Chapter 8

Solutions and Colloids

Figure 8.1 Coral reefs, such as this one at the Palmyra Atoll National Wildlife Refuge, are vital to the ecosystem of earth's oceans but are threatened by climate change and dissolved pollution. Marine life depends on the specific chemical composition of the complex mixture we know as seawater. (credit: modification of work by "USFWS – Pacific Region"/Wikimedia Commons)

Chapter Outline

8.1 The Dissolution Process

8.2 Electrolytes

8.3 Solubility and net ionic equations

8.4 Molarity

8.5 Other Units for Solution Concentrations

8.6 Colligative Properties

8.7 Colloids

Introduction

Coral reefs are home to about 25% of all marine species. They are being threatened by climate change, oceanic acidification, and water pollution, all of which change the composition of the solution we know as seawater. Dissolved oxygen in seawater is critical for sea creatures, but as the oceans warm, oxygen becomes less soluble. As the concentration of carbon dioxide in the atmosphere increases, the concentration of carbon dioxide in the oceans increases, contributing to oceanic acidification. Coral reefs are particularly sensitive to the acidification of the ocean, since the exoskeletons of the coral polyps are soluble in acidic solutions. Humans contribute to the changing of seawater composition by allowing agricultural runoff and other forms of pollution to affect our oceans.

Solutions are crucial to the processes that sustain life and to many other processes involving chemical reactions. In this chapter, we will consider the nature of solutions, and examine factors that determine whether a solution will form and what properties it may have. In addition, we will discuss colloids—systems that resemble solutions but consist of dispersions of particles somewhat larger than ordinary molecules or ions.

8.1 The Dissolution Process

By the end of this section, you will be able to:

- Describe the basic properties of solutions and how they form
- Predict whether a given mixture will yield a solution based on molecular properties of its components
- Explain why some solutions either produce or absorb heat when they form

Solutions are defined as homogeneous mixtures of two or more substances. Often, one component of a solution is present at a significantly greater concentration, in which case it is called the *solvent*. The other components of the solution present in relatively lesser concentrations are called *solutes*. Sugar is a covalent solid composed of sucrose molecules, $C_{12}H_{22}O_{11}$. When this compound dissolves in water, its molecules become uniformly distributed among the molecules of water:

$$C_{12}H_{22}O_{11}(s) \longrightarrow C_{12}H_{22}O_{11}(aq)$$

The subscript "*aq*" in the equation signifies that the sucrose molecules are solutes and are therefore *individually dispersed* throughout the *aqueous solution* (water is the solvent). Although sucrose molecules are heavier than water molecules, they remain dispersed throughout the solution; gravity does not cause them to "settle out" over time.

Potassium dichromate, $K_2Cr_2O_7$, is an ionic compound composed of colorless potassium ions, K^+, and orange dichromate ions, $Cr_2O_7^{2-}$. When a small amount of solid potassium dichromate is added to water, the compound dissolves and dissociates to yield potassium ions and dichromate ions uniformly distributed throughout the mixture (**Figure 8.2**), as indicated in this equation:

$$K_2Cr_2O_7(s) \longrightarrow 2K^+(aq) + Cr_2O_7^{2-}(aq)$$

As with the mixture of sugar and water, this mixture is also an aqueous solution. Its solutes, potassium and dichromate ions, remain individually dispersed among the solvent (water) molecules.

Figure 8.2 When potassium dichromate ($K_2Cr_2O_7$) is mixed with water, it forms a homogeneous orange solution. (credit: modification of work by Mark Ott)

Link to Learning

Visit this **virtual lab (http://openstaxcollege.org/l/16Phetsugar)** to view simulations of the dissolution of common covalent and ionic substances (sugar and salt) in water.

Water is used so often as a solvent that the word solution has come to imply an aqueous solution to many people. However, almost any gas, liquid, or solid can act as a solvent. Many **alloys** are solid solutions of one metal dissolved in another; for example, US five-cent coins contain nickel dissolved in copper. Air is a gaseous solution, a homogeneous mixture of nitrogen, oxygen, and several other gases. Oxygen (a gas), alcohol (a liquid), and sugar (a solid) all dissolve in water (a liquid) to form liquid solutions. **Table 8.1** gives examples of several different solutions and the phases of the solutes and solvents.

Solution	Solute	Solvent
air	$O_2(g)$	$N_2(g)$
soft drinks[1]	$CO_2(g)$	$H_2O(l)$
hydrogen in palladium	$H_2(g)$	$Pd(s)$
rubbing alcohol	$H_2O(l)$	$C_3H_8O(l)$ (2-propanol)
saltwater	$NaCl(s)$	$H_2O(l)$
brass	$Zn(s)$	$Cu(s)$

Table 8.1 Different Types of Solutions

Solutions exhibit these defining traits:

- They are homogeneous; that is, after a solution is mixed, it has the same composition at all points throughout (its composition is uniform).
- The physical state of a solution—solid, liquid, or gas—is typically the same as that of the solvent, as demonstrated by the examples in **Table 8.1**.
- The components of a solution are dispersed on a molecular scale; that is, they consist of a mixture of separated molecules, atoms, and/or ions.
- The dissolved solute in a solution will not settle out or separate from the solvent.
- The composition of a solution, or the concentrations of its components, can be varied continuously, within limits.

The Formation of Solutions

The formation of a solution is an example of a **spontaneous process**, a process that occurs under specified conditions without the requirement of energy from some external source. Sometimes we stir a mixture to speed up the dissolution

process, but this is not necessary; a homogeneous solution would form if we waited long enough. When the strengths of the intermolecular forces of attraction between solute and solvent species in a solution are no different than those present in the separated components, the solution is formed with no accompanying energy change. Such a solution is called an **ideal solution**. A mixture of ideal gases (or gases such as helium and argon, which closely approach ideal behavior) is an example of an ideal solution, since the entities comprising these gases experience no significant intermolecular attractions.

When containers of helium and argon are connected, the gases spontaneously mix due to diffusion and form a solution (**Figure 8.3**). The formation of this solution clearly involves an increase in disorder, since the helium and argon atoms occupy a volume twice as large as that which each occupied before mixing.

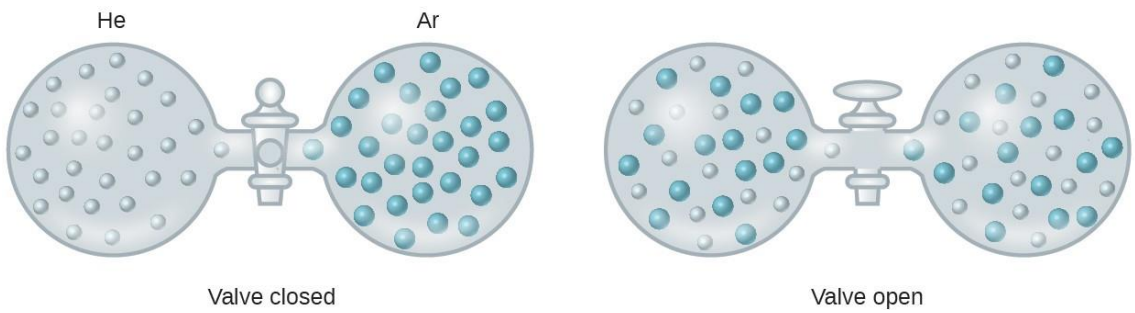


Figure 8.3 Samples of helium and argon spontaneously mix to give a solution in which the disorder of the atoms of the two gases is increased.

Ideal solutions may also form when structurally similar liquids are mixed. For example, mixtures of the alcohols methanol (CH_3OH) and ethanol (C_2H_5OH) form ideal solutions, as do mixtures of the hydrocarbons pentane, C_5H_{12}, and hexane, C_6H_{14}. Placing methanol and ethanol, or pentane and hexane, in the bulbs shown in **Figure 8.3** will result in the same diffusion and subsequent mixing of these liquids as is observed for the He and Ar gases (although at a much slower rate), yielding solutions with no significant change in energy. Unlike a mixture of gases, however, the components of these liquid-liquid solutions do, indeed, experience intermolecular attractive forces. But since the molecules of the two substances being mixed are structurally very similar, the intermolecular attractive forces between like and unlike molecules are essentially the same, and the dissolution process, therefore, does not entail any appreciable increase or decrease in energy. These examples illustrate how diffusion alone can provide the driving force required to cause the spontaneous formation of a solution. In some cases, however, the relative magnitudes of intermolecular forces of attraction between solute and solvent species may prevent dissolution.

Three types of intermolecular attractive forces are relevant to the dissolution process: solute-solute, solvent-solvent, and solute-solvent. As illustrated in **Figure 8.4**, the formation of a solution may be viewed as a stepwise process in which energy is consumed to overcome solute-solute and solvent-solvent attractions (endothermic processes) and released when solute-solvent attractions are established (an exothermic process referred to as **solvation**). The relative magnitudes of the energy changes associated with these stepwise processes determine whether the dissolution process overall will release or absorb energy. In some cases, solutions do not form because the energy required to separate solute and solvent species is so much greater than the energy released by solvation.

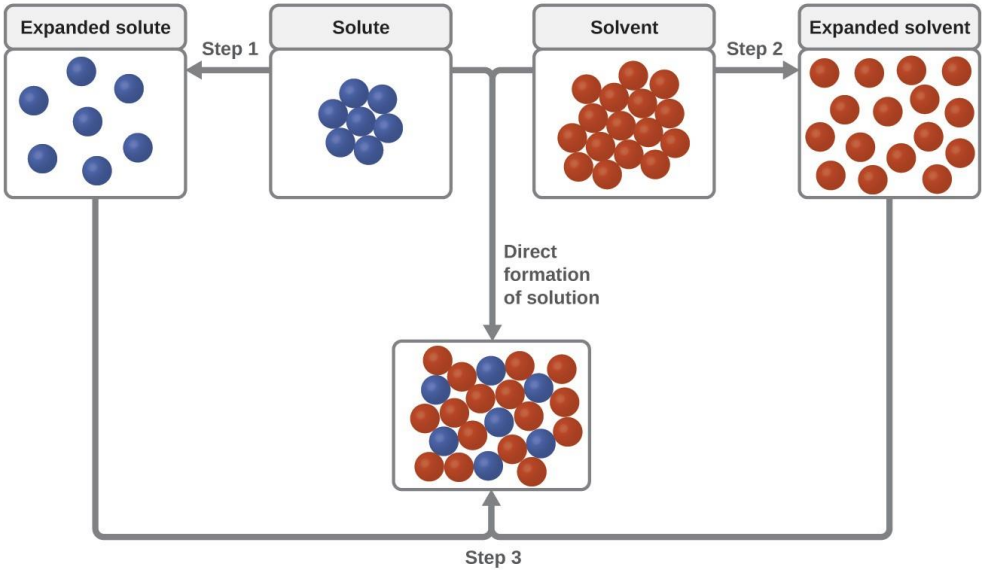


Figure 8.4 This schematic representation of dissolution shows a stepwise process involving the endothermic separation of solute and solvent species (Steps 1 and 2) and exothermic solvation (Step 3).

For example, cooking oils and water will not mix to any appreciable extent to yield solutions (**Figure 8.5**). Hydrogen bonding is the dominant intermolecular attractive force present in liquid water; the nonpolar hydrocarbon molecules of cooking oils are not capable of hydrogen bonding, instead being held together by dispersion forces. Forming an oil-water solution would require overcoming the very strong hydrogen bonding in water, as well as the significantly strong dispersion forces between the relatively large oil molecules. And, since the polar water molecules and nonpolar oil molecules would not experience very strong intermolecular attraction, very little energy would be released by solvation.

Figure 8.5 A mixture of nonpolar cooking oil and polar water does not yield a solution. (credit: Gautam Dogra)

On the other hand, a mixture of ethanol and water will mix in any proportions to yield a solution. In this case, both substances are capable of hydrogen bonding, and so the solvation process is sufficiently exothermic to compensate for the endothermic separations of solute and solvent molecules.

As noted at the beginning of this module, spontaneous solution formation is favored, but not guaranteed, by exothermic dissolution processes. While many soluble compounds do, indeed, dissolve with the release of heat, some dissolve endothermically. Ammonium nitrate (NH_4NO_3) is one such example and is used to make instant cold packs for treating injuries like the one pictured in **Figure 8.6**. A thin-walled plastic bag of water is sealed inside a larger bag with solid NH_4NO_3. When the smaller bag is broken, a solution of NH_4NO_3 forms, absorbing heat from the surroundings (the injured area to which the pack is applied) and providing a cold compress that decreases swelling. Endothermic dissolutions such as this one require a greater energy input to separate the solute species than is recovered when the solutes are solvated, but they are spontaneous nonetheless due to the increase in disorder that accompanies formation of the solution.

