones, has the formula $C_5H_4N_4O_3$. In aqueous solution, the solubility of uric acid is only 0.067 g/L. Express this concentration in (w/v)%, in parts per million, and in molarity.

9.99 Emergency treatment of cardiac arrest victims sometimes involves injection of a calcium chloride solution directly into the heart muscle. How many grams of $CaCl_2$ are administered in an injection of 5.0 mL of a 5.0% (w/v) solution? How many milliequivalents of Ca^{2+}?

9.100 Nitric acid, HNO_3, is available commercially at a concentration of 16 M. What volume would you use to prepare 750 mL of a 0.20 M solution?

9.101 One test for vitamin C (ascorbic acid, $C_6H_8O_6$) is based on the reaction of the vitamin with iodine:

$$C_6H_8O_6(aq) + I_2(aq) \longrightarrow C_6H_6O_6(aq) + 2\,HI(aq)$$

(a) If 25.0 mL of a fruit juice requires 13.0 mL of 0.0100 M I_2 solution for reaction, what is the molarity of the ascorbic acid in the fruit juice?
(b) The Food and Drug Administration recommends that 60 mg of ascorbic acid be consumed per day. How many milliliters of the fruit juice in part (a) must a person drink to obtain the recommended dosage?

9.102 *Ringer's solution*, used in the treatment of burns and wounds, is prepared by dissolving 8.6 g of NaCl, 0.30 g of KCl, and 0.33 g of $CaCl_2$ in water and diluting to a volume of 1.00 L. What is the molarity of each component?

9.103 What is the osmolarity of Ringer's solution (see Problem 9.102)? Is it hypotonic, isotonic, or hypertonic with blood plasma (0.30 osmol)?

9.104 The typical dosage of statin drugs for the treatment of high cholesterol is 10 mg. Assuming a total blood volume of 5.0 L, calculate the concentration of drug in the blood in units of (w/v)%.

9.105 Assuming the density of blood in healthy individuals is approximately 1.05 g/mL, report the concentration of drug in Problem 9.104 in units of ppm.

9.106 In many states, a person with a blood alcohol concentration of 0.080% (v/v) is considered legally drunk. What volume of total alcohol does this concentration represent, assuming a blood volume of 5.0 L?

9.107 Ammonia is very soluble in water (51.8 g/L at 20 °C and 760 mmHg).

(a) Show how NH_3 can hydrogen bond to water.
(b) What is the solubility of ammonia in water in moles per liter?

9.108 Cobalt(II) chloride, a blue solid, can absorb water from the air to form cobalt(II) chloride hexahydrate, a pink solid. The equilibrium is so sensitive to moisture in the air that $CoCl_2$ is used as a humidity indicator.

(a) Write a balanced equation for the equilibrium. Be sure to include water as a reactant to produce the hexahydrate.
(b) How many grams of water are released by the decomposition of 2.50 g of cobalt(II) chloride hexahydrate?

9.109 How many milliliters of 0.150 M $BaCl_2$ are needed to react completely with 35.0 mL of 0.200 M Na_2SO_4? How many grams of $BaSO_4$ will be formed?

9.110 Many compounds are only partially dissociated into ions in aqueous solution. Trichloroacetic acid (CCl_3CO_2H), for instance, is partially dissociated in water according to the equation

$$CCl_3CO_2H(aq) \rightleftharpoons H^+(aq) + CCl_3CO_2^{\,-}(aq)$$

For a solution prepared by dissolving 1.00 mol of trichloroacetic acid in 1.00 kg of water, 36.0% of the trichloroacetic acid dissociates to form H^+ and $CCl_3CO_2^{\,-}$ ions.

(a) What is the total concentration of dissolved ions and molecules in 1 kg of water?
(b) What is the freezing point of this solution? (The freezing point of 1 kg of water is lowered 1.86 °C for each mole of solute particles.)