Figure 8.6 An instant cold pack gets cold when certain salts, such as ammonium nitrate, dissolve in water—an endothermic process.

Link to Learning

Watch this brief **video (http://openstaxcollege.org/l/16endoexo)** illustrating endothermic and exothermic dissolution processes.

8.2 Electrolytes

By the end of this module, you will be able to:

- Define and give examples of electrolytes
- Distinguish between the physical and chemical changes that accompany dissolution of ionic and covalent electrolytes
- Relate electrolyte strength to solute-solvent attractive forces

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not

yield ions when dissolved are called **nonelectrolytes**. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a **strong electrolyte**. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a **weak electrolyte**. Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (**Figure 8.7**).

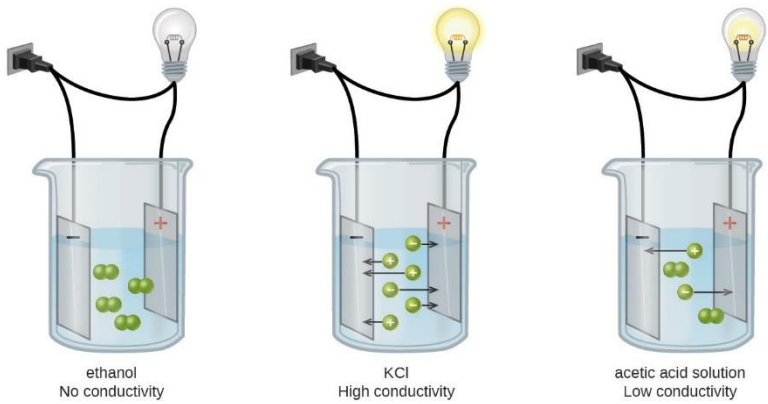


Figure 8.7 Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in **Figure 8.8**. The electrostatic attraction between an ion and a molecule with a dipole is called an **ion-dipole attraction**. These attractions play an important role in the dissolution of ionic compounds in water.

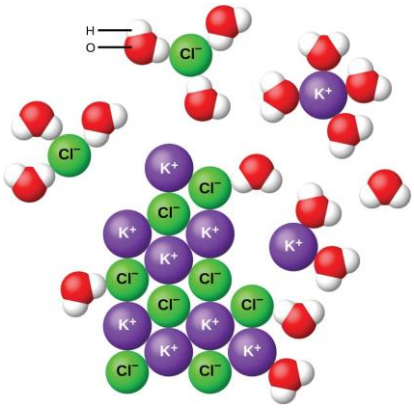


Figure 8.8 As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are

attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as **dissociation**. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. The water molecules penetrate between individual K^+ and Cl^- ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as **Figure 8.8** shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

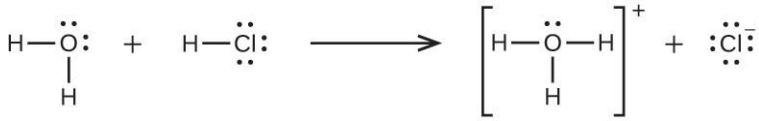
Covalent Electrolytes

Pure water is an extremely poor conductor of electricity because it is only very slightly ionized—only about two out of every 1 billion molecules ionize at 25 °C. Water ionizes when one molecule of water gives up a proton (H^+ ion) to another molecule of water, yielding hydronium and hydroxide ions.

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

In some cases, we find that solutions prepared from covalent compounds conduct electricity because the solute molecules react chemically with the solvent to produce ions. For example, pure hydrogen chloride is a gas consisting of covalent HCl molecules. This gas contains no ions. However, when we dissolve hydrogen chloride in water, we find that the solution is a very good conductor. The water molecules play an essential part in forming ions: Solutions of hydrogen chloride in many other solvents, such as benzene, do not conduct electricity and do not contain ions.

Hydrogen chloride is an *acid*, and so its molecules react with water, transferring H^+ ions to form hydronium ions (H_3O^+) and chloride ions (Cl^-):



This reaction is essentially 100% complete for HCl (i.e., it is a *strong acid* and, consequently, a strong electrolyte). Likewise, weak acids and bases that only react partially generate relatively low concentrations of ions when dissolved in water and are classified as weak electrolytes. The reader may wish to review the discussion of strong and weak

acids provided in the earlier chapter of this text on reaction classes and stoichiometry.

8.3 Solubility

By the end of this module, you will be able to:

- Describe the effects of temperature and pressure on solubility
- State Henry's law and use it in calculations involving the solubility of a gas in a liquid
- Explain the degrees of solubility possible for liquid-liquid solutions

Imagine adding a small amount of salt to a glass of water, stirring until all the salt has dissolved, and then adding a bit more. You can repeat this process until the salt concentration of the solution reaches its natural limit, a limit determined primarily by the relative strengths of the solute-solute, solute-solvent, and solvent-solvent attractive forces discussed in the previous two modules of this chapter. You can be certain that you have reached this limit because, no matter how long you stir the solution, undissolved salt remains. The concentration of salt in the solution at this point is known as its solubility.

The **solubility** of a solute in a particular solvent is the maximum concentration that may be achieved under given conditions when the dissolution process is *at equilibrium*. Referring to the example of salt in water:

$$\mathrm{NaCl}(s) \rightleftharpoons \mathrm{Na}^{+}(aq) + \mathrm{Cl}^{-}(aq)$$

When a solute's concentration is equal to its solubility, the solution is said to be **saturated** with that solute. If the solute's concentration is less than its solubility, the solution is said to be **unsaturated**. A solution that contains a relatively low concentration of solute is called dilute, and one with a relatively high concentration is called concentrated.

If we add more salt to a saturated solution of salt, we see it fall to the bottom and no more seems to dissolve. In fact, the added salt does dissolve, as represented by the forward direction of the dissolution equation. Accompanying this process, dissolved salt will precipitate, as depicted by the reverse direction of the equation. The system is said to be at equilibrium when these two reciprocal processes are occurring at equal rates, and so the amount of undissolved and dissolved salt remains constant. Support for the simultaneous occurrence of the dissolution and precipitation processes is provided by noting that the number and sizes of the undissolved salt crystals will change over time, though their combined mass will remain the same.

Link to Learning

Use this **interactive simulation (http://openstaxcollege.org/l/16Phetsoluble)** to prepare various saturated solutions.

Solutions may be prepared in which a solute concentration *exceeds* its solubility. Such solutions are said to be **supersaturated**, and they are interesting examples of *nonequilibrium* states. For example, the carbonated beverage in an open container that has not yet "gone flat" is supersaturated with carbon dioxide gas; given time, the CO_2

concentration will decrease until it reaches its equilibrium value.

Link to Learning

Watch this **impressive video (http://openstaxcollege.org/l/16NaAcetate)** showing the precipitation of sodium acetate from a supersaturated solution.

Solutions of Gases in Liquids

In an earlier module of this chapter, the effect of intermolecular attractive forces on solution formation was discussed. The chemical structures of the solute and solvent dictate the types of forces possible and, consequently, are important factors in determining solubility. For example, under similar conditions, the water solubility of oxygen is approximately three times greater than that of helium, but 100 times less than the solubility of chloromethane, $CHCl_3$. Considering the role of the solvent's chemical structure, note that the solubility of oxygen in the liquid hydrocarbon hexane, C_6H_{14}, is approximately 20 times greater than it is in water.

Other factors also affect the solubility of a given substance in a given solvent. Temperature is one such factor, with gas solubility typically decreasing as temperature increases (**Figure 8.9**). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.

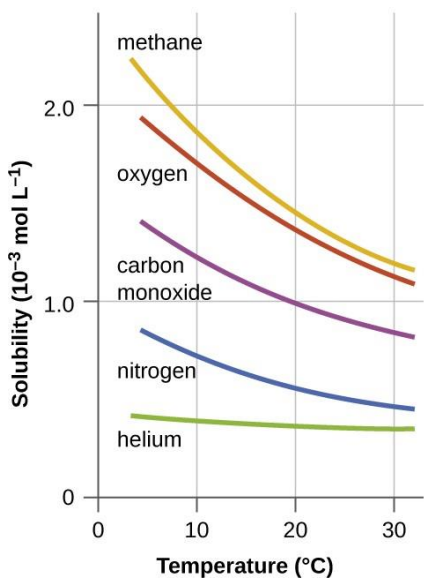


Figure 8.9 The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills (**Figure 8.10**).

(a) (b)

Figure 8.10 (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (credit a: modification of work by Liz West; credit b: modification of work by U.S. Fish and Wildlife Service)

The solubility of a gaseous solute is also affected by the partial pressure of solute in the gas to which the solution is exposed. Gas solubility increases as the pressure of the gas increases. Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at this pressure. When the beverage container is opened, a familiar hiss is heard as the carbon dioxide gas pressure is released, and some of the dissolved carbon dioxide is typically seen leaving solution in the form of small bubbles (**Figure 8.11**). At this point, the beverage is *supersaturated* with carbon dioxide and, with time, the dissolved carbon dioxide concentration will decrease to its equilibrium value and the beverage will become "flat."

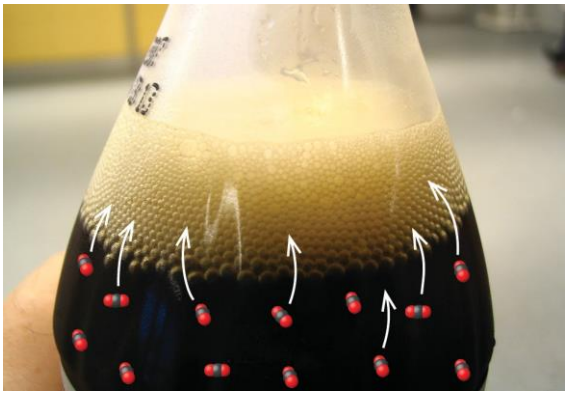


Figure 8.11 Opening the bottle of carbonated beverage reduces the pressure of the *gaseous* carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some *dissolved* carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Derrick Coetzee)

For many gaseous solutes, the relation between solubility, C_g, and partial pressure, P_g, is a proportional one:

$$C_g = kP_g$$

where k is a proportionality constant that depends on the identities of the gaseous solute and solvent, and on the solution temperature. This is a mathematical statement of **Henry's law**: *The quantity of an ideal gas that dissolves in a definite volume of liquid is directly proportional to the pressure of the gas.*

Example 8.1

Application of Henry's Law

At 20 °C, the concentration of dissolved oxygen in water exposed to gaseous oxygen at a partial pressure of 101.3 kPa (760 torr) is 1.38 × 10−3 mol L−1. Use Henry's law to determine the solubility of oxygen when its partial pressure is 20.7 kPa (155 torr), the approximate pressure of oxygen in earth's atmosphere.

Solution:

According to Henry's law, for an ideal solution the solubility, Cg, of a gas (1.38 × 10−3 mol L−1, in this case) is directly proportional to the pressure, Pg, of the undissolved gas above the solution (101.3 kPa, or 760 torr, in this case). Because we know both Cg and Pg, we can rearrange this expression to solve for k.

C_g = kPg

$$K = \frac{C_g}{P_g}$$

$$= \frac{1.38 * 10^{-3} molL^{-1}}{101.3kpa}$$

$$= 1.36 * 10^{-5}\ mol\ L^{-1}\ kPa^{-1}$$

Or

$$= 1.82 * 10^{-6}\ mol\ L^{-1}$$

Now we can use K to find the solubility at the lower pressure.

$$1.36 * 10^{-5} molL^{-1} * 20.7\ kpa$$

Or

$$1.82 * 10^{-6}\ molL^{-1} torr^{-1} * 155\ torr$$

$$= 2.82 * 10^{-4}\ molL^{-1}$$

Note that various units may be used to express the quantities involved in these sorts of computations. Any combination of units that yield to the constraints of dimensional analysis are acceptable.

Check Your Learning

Exposing a 100.0 mL sample of water at 0 °C to an atmosphere containing a gaseous solute at 20.26 kPa (152 torr) resulted in the dissolution of 1.45 × 10−3 g of the solute. Use Henry's law to determine the solubility of this gaseous solute when its pressure is 101.3 kPa (760 torr).

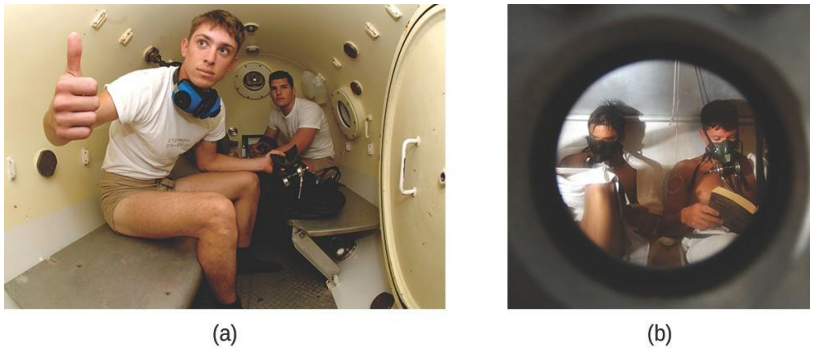
Answer: 7.25 × 10−3 in 100.0 mL or 0.0725 g/L

Chemistry in Everyday Life: Decompression Sickness or "The Bends"

Decompression sickness (DCS), or "the bends," is an effect of the increased pressure of the air inhaled by scuba divers when swimming underwater at considerable depths. In addition to the pressure exerted by the atmosphere, divers are subjected to additional pressure due to the water above them, experiencing an increase of approximately 1 atm for each 10 m of depth. Therefore, the air inhaled by a diver while submerged contains gases at the corresponding higher ambient pressure, and the concentrations of the gases dissolved in the diver's blood are proportionally higher per Henry's law.

As the diver ascends to the surface of the water, the ambient pressure decreases, and the dissolved gases becomes less soluble. If the ascent is too rapid, the gases escaping from the diver's blood may form bubbles that can cause a variety

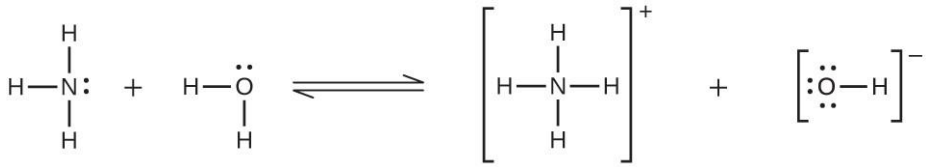
of symptoms ranging from rashes and joint pain to paralysis and death. To avoid DCS, divers must ascend from depths at relatively slow speeds (10 or 20 m/min) or otherwise make several decompressions stops, pausing for several minutes at given depths during the ascent. When these preventive measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers (**Figure 8.12**).



(a) (b)

Figure 8.12 (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.

Deviations from Henry's law are observed when a chemical reaction takes place between the gaseous solute and the solvent. Thus, for example, the solubility of ammonia in water does not increase as rapidly with increasing pressure as predicted by the law because ammonia, being a base, reacts to some extent with water to form ammonium ions and hydroxide ions.



Gases can form supersaturated solutions. If a solution of a gas in a liquid is prepared either at low temperature or under pressure (or both), then as the solution warms or as the gas pressure is reduced, the solution may become supersaturated. In 1986, more than 1700 people in Cameroon were killed when a cloud of gas, almost certainly carbon dioxide, bubbled from Lake Nyos (**Figure 8.13**), a deep lake in a volcanic crater. The water at the bottom of Lake Nyos is saturated with carbon dioxide by volcanic activity beneath the lake. It is believed that the lake underwent a turnover due to gradual heating from below the lake, and the warmer, less-dense water saturated with carbon dioxide reached the surface. Consequently, tremendous quantities of dissolved CO_2 were released, and the colorless gas, which is denser than air, flowed down the valley below the lake and suffocated humans and animals living in the valley.

(a)

(b)

Figure 8.13 (a) It is believed that the 1986 disaster that killed more than 1700 people near Lake Nyos in Cameroon resulted when a large volume of carbon dioxide gas was released from the lake. (b) A CO_2 vent has since been installed to help outgas the lake in a slow, controlled fashion and prevent a similar catastrophe from happening in the future. (credit a: modification of work by Jack Lockwood; credit b: modification of work by Bill Evans)

Solutions of Liquids in Liquids

We know that some liquids mix with each other in all proportions; in other words, they have infinite mutual solubility and are said to be **miscible**. Ethanol, sulfuric acid, and ethylene glycol (popular for use as antifreeze, pictured in **Figure 8.14**) are examples of liquids that are completely miscible with water. Two-cycle motor oil is miscible with gasoline.

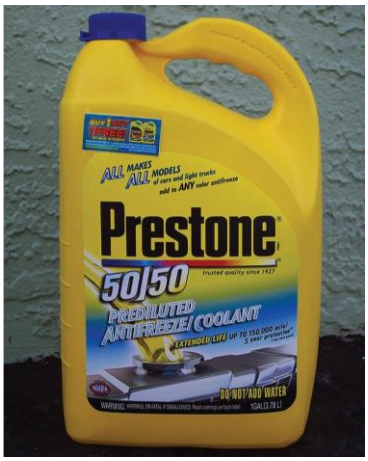


Figure 8.14 Water and antifreeze are miscible; mixtures of the two are homogeneous in all proportions. (credit: "dno1967"/Wikimedia commons)

Liquids that mix with water in all proportions are usually polar substances or substances that form hydrogen bonds. For such liquids, the dipole-dipole attractions (or hydrogen bonding) of the solute molecules with the solvent molecules are at least as strong as those between molecules in the pure solute or in the pure solvent. Hence, the two kinds of molecules mix easily. Likewise, nonpolar liquids are miscible with each other because there is no appreciable difference in the strengths of solute-solute, solvent-solvent, and solute-solvent intermolecular attractions. The solubility of polar molecules in polar solvents and of

nonpolar molecules in nonpolar solvents is, again, an illustration of the chemical axiom “like dissolves like.”

Two liquids that do not mix to an appreciable extent are called **immiscible**. Layers are formed when we pour immiscible liquids into the same container. Gasoline, oil (**Figure 8.15**), benzene, carbon tetrachloride, some paints, and many other nonpolar liquids are immiscible with water. The attraction between the molecules of such nonpolar liquids and polar water molecules is ineffectively weak. The only strong attractions in such a mixture are between the water molecules, so they effectively squeeze out the molecules of the nonpolar liquid. The distinction between immiscibility and miscibility is really one of degrees, so that miscible liquids are of infinite mutual solubility, while liquids said to be immiscible are of very low (though not zero) mutual solubility.

Figure 8.15 Water and oil are immiscible. Mixtures of these two substances will form two separate layers with the less dense oil floating on top of the water. (credit: “Yortw”/Flickr)

Two liquids, such as bromine and water, that are of *moderate* mutual solubility are said to be **partially miscible**. Two partially miscible liquids usually form two layers when mixed. In the case of the bromine and water mixture, the upper layer is water, saturated with bromine, and the lower layer is bromine saturated with water. Since bromine is nonpolar, and, thus, not very soluble in water, the water layer is only slightly discolored by the bright orange bromine dissolved in it. Since the solubility of water in bromine is very low, there is no noticeable effect on the dark color of the bromine layer (**Figure 8.16**).

Figure 8.16 Bromine (the deep orange liquid on the left) and water (the clear liquid in the middle) are partially miscible. The top layer in the mixture on the right is a saturated solution of bromine in water; the bottom layer is a saturated solution of water in bromine. (credit: Paul Flowers)

Solutions of Solids in Liquids

The dependence of solubility on temperature for a number of solids in water is shown by the solubility curves in **Figure 8.17**. Reviewing these data indicate a general trend of increasing solubility with temperature, although there are exceptions, as illustrated by the ionic compound cerium sulfate.

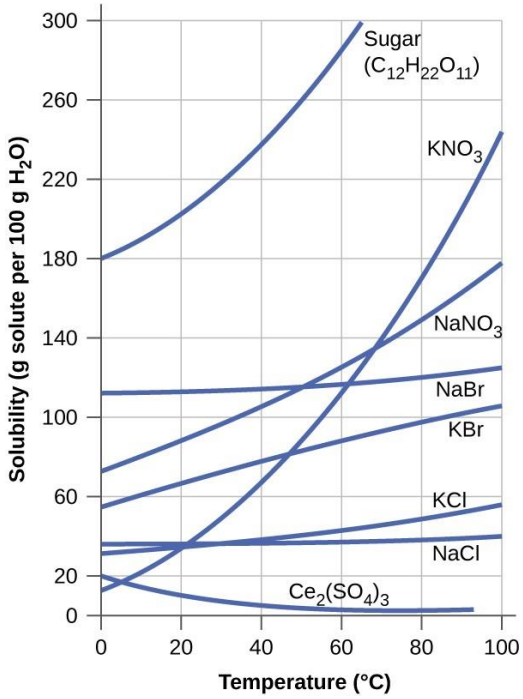


Figure 8.17 This graph shows how the solubility of several solids changes with temperature.

The temperature dependence of solubility can be exploited to prepare supersaturated solutions of certain compounds. A solution may be saturated with the compound at an elevated temperature (where the solute is more soluble) and subsequently cooled to a lower temperature without precipitating the solute. The resultant solution contains solute at a concentration greater than its equilibrium solubility at the lower temperature (i.e., it is supersaturated) and is relatively stable. Precipitation of the excess solute can be initiated by adding a seed crystal (see the video in the Link to Learning earlier in this module) or by mechanically agitating the solution. Some hand warmers, such as the one pictured in **Figure 8.18**, take advantage of this behavior.

Figure 8.18 This hand warmer produces heat when the sodium acetate in a supersaturated solution precipitate. Precipitation of the solute is initiated by a mechanical shockwave generated when the flexible metal disk within the solution is "clicked." (credit: modification of work by "Velela"/Wikimedia Commons)

Link to Learning

This video (http://openstaxcollege.org/l/16handwarmer) shows the crystallization process occurring in a hand warmer.

Precipitation Reactions and Solubility Rules

A **precipitation reaction** is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as *double displacement*, *double replacement*, or *metathesis* reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals. Precipitation reactions also play a central role in many chemical analysis techniques, including spot tests used to identify metal ions and *gravimetric methods* for determining the composition of matter (see the last module of this chapter).

The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its **solubility**, defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be **soluble**. A substance will **precipitate** when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be **insoluble**, and these are the substances that readily precipitate from solution. More information on these important concepts is provided in the next chapter on solutions. To predict the identities of solids formed by precipitation reactions, one may simply refer to patterns of solubility that have been observed for many ionic compound (Table8.2).

Solubilities of Common Ionic Compounds in Water

Soluble compounds contain • group 1 metal cations (Li^+, Na^+, K^+, Rb^+, and Cs^+) and ammonium ion (NH_4^+) • the halide ions (Cl^-, Br^-, and I^-) • the acetate ($C_2H_3O_2^-$), bicarbonate (HCO_3^-), nitrate (NO_3^-), and chlorate (ClO_3^-) ions • the sulfate (SO_4^-) ion	**Exceptions** to these solubility rules include • halides of Ag^+, Hg_2^{2+}, and Pb^{2+} • sulfates of Ag^+, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+}, and Sr^{2+}
Insoluble compounds contain • carbonate (CO_3^{2-}), chromate (CrO_3^{2-}), phosphate (PO_4^{3-}), and sulfide (S^{2-}) ions • hydroxide ion (OH^-)	**Exceptions** to these insolubility rules include • compounds of these anions with group 1 metal cations and ammonium ion • hydroxides of group 1 metal cations and Ba^{2+}

Table 8.2 Solubilities of Common Ionic Compounds in Water

A vivid example of precipitation is observed when solutions of potassium iodide and lead nitrate are mixed, resulting in the formation of solid lead iodide:

$$2KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$$

This observation is consistent with the solubility guidelines: The only insoluble compound among all those involved is lead iodide, one of the exceptions to the general solubility of iodide salts.

The net ionic equation representing this reaction is:

$$Pb^{2+}(aq) + 2I^{-}(aq) \longrightarrow PbI_2(s)$$

Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow (**Figure 8.19**). The properties of pure PbI_2 crystals make them useful for fabrication of X-ray and gamma ray detectors.

Figure 8.19 A precipitate of PbI_2 forms when solutions containing Pb^{2+} and I^{-} are mixed. (credit: Der Kreole/ Wikimedia Commons)

The solubility guidelines in **Table 8.2** may be used to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. One merely needs to identify all the ions present in the solution and then consider if possible cation/anion pairing could result in an insoluble compound. For example, mixing solutions of silver nitrate and sodium fluoride will yield a solution containing Ag^{+}, NO_3^{-}, Na^{+}, and F^{-}ions. Aside from the two ionic compounds originally present in the solutions, $AgNO_3$ and NaF, two additional ionic compounds may be derived from this collection of ions: $NaNO_3$ and AgF. The solubility guidelines indicate all nitrate salts are soluble but that AgF is one of the exceptions to the general solubility of fluoride salts. A precipitation reaction, therefore, is predicted to occur, as described by the following equations:

$$NaF(aq) + AgNO_3(aq) \longrightarrow AgF(s) + NaNO_3(aq) \text{ (molecular)}$$

Equations for Ionic Reactions

Given the abundance of water on earth, it stands to reason that a great many chemical reactions take place in aqueous media. When ions are involved in these reactions, the chemical equations may be written with various levels of detail appropriate to their intended use. To illustrate this, consider a reaction between ionic compounds taking place in an

aqueous solution. When aqueous solutions of $CaCl_2$ and $AgNO_3$ are mixed, a reaction takes place producing aqueous $Ca(NO_3)_2$ and solid AgCl:

$$CaCl_2(aq) + 2AgNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$$

This balanced equation, derived in the usual fashion, is called a **molecular equation** because it doesn't explicitly represent the ionic species that are present in solution. When ionic compounds dissolve in water, they may *dissociate* into their constituent ions, which are subsequently dispersed homogenously throughout the resulting solution (a thorough discussion of this important process is provided in the chapter on solutions). Ionic compounds dissolved in water are, therefore, more realistically represented as dissociated ions, in this case:

$$CaCl_2(aq) \longrightarrow Ca^{2+}(aq) + 2Cl^{-}(aq) \quad 2AgNO_3(aq) \longrightarrow 2Ag^{+}(aq) + 2NO_3^{-}(aq) \quad Ca(NO_3)_2(aq) \longrightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq)$$

Unlike these three ionic compounds, AgCl does not dissolve in water to a significant extent, as signified by its physical state notation, *s*. Explicitly representing all dissolved ions results in a **complete ionic equation**. In this particular case, the formulas for the dissolved ionic compounds are replaced by formulas for their dissociated ions:

$$Ca^{2+}(aq) + 2Cl^{-}(aq) + 2Ag^{+}(aq) + 2NO_3^{-}(aq) \longrightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq) + 2AgCl(s)$$

Examining this equation shows that two chemical species are present in identical form on both sides of the arrow, $Ca^{2+}(aq)$ and $NO_3^{-}(aq)$. These **spectator ions**—ions whose presence is required to maintain charge neutrality—are neither chemically nor physically changed by the process, and so they may be eliminated from the equation to yield more succinct representation called a **net ionic equation**: $Ca^{2+}(aq) + 2Cl^{-}(aq) + 2Ag^{+}(aq) + 2NO_3^{-}(aq) \longrightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq) + 2AgCl(s)$ $2Cl^{-}(aq) + 2Ag^{+}(aq) \longrightarrow 2AgCl(s)$

Following the convention of using the smallest possible integers as coefficients, this equation is then written:

$$Cl^{-}(aq) + Ag^{+}(aq) \longrightarrow AgCl(s)$$

This net ionic equation indicates that solid silver chloride may be produced from dissolved chloride and silver(I) ions, regardless of the source of these ions. These molecular and complete ionic equations provide additional information, namely, the ionic compounds used as sources of Cl^{-} and Ag^{+}.

Example 8.2

Molecular and Ionic Equations: When carbon dioxide is dissolved in an aqueous solution of sodium hydroxide, the mixture reacts to yield aqueous sodium carbonate and liquid water. Write balanced molecular, complete ionic, and net ionic equations for this process.

Solution:

Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

$$CO2\ (aq) + NaOH(aq) \longrightarrow Na2\ CO3(aq) + H2\ O(l) \quad \text{(unbalanced)}$$

Balance is achieved easily in this case by changing the coefficient for NaOH to 2, resulting in the molecular equation for this reaction:

$$CO2\ (aq) + 2NaOH(aq) \longrightarrow Na2\ CO3(aq) + H2\ O(l)$$

The two dissolved ionic compounds, NaOH and Na2CO3, can be represented as dissociated ions to yield the complete

ionic equation:

$$CO_2(aq) + 2Na+(aq) + 2OH-(aq) \longrightarrow 2Na+(aq) + CO_3^{2-}9aq) + H2\ O(l)$$

Finally, identify the spectator ion(s), in this case Na+(aq), and remove it from each side of the equation to generate the net ionic equation:

$$CO_2(aq) + 2Na+(aq) + 2OH-(aq) \longrightarrow 2Na+(aq) + CO_3^{2-}(aq) + H2\ O(l)\ CO2(aq) + 2OH-(aq) \longrightarrow CO3\ 2-(aq) + H2\ O(l)$$

Check Your Learning

Diatomic chlorine and sodium hydroxide (lye) are commodity chemicals produced in large quantities, along with diatomic hydrogen, via the electrolysis of brine, according to the following unbalanced equation:

$$NaCl(aq) + H2\ O(l) \xrightarrow{electricity} NaOH(aq) + H2(g) + Cl2(g)$$

Write balanced molecular, complete ionic, and net ionic equations for this process.

Answer: 2NaCl(aq) + 2H2 O(l) ⟶ 2NaOH(aq) + H2(g) + Cl2(g) (molecular)

2Na+(aq) + 2Cl−(aq) + 2H2 O(l) ⟶ 2Na+(aq) + 2OH−(aq) + H2(g) + Cl2(g) (complete ionic)

2Cl−(aq) + 2H2 O(l) ⟶ 2OH−(aq) + H2(g) + Cl2(g) (net ionic)

Example 8.3

Predicting Precipitation Reactions Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If precipitation is expected, write a balanced net ionic equation for the reaction.

(a) potassium sulfate and barium nitrate

(b) lithium chloride and silver acetate

(c) lead nitrate and ammonium carbonate

Solution:

(a) The two possible products for this combination are KNO3 and BaSO4. The solubility guidelines indicate BaSO4 is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction,

Derived in the manner detailed in the previous module, is

$$Ba^{2+}(aq) + So_4^{2-}(aq) \rightarrow$$

(b) The two possible products for this combination are LiC2H3O2 and AgCl. The solubility guidelines indicate AgCl is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$Ag+(aq) + Cl-(aq) \longrightarrow AgCl(s)$$

(c) The two possible products for this combination are PbCO3 and NH4NO3. The solubility guidelines indicate PbCO3 is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$Pb2+(aq) + CO_3^{2-}\ aq \longrightarrow$$

Check Your Learning

Which solution could be used to precipitate the barium ion, Ba2+, in a water sample: sodium chloride, sodium hydroxide, or sodium sulfate? What is the formula for the expected precipitate?

Answer: sodium sulfate, $BaSO_4$

Assessment 8.1:

Saline, a solution of salt (sodium chloride) in water, is used to clean wounds and remove contact lenses. It is prepared by dissolving sodium chloride in water, 0.90% mass/volume. When saline is made on an industrial scale, small portions of the solution are titrated to precipitate out the chloride ions as an insoluble salt, and the mass of the precipitate thus obtained is used to determine the concentration of sodium chloride. We will examine this quantitatively in Assessment 8.11.

(a). What metal ions can be used to precipitate out chloride ions?

(b). Write the net ionic equations for (a).

Answer:

Strategy: Refer to the solubility rules in the table above.

(a). The chlorides of metals are soluble in water, EXCEPT those of silver, mercury or lead. In other words, AgCl, Hg_2Cl_2 and $PbCl_2$ are insoluble in water and would precipitate out. In practice, nitrates of these metal ions, which are water soluble, can be added to a solution containing chloride ions to precipitate out the chloride ions as insoluble salts.

(b). The net ionic equations are given below.

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

$Hg_2^{2+}(aq) + 2Cl^-(aq) \rightarrow Hg_2Cl_2(s)$

$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

Assessment 8.2:

Because of percolation in rocks, water contains calcium and magnesium ions. "Hard water" is defined, based on high content of these metal ions, and has been linked with eczema in children, specially infants.

(a). What anions can be used to precipitate out calcium and magnesium ions?

(b). Write the net ionic equations for (a).

Answer:

Strategy: Refer to the solubility rules in the table above.

(a). Calcium and magnesium ions can be precipitated out as carbonates, sulfides, chromates, phosphates, or hydroxides.

(b). The net ionic equations are given below for calcium ions; the same applies for magnesium ions. We note that the charge on these ions is +2.

$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

$Ca^{2+}(aq) + S^{2-}(aq) \rightarrow CaS(s)$

$Ca^{2+}(aq) + CrO_4^{2-}(aq) \rightarrow CaCrO_4(s)$

$3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

$Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$

8.4 Molarity

By the end of this section, you will be able to:

- Describe the fundamental properties of solutions
- Calculate solution concentrations using molarity

In preceding sections, we focused on the composition of substances: samples of matter that contain only one type of element or compound. However, mixtures—samples of matter containing two or more substances physically combined—are more commonly encountered in nature than are pure substances. Similar to a pure substance, the relative composition of a mixture plays an important role in determining its properties. The relative amount of oxygen in a planet's atmosphere determines its ability to sustain aerobic life. The relative amounts of iron, carbon, nickel, and other elements in steel (a mixture known as an "alloy") determine its physical strength and resistance to corrosion. The relative amount of the active ingredient in a medicine determines its effectiveness in achieving the desired pharmacological effect. The relative amount of sugar in a beverage determines its sweetness (see **Figure 8.20**). In this section, we describe one of the most common ways in which the relative compositions of mixtures may be quantified.

Figure 8.20 Sugar is one of many components in the complex mixture known as coffee. The amount of sugar in a given amount of coffee is an important determinant of the beverage's sweetness. (credit: Jane Whitney)

Solutions

We have previously defined solutions as homogeneous mixtures, meaning that the composition of the mixture (and therefore its properties) is uniform throughout its entire volume. Solutions occur frequently in nature and have also been implemented in many forms of manmade technology. We will explore a more thorough treatment of solution properties in the chapter on solutions and colloids, but here we will introduce some of the basic properties of solutions.

The relative amount of a given solution component is known as its **concentration**. Often, though not always, a solution contains one component with a concentration that is significantly greater than that of all other components. This component is called the **solvent** and may be viewed as the medium in which the other components are dispersed, or **dissolved**. Solutions in which water is the solvent are, of course, very common on our planet. A solution in which water is the solvent is called an **aqueous solution**.

A **solute** is a component of a solution that is typically present at a much lower concentration than the solvent. Solute concentrations are often described with qualitative terms such as **dilute** (of relatively low concentration) and **concentrated** (of relatively high concentration). Concentrations may be quantitatively assessed using a wide variety of measurement units, each convenient for particular applications. **Molarity (*M*)** is a useful concentration unit for

many applications in chemistry. Molarity is defined as the number of moles of solute in exactly 1 liter (1 L) of the solution:

$$M = \frac{mol\ solute}{L\ solution}$$

Example 8.4

Calculating Molar Concentrations: A 355-mL soft drink sample contains 0.133 mol of sucrose (table sugar). What is the molar concentration of sucrose in the beverage?

Solution:

Since the molar amount of solute and the volume of solution are both given, the molarity can be calculated using the definition of molarity. Per this definition, the solution volume must be converted from mL to L

$$M = \frac{mol\ solute}{L\ solution} = \frac{0.133\ mol}{355\ mL * \frac{1\ L}{1000\ mL}} = 0.375\ M$$

Check Your Learning

A teaspoon of table sugar contains about 0.01 mol sucrose. What is the molarity of sucrose if a teaspoon of sugar has been dissolved in a cup of tea with a volume of 200 mL?

Answer: 0.05 M

Example 8.5

Deriving Moles and Volumes from Molar Concentrations: How much sugar (mol) is contained in a modest sip (~10 mL) of the soft drink from Example 8.4?

Solution:

In this case, we can rearrange the definition of molarity to isolate the quantity sought, moles of sugar. We then substitute the value for molarity that we derived in Example 8.4, 0.375 M:

$$M = \frac{mol\ solute}{L\ solution}$$

$$mol\ solute = M * L\ solution$$

$$mol\ solute = 0.375\ \frac{mol\ sugar}{L} * 10\ mL * \frac{1\ L}{1000\ mL} = 0.004\ mol\ sugar$$

Check Your Learning

What volume (mL) of the sweetened tea described in Example 8.4 contains the same amount of sugar (mol) as 10 mL of the soft drink in this example?

Answer: 80 mL

Example 8.6

Calculating Molar Concentrations from the Mass of Solute: Distilled white vinegar (**Figure 8.21**) is a solution of acetic acid, CH3CO2H, in water. A 0.500-L vinegar solution contains 25.2 g of acetic acid. What is the concentration of the acetic acid solution in units of molarity?

Figure 8.21 Distilled white vinegar is a solution of acetic acid in water.

Solution:

As in previous examples, the definition of molarity is the primary equation used to calculate the quantity sought. In this case, the mass of solute is provided instead of its molar amount, so we must use the solute's molar mass to obtain the amount of solute in moles:

$$M = \frac{mol\ solute}{L\ solution} = \frac{25.2\ g\ CH_3CO_2H * \frac{1\ mol\ CH_3CO_2H}{60.052 g\ CH_3CO_2H}}{0.500\ L\ solution} = 0.839M$$

$$M = \frac{mol\ solute}{L\ solution} = 0.839M$$

$$M = \frac{0.839\ mol\ solute}{1.00\ L\ solution}$$

Check Your Learning

Calculate the molarity of 6.52 g of CoCl2 (128.9 g/mol) dissolved in an aqueous solution with a total volume of 75.0 mL.

Answer: 0.674 M

Example 8.7

Determining the Mass of Solute in a Given Volume of Solution: How many grams of NaCl are contained in 0.250 L of a 5.30-M solution?

Solution:

The volume and molarity of the solution are specified, so the amount (mol) of solute is easily computed as demonstrated in **Example 8.5**:

$$M = \frac{mol\ solute}{L\ solution}$$

$$mol\ slute = M * L\ solution$$

$$mol\ solute = 5.30\ \frac{mol\ NaCl}{L} * 0.250\ l = 1.325\ MOL\ NaCl$$

Finally, this molar amount is used to derive the mass of NaCl:

$$1.325\ mol\ NaCl * \frac{58.44\ g\ NaCl}{mol\ NaCl} = 77.4\ g\ NaCl$$

Check Your Learning

How many grams of CaCl2 (110.98 g/mol) are contained in 250.0 mL of a 0.200-M solution of calcium chloride?

Answer: 5.55 g CaCl2

When performing calculations stepwise, as in **Example 8.7**, it is important to refrain from rounding any intermediate calculation results, which can lead to rounding errors in the final result. In **Example 8.7**, the molar amount of NaCl computed in the first step, 1.325 mol, would be properly rounded to 1.32 mol if it were to be reported; however, although the last digit (5) is not significant, it must be retained as a guard digit in the intermediate calculation. If we had not retained this guard digit, the final calculation for the mass of NaCl would have been 77.1 g, a difference of 0.3 g.In addition to retaining a guard digit for intermediate calculations, we can also avoid rounding errors by performing computations in a single step (see **Example 8.8**). This eliminates intermediate steps so that only the final result is rounded.

Example 8.8

Determining the Volume of Solution Containing a Given Mass of Solute: In **Example 8.6**, we found the typical concentration of vinegar to be 0.839 *M*. What volume of vinegar contains 75.6 g of acetic acid?

Solution:

First, use the molar mass to calculate moles of acetic acid from the given mass:

$$g\ solute * \frac{mol\ solute}{g\ solute} = mol\ solute$$

Then, use the molarity of the solution to calculate the volume of solution containing this molar amount of solute:

$$mol\ solute * \frac{L\ solute}{mol\ solute} = L\ solution$$

Combining these two steps into one yields:

$$g\ solute * \frac{mol\ solute}{g\ solute} * \frac{L\ solution}{mol\ solute} = L\ solution$$

$$75.6\ g\ CH_3CO_2H * \frac{1\ mol\ CH_3CO_2H}{60.05g} * \frac{L\ solution}{0.839\ molCH_3CO_2H} = 1.50\ L\ solution$$

Check Your Learning

What volume of a 1.50-M KBr solution contains 66.0 g KBr?

Answer: 0.370 L

Assessment 8.3:

Saline, a solution of salt (sodium chloride) in water is used to clean wounds and remove contact lenses, as noted in Assessment 8.1. It is prepared by dissolving sodium chloride in water, 0.90% mass/volume. Calculate the molarity of sodium chloride in saline.

Answer:

Strategy: We apply the definition of molarity, given below.

Key equation:

$$M = \frac{mol\ solute}{L\ of\ solution}$$

Consider 1.00 L of solution. The amount of the solute, NaCl, is given by:

$$Mass\ of\ \ sodium\ chloride = \frac{0.90\ g\ sodium\ chloride}{100.0\ mL} \text{x} \frac{1000\ mL}{1\ L} = 9.0g$$

$$Mol\ of\ \ sodium\ chloride = 9.0g\ \text{x} \frac{1\ mol}{58.44\ g} = 0.154\ mol$$

We round this to 0.15 *mol, with* 2 sf. Since the solution has a volume of 1.00 *L*, the molarity of sodium chloride is 0.15 *M*.

8.5 Other Units for Solution Concentrations

By the end of this section, you will be able to:

- Define the concentration units of mass percentage, volume percentage, mass-volume percentage, parts-per- million (ppm), and parts-per-billion (ppb)
- Perform computations relating a solution's concentration and its components' volumes and/or masses using these units
- Perform dilution calculations using the dilution equation

In the previous section, we introduced molarity, a very useful measurement unit for evaluating the concentration of solutions. However, molarity is only one measure of concentration. In this section, we will introduce some other units of concentration that are commonly used in various applications, either for convenience or by convention.

Mass Percentage

Earlier in this chapter, we introduced percent composition as a measure of the relative amount of a given element in a compound. Percentages are also commonly used to express the composition of mixtures, including solutions. The **mass percentage** of a solution component is defined as the ratio of the component's mass to the solution's mass, expressed as a percentage:

$$mass\ percentage = \frac{mass\ of\ component}{mass\ of\ solution} * 100\%$$

We are generally most interested in the mass percentages of solutes, but it is also possible to compute the mass percentage of solvent.Mass percentage is also referred to by similar names such as *percent mass, percent weight, weight/weight percent*, and other variations on this theme. The most common symbol for mass percentage is simply the percent sign, %, although more detailed symbols are often used including %mass, %weight, and (w/w)%. Use of these more detailed symbols can prevent confusion of mass percentages with other types of percentages, such as volume percentages (to be discussed later in this section).

Mass percentages are popular concentration units for consumer products. The label of a typical liquid bleach bottle (**Figure 8.22**) cites the concentration of its active ingredient, sodium hypochlorite (NaOCl), as being 7.4%. A 100.0-g sample of bleach would therefore contain 7.4 g of NaOCl.

Figure 8.22 Liquid bleach is an aqueous solution of sodium hypochlorite (NaOCl). This brand has a concentration of 7.4% NaOCl by mass.

Example 8.9

Calculation of Percent by Mass: A 5.0-g sample of spinal fluid contains 3.75 mg (0.00375 g) of glucose. What is the percent by mass of glucose in spinal fluid?

Solution:

The spinal fluid sample contains roughly 4 mg of glucose in 5000 mg of fluid, so the mass fraction of glucose should be a bit less than one part in 1000, or about 0.1%. Substituting the given masses into the equation defining mass percentage yields:

$$\%\ glucose = \frac{3.75\ mg\ glucose * \frac{1g}{1000\ mg}}{5.0\ g\ spinal\ flui} = 0.075\%$$

The computed mass percentage agrees with our rough estimate (it's a bit less than 0.1%).

Note that while any mass unit may be used to compute a mass percentage (mg, g, kg, oz, and so on), the same unit must be used for both the solute and the solution so that the mass units cancel, yielding a dimensionless ratio. In this case, we converted the units of solute in the numerator from mg to g to match the units in the denominator. We could just as easily have converted the denominator from g to mg instead. As long as identical mass units are used for both solute and solution, the computed mass percentage will be correct.

Check Your Learning

A bottle of a tile cleanser contains 135 g of HCl and 775 g of water. What is the percent by mass of HCl in this cleanser?

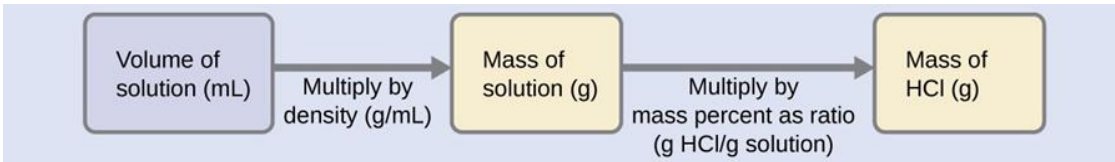
Answer: 14.8%

Example 8.10

Calculations using Mass Percentage:"Concentrated" hydrochloric acid is an aqueous solution of 37.2% HCl that is commonly used as a laboratory reagent. The density of this solution is 1.19 g/mL. What mass of HCl is contained in 0.500 L of this solution?

Solution:

The HCl concentration is near 40%, so a 100-g portion of this solution would contain about 40 g of HCl. Since the solution density isn't greatly different from that of water (1 g/mL), a reasonable estimate of the HCl mass in 500 g (0.5 L) of the solution is about five times greater than that in a 100 g portion, or 5 × 40 = 200 g. In order to derive the mass of solute in a solution from its mass percentage, we need to know the corresponding mass of the solution. Using the solution density given, we can convert the solution's volume to mass, and then use the given mass percentage to calculate the solute mass. This mathematical approach is outlined in this flowchart:



For proper unit cancellation, the 0.500-L volume is converted into 500 mL, and the mass percentage is expressed as a ratio, 37.2 g HCl/g solution:

$$500\ mL\ solution * \frac{1.19 g\ solution}{mL\ solution} * \frac{37.2\ g\ HCl}{100\ g\ solution} = 221\ g\ HCl$$

This mass of HCl is consistent with our rough estimate of approximately 200 g.

Check Your Learning

What volume of concentrated HCl solution contains 125 g of HCl?

Answer: 282 mL

Volume Percentage

Liquid volumes over a wide range of magnitudes are conveniently measured using common and relatively inexpensive laboratory equipment. The concentration of a solution formed by dissolving a liquid solute in a liquid solvent is therefore often expressed as a **volume percentage**, %vol or (v/v)%:

$$Volume\ percentage = \frac{volume\ solute}{volume\ solution} * 100\%$$

Example 8.11

Calculations using Volume Percentage: Rubbing alcohol (isopropanol) is sold as a 70%vol aqueous solution. If the density of isopropyl alcohol is 0.785g/mL, how many grams of alcohol are present in a 355 mL bottle?

Solution:

Per the definition of volume percentage, the isopropanol volume is 70% of the total solution volume. Multiplying the isopropanol volume by its density yields the requested mass:

$$355\ mL\ solution * \frac{70 mL\ isoprophyl\ alcohol}{100\ mL\ solution} * \frac{0.785\ g\ isopropyl\ alcohol}{1\ mL\ isopropyalcohol} = 195\ g\ isopropyl\ alcohol$$

Check your learning

Wine is approximately 12% ethanol (CH3CH2OH) by volume. Ethanol has a molar mass of 46.06 g/mol and a density 0.789 g/mL. How many moles of ethanol are present in a 750-mL bottle of wine?

Answer: 1.5 mol ethanol

Mass-Volume Percentage

"Mixed" percentage units, derived from the mass of solute and the volume of solution, are popular for certain biochemical and medical applications. A **mass-volume percent** is a ratio of a solute's mass to the solution's volume expressed as a percentage. The specific units used for solute mass and solution volume may vary, depending on the solution. For example, physiological saline solution, used to prepare intravenous fluids, has a concentration of 0.9% mass/volume (m/v), indicating that the composition is 0.9 g of solute per 100 mL of solution. The concentration of glucose in blood (commonly referred to as "blood sugar") is also typically expressed in terms of a mass-volume ratio. Though not expressed explicitly as a percentage, its concentration is usually given in milligrams of glucose per deciliter (100 mL) of blood (**Figure 8.23**).

(a) (b)

Figure 8.23"Mixed" mass-volume units are commonly encountered in medical settings. (a) The NaCl concentration of physiological saline is 0.9% (m/v). (b) This device measures glucose levels in a sample of blood. The normal range for glucose concentration in blood (fasting) is around 70–100 mg/dL. (credit a: modification of work by "The National Guard"/Flickr; credit b: modification of work by Biswarup Ganguly)

Assessment 8.4:

Modestly priced, hand-held devices (shown above) are conveniently used at home by individuals to measure blood glucose levels, which are typically expressed in units of *mg/dL*. An individual measure a healthy glucose level of 100 mg/dL after fasting for 12 hours. Convert this to units of *mmol/L*. For this problem, assume 3 sf.

Answer:

Strategy: We apply the definition of molarity, given below. We also note that glucose has the molecular formula, $C_6H_{12}O_6$.

Key equations:

$$M = \frac{mol\ solute}{L\ of\ solution}$$

1 *mol* = 1000 *mmol*

$$M = \frac{100\ mg\ C_6H_{12}O_6}{100.0\ mL} \text{ x } \frac{1\ g\ C_6H_{12}O_6}{1000\ mg\ C_6H_{12}O_6} \text{ x } \frac{1\ mol\ C_6H_{12}O_6}{180.156\ g\ C_6H_{12}O_6} \text{ x } \frac{1000\ mL}{1\ L} = 5.55 \text{ x } 10^{-3} M$$

We can replace 10^{-3} M with *mM*. Thus, the concentration is 5.55 *mM*.

Parts per Million and Parts per Billion

Very low solute concentrations are often expressed using appropriately small units such as **parts per million (ppm)** or **parts per billion (ppb)**. Like percentage ("part per hundred") units, ppm and ppb may be defined in terms of masses, volumes, or mixed mass-volume units. There are also ppm and ppb units defined with respect to numbers of atoms and molecules.

The mass-based definitions of ppm and ppb are given here:

$$\text{ppm} = \frac{mass\ solute}{mass\ solution} * 10^6\ ppm$$

$$\text{ppb} = \frac{mass\ solute}{mass\ solution} * 10^9\ \text{ppb}$$

Both ppm and ppb are convenient units for reporting the concentrations of pollutants and other trace contaminants in water. Concentrations of these contaminants are typically very low in treated and natural waters, and their levels cannot exceed relatively low concentration thresholds without causing adverse effects on health and wildlife. For example, the EPA has identified the maximum safe level of fluoride ion in tap water to be 4 ppm. Inline water filters are designed to reduce the concentration of fluoride and several other trace-level contaminants in tap water (**Figure 8.24**).

(a) (b)

Figure 8.24 (a) In some areas, trace-level concentrations of contaminants can render unfiltered tap water unsafe for drinking and cooking. (b) Inline water filters reduce the concentration of solutes in tap water. (credit a: modification of work by Jenn Durfey; credit b: modification of work by "vastateparkstaff"/Wikimedia commons)

Example 8.12

Calculation of Parts per Million and Parts per Billion Concentrations: According to the EPA, when the concentration of lead in tap water reaches 15 ppb, certain remedial actions must be taken. What is this concentration in ppm? At this concentration, what mass of lead (µg) would be contained in a typical glass of water (300 mL)?

Solution:

The definitions of the ppm and ppb units may be used to convert the given concentration from ppb to ppm. Comparing these two-unit definitions shows that ppm is 1000 times greater than ppb (1 ppm = 103 ppb). Thus:

$$15\ ppb * \frac{1ppm}{10^3 ppb} = 0.015\ ppm$$

The definition of the ppb unit may be used to calculate the requested mass if the mass of the solution is provided. However, only the volume of solution (300 mL) is given, so we must use the density to derive the corresponding mass.

We can assume the density of tap water to be roughly the same as that of pure water (~1.00 g/mL), since the concentrations of any dissolved substances should not be very large. Rearranging the equation defining the ppb unit and substituting the given quantities yields:

$$ppb = \frac{mass\ solute}{mass\ solution} * 10^9 ppb$$

$$mass\ solute = \frac{ppb * mass\ solution}{10^9 ppb}$$

$$mass\ solute = \frac{15\ ppb * 300\ mL * \frac{1.00\ g}{mL}}{10^9\ ppb} = 4.5 * 10^{-6} g$$

Finally, convert this mass to the requested unit of micrograms

$$4.5 * 10^{-6} g * \frac{1 \mu g}{10^{-6} g} = 4.5 \mu g$$

Check Your Learning

A 50.0-g sample of industrial wastewater was determined to contain 0.48 mg of mercury. Express the mercury concentration of the wastewater in ppm and ppb units.

Answer: 9.6 ppm, 9600 ppb

Assessment 8.5:

Intravenous medications are sometimes administered in a solution of glucose, 5.0% by mass in water. A solution whose volume is 200.0 mL contains 10.0 mg of glucose ($C_6H_{12}O_6$). What is the concentration of glucose in ppm, assuming that the density of the solution is 1.00 g/mL?

Answer:

Strategy: We apply the definition, given below.

Key Equation:

$$ppm = \frac{mass\ of\ solute}{mass\ of\ solution} \text{ x } 10^6$$

$$ppm\ C_6H_{12}O_6 = 10.0 mg\ C_6H_{12}O_6 \text{ x } \frac{1\ g\ C_6H_{12}O_6}{1000\ mg\ C_6H_{12}O_6} \text{ x } \frac{1}{200.0\ mL\ solution} \text{ x } \frac{1.00\ mL\ solution}{1.00\ g\ solution} \text{ x } 10^6 = 50.0\ ppm$$

(3 sf)

Assessment 8.6:

(From wqa.org) Cadmium occurs naturally in zinc, lead and copper ores, in coal and other fossil fuels, in shales and is released during volcanic action. These deposits can serve as sources to ground and surface waters. Cadmium (in the form of Cd^{2+} ions) has the chronic potential to cause kidney, liver, bone and blood damage from long- term exposure at levels above the maximum contaminant level (MCL), 0.005 mg/L. Convert this to ppb.

Answer:

Strategy: We apply the definition, given below. Since the solution is very dilute, its density is 1.00 g/mL, or the mass of 1.00 L of solution is 1000 g.

Key Equation:

$$ppb = \frac{mass\ of\ solute}{mass\ of\ solution} \text{ x } 10^9$$

$$ppb = \frac{mass\ of\ Cd^{2+}}{mass\ of\ solution} \text{ x } 10^9 = \frac{0.005 mg Cd^{2+}}{1\ L\ solution} \text{ x } \frac{1\ g\ Cd^{2+}}{1000\ mg\ Cd^{2+}} \text{ x } \frac{1\ L\ solution}{1000\ g\ solution} \text{ x} 10^9 = 5\ ppb$$

Dilution of Solutions

Dilution is the process whereby the concentration of a solution is lessened by the addition of solvent. For example, we might say that a glass of iced tea becomes increasingly diluted as the ice melts. The water from the melting ice increases the volume of the solvent (water) and the overall volume of the solution (iced tea), thereby reducing the relative concentrations of the solutes that give the beverage its taste (**Figure 8.24**).

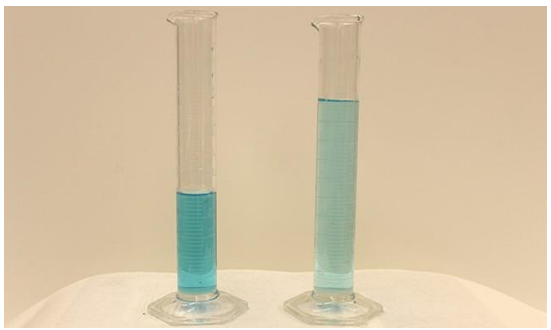


Figure 8.24 Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott)

Dilution is also a common means of preparing solutions of a desired concentration. By adding solvent to a measured portion of a more concentrated *stock solution*, we can achieve a particular concentration. For example, commercial pesticides are typically sold as solutions in which the active ingredients are far more concentrated than is appropriate for their application. Before they can be used on crops, the pesticides must be diluted. This is also a very common practice for the preparation of a number of common laboratory reagents (**Figure 8.25**).

Figure 8.25 A solution of $KMnO_4$ is prepared by mixing water with 4.74 g of $KMnO_4$ in a flask. (credit: modification of work by Mark Ott)

A simple mathematical relationship can be used to relate the volumes and concentrations of a solution before and after the dilution process. According to the definition of molarity, the molar amount of solute in a solution is equal to the

product of the solution's molarity and its volume in liters:

$$n = ML$$

Expressions like these may be written for a solution before and after it is diluted:

$$n_1 = M_1 L_1$$

$$n_2 = M_2 L_2$$

where the subscripts "1" and "2" refer to the solution before and after the dilution, respectively. Since the dilution process *does not change the amount of solute in the solution*, $n_1 = n_2$. Thus, these two equations may be set equal to one another:

$$M_1 L_1 = M_2 L_2$$

This relation is commonly referred to as the dilution equation. Although we derived this equation using molarity as the unit of concentration and liters as the unit of volume, other units of concentration and volume may be used, so long as the units properly cancel per the factor-label method. Reflecting this versatility, the dilution equation is often written in the more general form:

$$C_1 V_1 = C_2 V_2$$

where C and V are concentration and volume, respectively.

Link to Learning

Use the **simulation (http://openstaxcollege.org/l/16Phetsolvents)** to explore the relations between solute amount, solution volume, and concentration and to confirm the dilution equation.

Example 8.13

Determining the Concentration of a Diluted Solution: If 0.850 L of a 5.00-*M* solution of copper nitrate, $Cu(NO_3)_2$, is diluted to a volume of 1.80 L by the addition of water, what is the molarity of the diluted solution?

Solution:

We are given the volume and concentration of a stock solution, V_1 and C_1, and the volume of the resultant diluted solution, V_2. We need to find the concentration of the diluted solution, C_2. We thus rearrange the dilution equation in order to isolate C_2:

$$C_1V_1 = C_2V_2$$

$$C_2 = \frac{C_1V_1}{V_2}$$

Since the stock solution is being diluted by more than two-fold (volume is increased from 0.85 L to 1.80 L), we would expect the diluted solution's concentration to be less than one-half 5 *M*. We will compare this ballpark estimate to the

calculated result to check for any gross errors in computation (for example, such as an improper substitution of the given quantities). Substituting the given values for the terms on the right side of this equation yields:

$$C_2 = \frac{0.850\ L * 5.00\ \frac{mol}{L}}{1.80L} = 2.36\ M$$

Check Your Learning

What is the concentration of the solution that results from diluting 25.0 mL of a 2.04-M solution of CH3OH to 500.0 mL?

Answer: 0.102 M CH3OH

Assessment 8.7:

The solution of HCl (37.2% by mass, density of 1.19 *g*/*mL*), described above, has a concentration of 12.1 *M*. A student uses this solution to prepare 1.00 *L* of a dilute solution in which the concentration of HCl mirrors that of the dilute HCl present in the stomach at a concentration of 0.010 *M* (which in turn corresponds to a pH of 2.00). What volume of the concentrated acid is needed?

Answer:

Strategy: We recognize this as a dilution problem.

Key equation:

$C_1V_1 = C_2V_2$

We can rewrite this as:

$C_{concentraed}V_{concentraed} = C_{dilute}V_{dilute}$

Solving for $V_{\text{concentrated}}$,

$$V_{concentraed} = \frac{C_{dilute}V_{dilute}}{C_{concentraed}} = \frac{(0.010M)(1.00\ L)}{12.1M} = 8.3 \text{ x } 10^{-4}\ L = 0.83\ mL \text{ (2sf)}$$

Example 8.14

Volume of a Diluted Solution: What volume of 0.12 *M* HBr can be prepared from 11 mL (0.011 L) of 0.45 *M* HBr?

Solution:

We are given the volume and concentration of a stock solution, V_1 and C_1, and the concentration of the resultant diluted solution, C_2. We need to find the volume of the diluted solution, V_2. We thus rearrange the dilution equation in order to isolate V_2:

$$C_1V_1 = C_2V_2$$

$$V_2 = \frac{C_1V_1}{C_2}$$

Since the diluted concentration (0.12 *M*) is slightly more than one-fourth the original concentration (0.45 *M*), we would expect the volume of the diluted solution to be roughly four times the original volume, or around 44 mL. Substituting the given values and solving for the unknown volume yields:

$$V_2 = \frac{0.45\ M * 0.011\ L}{0.12\ M}$$

V_2 = 0.041 L

The volume of the 0.12-*M* solution is 0.041 L (41 mL). The result is reasonable and compares well with our rough estimate.

Check Your Learning

A laboratory experiment calls for 0.125 *M* HNO_3. What volume of 0.125 *M* HNO_3 can be prepared from 0.250 L of 1.88 *M* HNO_3?

Answer: 3.76 L

Example 8.15

Volume of a Concentrated Solution Needed for Dilution: What volume of 1.59 M KOH is required to prepare 5.00 L of 0.100 M KOH?

Solution:

We are given the concentration of a stock solution, C1, and the volume and concentration of the resultant diluted solution, V2 and C2. We need to find the volume of the stock solution, V1. We thus rearrange the dilution equation in order to isolate V1:

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1}$$

(1.59 M), we would expect the volume of the stock solution to be about one-sixteenth that of the diluted solution, or around 0.3 liters. Substituting the given values and solving for the unknown volume yields:

$$V_1 = \frac{0.100\ M * 5.00L}{1.59\ M}$$

V_1 = 0.314 L

Thus, we would need 0.314 L of the 1.59-M solution to prepare the desired solution. This result is consistent with our rough estimate.

Check Your Learning

What volume of a 0.575-M solution of glucose, C6H12O6, can be prepared from 50.00 mL of a 3.00-M glucose solution?

Answer: 0.261 L

8.6 Colligative Properties

By the end of this section, you will be able to:

8.6.1 Explain the process of osmosis and describe how it is applied industrially and in nature

The properties of a solution are different from those of either the pure solute(s) or solvent. Many solution properties are dependent upon the chemical identity of the solute. Compared to pure water, a solution of hydrogen chloride is more acidic, a solution of ammonia is more basic, a solution of sodium chloride is more dense, and a solution of

sucrose is more viscous. There are a few solution properties, however, that depend *only* upon the total concentration of solute species, regardless of their identities. These **colligative properties** include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure. This small set of properties is of central importance to many natural phenomena and technological applications, as will be described in this module.

Osmosis and Osmotic Pressure of Solutions

A number of natural and synthetic materials exhibit *selective permeation*, meaning that only molecules or ions of a certain size, shape, polarity, charge, and so forth, are capable of passing through (permeating) the material. Biological cell membranes provide elegant examples of selective permeation in nature, while dialysis tubing used to remove metabolic wastes from blood is a more simplistic technological example. Regardless of how they may be fabricated, these materials are generally referred to as **semipermeable membranes**.

Consider the apparatus illustrated in **Figure 8.26**, in which samples of pure solvent and a solution are separated by a membrane that only solvent molecules may permeate. Solvent molecules will diffuse across the membrane in both directions. Since the concentration of *solvent* is greater in the pure solvent than the solution, these molecules will diffuse from the solvent side of the membrane to the solution side at a faster rate than they will in the reverse direction. The result is a net transfer of solvent molecules from the pure solvent to the solution. Diffusion-driven transfer of solvent molecules through a semipermeable membrane is a process known as **osmosis**.

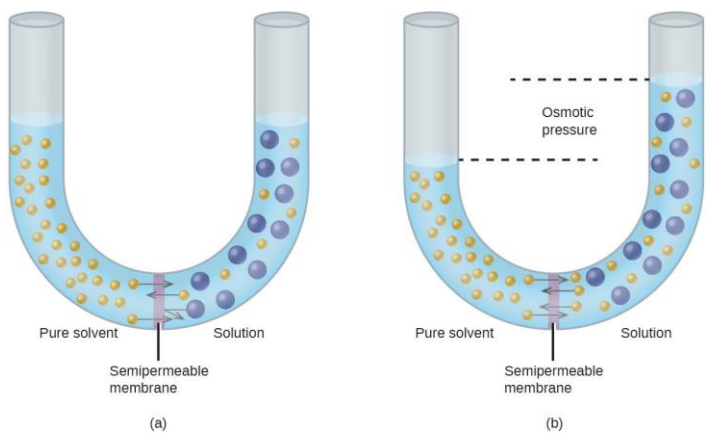


Figure 8.26 Osmosis results in the transfer of solvent molecules from a sample of low (or zero) solute concentration to a sample of higher solute concentration.

When osmosis is carried out in an apparatus like that shown in **Figure 8.26**, the volume of the solution increases as it becomes diluted by accumulation of solvent. This causes the level of the solution to rise, increasing its hydrostatic pressure (due to the weight of the column of solution in the tube) and resulting in a faster transfer of solvent molecules back to the pure solvent side. When the pressure reaches a value that yields a reverse solvent transfer rate equal to the osmosis rate, bulk transfer of solvent ceases. This pressure is called the **osmotic pressure (Π)** of the solution. The osmotic pressure of a dilute solution is related to its solute molarity, M, and absolute temperature, T, according to the equation:

$$n = MRT$$

where R is the universal gas constant.

Example 8.16

Calculation of Osmotic Pressure: What is the osmotic pressure (atm) of a 0.30 M solution of glucose in water that is used for intravenous infusion at body temperature, 37 °C?

Solution:

We can find the osmotic pressure, Π, using the formula Π = MRT, where T is on the Kelvin scale (310 K) and the value of R is expressed in appropriate units (0.08206 L atm/mol K).

$$n = MRT$$

$$= 0.03\frac{\text{mol}}{\text{l}} * 0.08206\frac{\text{l atm}}{\text{mol K}} * 310\text{ K}$$

$$= 7.6\text{ atm}$$

Check Your Learning

What is the osmotic pressure (atm) a solution with a volume of 0.750 L that contains 5.0 g of methanol, CH3OH, in water at 37 °C?

Answer: 5.3 atm

If a solution is placed in an apparatus like the one shown in **Figure 8.27**, applying pressure greater than the osmotic pressure of the solution reverses the osmosis and pushes solvent molecules from the solution into the pure solvent. This technique of reverse osmosis is used for large-scale desalination of seawater and on smaller scales to produce high-purity tap water for drinking.

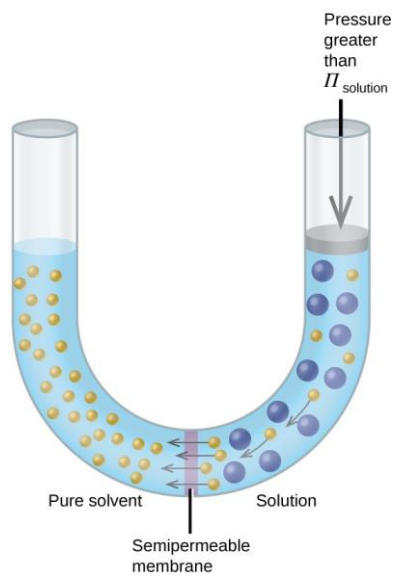


Figure 8.27 Applying a pressure greater than the osmotic pressure of a solution will reverse osmosis. Solvent molecules from the solution are pushed into the pure solvent.

Chemistry in Everyday Life: Reverse Osmosis Water Purification

In the process of osmosis, diffusion serves to move water through a semipermeable membrane from a less concentrated solution to a more concentrated solution. Osmotic pressure is the amount of pressure that must be applied to the more

concentrated solution to cause osmosis to stop. If greater pressure is applied, the water will go from the more concentrated solution to a less concentrated (more pure) solution. This is called reverse osmosis. Reverse osmosis (RO) is used to purify water in many applications, from desalination plants in coastal cities, to water-purifying machines in grocery stores (**Figure 8.28**), and smaller reverse- osmosis household units. With a hand-operated pump, small RO units can be used in third-world countries, disaster areas, and in lifeboats. Our military forces have a variety of generator-operated RO units that can be disaster areas, and in lifeboats. Our military forces have a variety of generator- operated RO units that can be transported in vehicles to remote locations.

(a) (b)

Figure 8.28 Reverse osmosis systems for purifying drinking water are shown here on (a) small and (b) large scales. (credit a: modification of work by Jerry Kirkhart; credit b: modification of work by Willard J. Lathrop)

Examples of osmosis are evident in many biological systems because cells are surrounded by semipermeable membranes. Carrots and celery that have become limp because they have lost water can be made crisp again by placing them in water. Water moves into the carrot or celery cells by osmosis. A cucumber placed in a concentrated salt solution loses water by osmosis and absorbs some salt to become a pickle. Osmosis can also affect animal cells. Solute concentrations are particularly important when solutions are injected into the body. Solutes in body cell fluids and blood serum give these solutions an osmotic pressure of approximately 7.7 atm. Solutions injected into the body must have the same osmotic pressure as blood serum; that is, they should be **isotonic** with blood serum. If a less concentrated solution, a **hypotonic** solution, is injected in sufficient quantity to dilute the blood serum, water from the diluted serum passes into the blood cells by osmosis, causing the cells to expand and rupture. This process is called **hemolysis**. When a more concentrated solution, a **hypertonic** solution, is injected, the cells lose water to the more concentrated solution, shrivel, and possibly die in a process called **crenation**. These effects are illustrated in **Figure 8.29**.

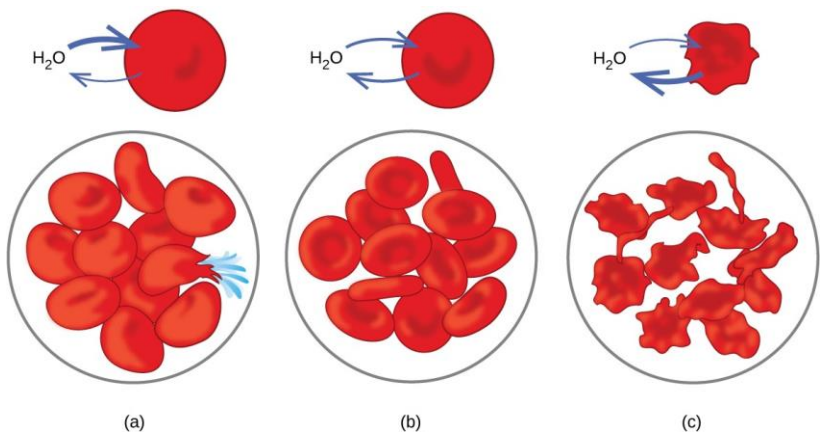


Figure 8.29 Red blood cell membranes are water permeable and will (a) swell and possibly rupture in a hypotonic solution; (b) maintain normal volume and shape in an isotonic solution; and (c) shrivel and possibly die in a hypertonic solution. (credit a/b/c: modifications of work by "LadyofHats"/Wikimedia commons)

Determination of Molar Masses

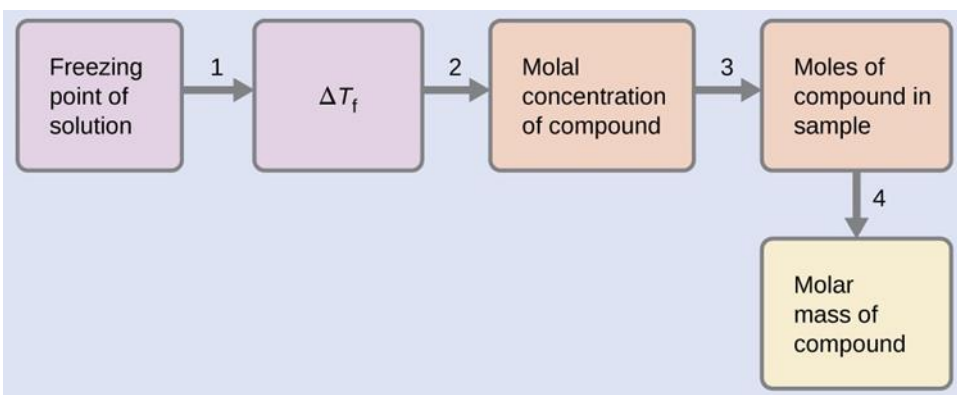
Osmotic pressure and changes in freezing point, boiling point, and vapor pressure are directly proportional to the concentration of solute present. Consequently, we can use a measurement of one of these properties to determine the molar mass of the solute from the measurements.

Example 8.17

Determination of a Molar Mass from a Freezing Point Depression: A solution of 4.00 g of a nonelectrolyte dissolved in 55.0 g of benzene is found to freeze at 2.32 °C. What is the molar mass of this compound?

Solution:

We can solve this problem using the following steps.



Step 1. Determine the change in freezing point from the observed freezing point and the freezing point of pure benzene (**Table 8.2**).

$\Delta Tf = 5.5\ °C - 2.32\ °C = 3.2\ °C$

Step 2. Determine the molal concentration from Kf, the freezing point depression constant for

benzene (**Table 8.2**), and ΔTf.

$$T_f = k_f m$$

$$m = \frac{T_f}{K_r} = \frac{3.2°C}{5.12°C\ m^{-1}} = 0.63\ m$$

Step 3. Determine the number of moles of compound in the solution from the molal concentration and the mass of solvent used to make the solution

$$moles\ of\ solte = \frac{0.63\ mol\ solute}{1.00\ kg\ solvent} * 0.0550\ kg\ solvent = 0.035\ mol$$

Step 4. Determine the molar mass from the mass of the solute and the number of moles in that mass.

$$Molar\ mass = \frac{4.00\ g}{0.035\ mol} = 1.1 * 10^2\ \frac{g}{mol}$$

Check Your Learning

A solution of 35.7 g of a nonelectrolyte in 220.0 g of chloroform has a boiling point of 64.5 °C. What is the molar mass of this compound?

Answer: 1.8 × 102 g/mol

Example 8.18

Determination of a Molar Mass from Osmotic Pressure: A 0.500 L sample of an aqueous solution containing 10.0 g of hemoglobin has an osmotic pressure of 5.9 torr at 22 °C. What is the molar mass of hemoglobin?

Solution:

Here is one set of steps that can be used to solve the problem:

Step 1. *Convert the osmotic pressure to atmospheres, then determine the molar concentration from the osmotic pressure.*

$$n = \frac{5.9\ torr * 1atm}{760\ torr} = 7.8 * 10^{-3} atm$$

$$n = MRT$$

$$M = \frac{N}{RT} = \frac{7.8 * 10^{-3} atm}{\left(0.08206 \frac{L\ atm}{mol\ K}\right)(295\ K\ )} = 3.2 * 10^{-4}\ M$$

Step 2. *Determine the number of moles of hemoglobin in the solution from the concentration and the volume of the solution.*

$$\text{moles of hemoglobin} = \frac{3.2 * 10^{-4}\text{mol}}{1\ \text{L solution}} * 0.500\ \text{L solution} = 1.6 * 10^{-4}\text{mol}$$

Step 3. *Determine the molar mass from the mass of hemoglobin and the number of moles in that mass.*

$$\text{molar mass} = \frac{10.0\ g}{1.6 * 10^{-4} mol} = 6.2 * 10^4\ \frac{g}{mol}$$

Check Your Learning

What is the molar mass of a protein if a solution of 0.02 g of the protein in 25.0 mL of solution has an osmotic pressure of 0.56 torr at 25 °C?

Answer: 3 × 104 g/mol

Assessment 8.8:

Which of the following aqueous solutions contains the **largest** total concentration of ions, assuming complete dissociation?

(a) 0.010 *m* sodium chloride

(b) 0.010 *m* calcium chloride

(c) 0.010 *m* aluminum chloride

(d) 0.010 *m* aluminum sulfate

Answer:

Strategy: Each of the solutions contains 0.010 mole of solute per kg of solvent, based on the definition of molality. We will assume that in a dilute solution, the compounds will fully dissociate in water.

The ions in each of the compounds are listed below. The total number of ions per formula unit is also listed.

Compound	**Formula**	**Ions in 1 formula unit**	**Total number of ions per formula unit**
Sodium chloride	NaCl	1 Na^+; 1 Cl^-	2
Calcium chloride	$CaCl_2$	1 $Ca2^+$; 2 Cl^-	3
Aluminum chloride	$AlCl_3$	1 Al^{3+}; 3 Cl^-	4
Aluminum sulfate	$Al_2(SO_4)_3$	2 Al^{3+}; 3 SO_4^{2-}	5

We note that aluminum sulfate produces the most number of ions per formula unit. Thus, the highest concentration of ions will result from aluminum sulfate, and the effective molality due to all ions will be 0.050 *m*.

Assessment 8.9:

(a). A grape is placed in water. Over time, _____.

(i) The grape expands

(ii) The size of the grape remains the same

(iii) The grape shrinks

(b). A cucumber is placed in brine, a concentrated aqueous solution of sodium chloride. As a result of osmosis, over time, _____.

(a) The cucumber expands

(b) The size of the cucumber remains the same

(c) The cucumber shrinks

Answer:

Strategy: Both of these are real-life examples of osmosis across the semi-permeable membranes of the grape and cucumber.

(a). We can consider the inside of a grape as a dilute aqueous solution. The outside is pure water. Water will travel from a region of higher vapor pressure (outside the grape) to lower vapor pressure (inside the grape) which thus expands over time.

(b). We can consider the inside of a cucumber as a dilute aqueous solution. The outside is a concentrated brine (NaCl)

solution. Water will travel from a region of higher vapor pressure (inside the cucumber) to lower vapor pressure (brine outside the cucumber). The cucumber thus shrinks over time.

8.7 Colloids

By the end of this section, you will be able to:

- Describe the composition and properties of colloidal dispersions
- List and explain several technological applications of colloids

As a child, you may have made suspensions such as mixtures of mud and water, flour and water, or a suspension of solid pigments in water, known as tempera paint. These suspensions are heterogeneous mixtures composed of relatively large particles that are visible (or that can be seen with a magnifying glass). They are cloudy, and the suspended particles settle out after mixing. On the other hand, when we make a solution, we prepare a homogeneous mixture in which no settling occurs and in which the dissolved species are molecules or ions. Solutions exhibit completely different behavior from suspensions. A solution may be colored, but it is transparent, the molecules or ions are invisible, and they do not settle out on standing. A group of mixtures called **colloids** (or **colloidal dispersions**) exhibit properties intermediate between those of suspensions and solutions (**Figure 8.30**). The particles in a colloid are larger than most simple molecules; however, colloidal particles are small enough that they do not settle out upon standing.

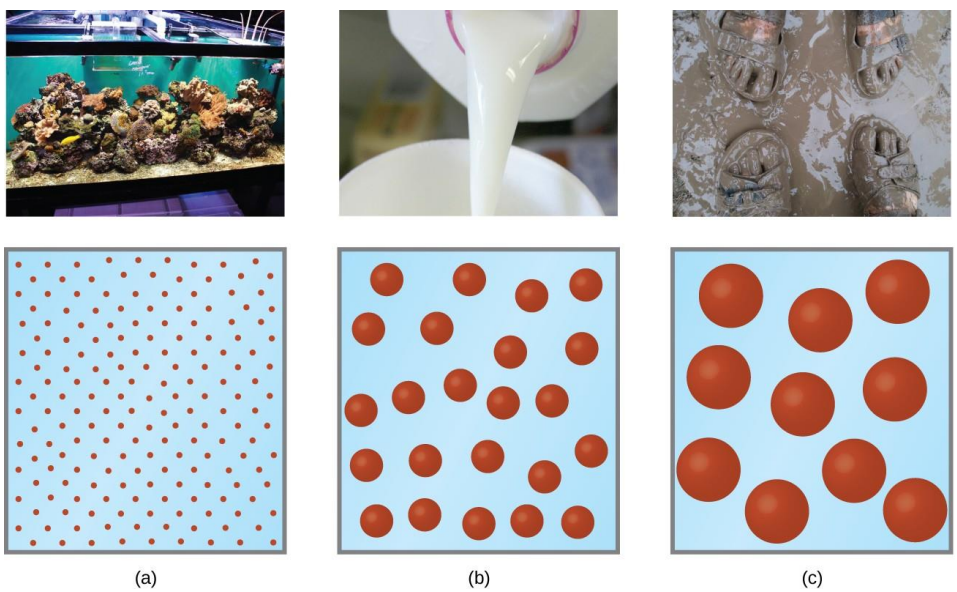


Figure 8.30 (a) A solution is a homogeneous mixture that appears clear, such as the saltwater in this aquarium. (b) In a colloid, such as milk, the particles are much larger but remain dispersed and do not settle. (c) A suspension, such as mud, is a heterogeneous mixture of suspended particles that appears cloudy and in which the particles can settle. (credit a photo: modification of work by Adam Wimsatt; credit b photo: modification of work by Melissa Wiese; credit c photo: modification of work by Peter Burgess)

The particles in a colloid are large enough to scatter light, a phenomenon called the **Tyndall effect**. This can make colloidal mixtures appear cloudy or opaque, such as the searchlight beams shown in **Figure 8.31**. Clouds are colloidal mixtures. They are composed of water droplets that are much larger than molecules, but that are small enough that they do not settle out.

Figure 8.31 The paths of searchlight beams are made visible when light is scattered by colloidal-size particles in the air (fog, smoke, etc.). (credit: "Bahman"/Wikimedia Commons)

The term "colloid"—from the Greek words *kolla*, meaning "glue," and *eidos*, meaning "like"—was first used in 1861 by Thomas Graham to classify mixtures such as starch in water and gelatin. Many colloidal particles are aggregates of hundreds or thousands of molecules, but others (such as proteins and polymer molecules) consist of a single extremely large molecule. The protein and synthetic polymer molecules that form colloids may have molecular masses ranging from a few thousand to many million atomic mass units.

Analogous to the identification of solution components as "solute" and "solvent," the components of a colloid are likewise classified according to their relative amounts. The particulate component typically present in a relatively minor amount is called the **dispersed phase** and the substance or solution throughout which the particulate is dispersed is called the **dispersion medium**. Colloids may involve virtually any combination of physical states (gas in liquid, liquid in solid, solid in gas, etc.), as illustrated by the examples of colloidal systems given in **Table 8.4**.

Dispersed Phase	Dispersion Medium	Common Examples	Name
solid	gas	smoke, dust	—
solid	liquid	starch in water, some inks, paints, milk of magnesia	sol
solid	solid	some colored gems, some alloys	—
liquid	gas	clouds, fogs, mists, sprays	aerosol
liquid	liquid	milk, mayonnaise, butter	emulsion
liquid	solid	jellies, gels, pearl, opal (H_2O in SiO_2)	gel
gas	liquid	foams, whipped cream, beaten egg whites	foam
gas	solid	Pumice, floating soaps	

Table 8.4 Examples of Colloidal Systems

Preparation of Colloidal Systems

We can prepare a colloidal system by producing particles of colloidal dimensions and distributing these particles throughout a dispersion medium. Particles of colloidal size are formed by two methods:

1. Dispersion methods: that is, by breaking down larger particles. For example, paint pigments are produced by dispersing large particles by grinding in special mills.
2. Condensation methods: that is, growth from smaller units, such as molecules or ions. For example, clouds form when water molecules condense and form very small droplets.

A few solid substances, when brought into contact with water, disperse spontaneously and form colloidal systems. Gelatin, glue, starch, and dehydrated milk powder behave in this manner. The particles are already of colloidal size; the water simply disperses them. Powdered milk particles of colloidal size are produced by dehydrating milk spray. Some atomizers produce colloidal dispersions of a liquid in air.

We can prepare an **emulsion** by shaking together or blending two immiscible liquids. This breaks one liquid into droplets of colloidal size, which then disperse throughout the other liquid. Oil spills in the ocean may be difficult to clean up, partly because wave action can cause the oil and water to form an emulsion. In many emulsions, however, the dispersed phase tends to coalesce, form large drops, and separate. Therefore, emulsions are usually stabilized by an **emulsifying agent**, a substance that inhibits the coalescence of the dispersed liquid. For example, a little soap will stabilize an emulsion of kerosene in water. Milk is an emulsion of butterfat in water, with the protein casein as the emulsifying agent. Mayonnaise is an emulsion of oil in vinegar, with egg yolk components as the emulsifying agents.

Condensation methods form colloidal particles by aggregation of molecules or ions. If the particles grow beyond the colloidal size range, drops or precipitates form, and no colloidal system results. Clouds form when water molecules aggregate and form colloid-sized particles. If these water particles coalesce to form adequately large water drops of liquid water or crystals of solid water, they settle from the sky as rain, sleet, or snow. Many condensation methods involve chemical reactions. We can prepare a red colloidal suspension of iron(III) hydroxide by mixing a concentrated solution of iron(III) chloride with hot water:

$$Fe^{3+}(aq) + 3Cl^{-}(aq) + 6H_2O(l) \longrightarrow Fe(OH)_3(s) + H_3O^{+}(aq) + 3Cl^{-}(aq).$$

A colloidal gold sol results from the reduction of a very dilute solution of gold(III) chloride by a reducing agent such as formaldehyde, tin(II) chloride, or iron(II) sulfate:

$$Au^{3+} + 3e^{-} \longrightarrow Au$$

Some gold sols prepared in 1857 are still intact (the particles have not coalesced and settled), illustrating the long-term stability of many colloids.

Soaps and Detergents

Pioneers made soap by boiling fats with a strongly basic solution made by leaching potassium carbonate, K_2CO_3, from wood ashes with hot water. Animal fats contain polyesters of fatty acids (long-chain carboxylic acids). When animal fats are treated with a base like potassium carbonate or sodium hydroxide, glycerol and salts of fatty acids such as palmitic, oleic, and stearic acid are formed. The salts of fatty acids are called *soaps*. The sodium salt of stearic acid,

sodium stearate, has the formula $C_{17}H_{35}CO_2Na$ and contains an uncharged nonpolar hydrocarbon chain, the $C_{17}H_{35}$— unit, and an ionic carboxylate group, the — CO_2^- unit (**Figure 8.32**).

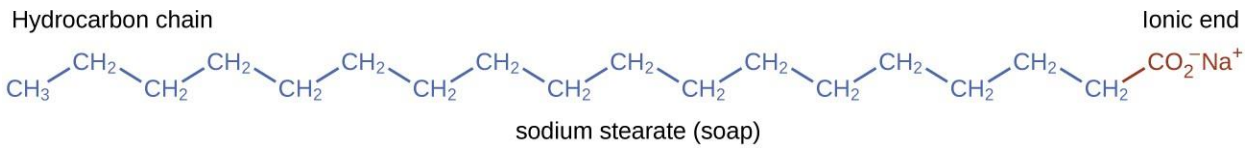


Figure 8.32 Soaps contain a nonpolar hydrocarbon end (blue) and an ionic end (red). The ionic end is a carboxylate group. The length of the hydrocarbon end can vary from soap to soap.

Detergents (soap substitutes) also contain nonpolar hydrocarbon chains, such as $C_{12}H_{25}$—, and an ionic group, such as a sulfate— OSO_3^-, or a sulfonate— SO_3^- (**Figure 8.33**). Soaps form insoluble calcium and magnesium compounds in hard water; detergents form water-soluble products—a definite advantage for detergents.

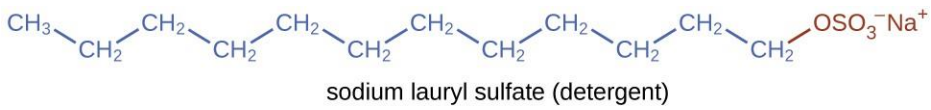


Figure 8.33 Detergents contain a nonpolar hydrocarbon end (blue) and an ionic end (red). The ionic end can be either a sulfate or a sulfonate. The length of the hydrocarbon end can vary from detergent to detergent.

The cleaning action of soaps and detergents can be explained in terms of the structures of the molecules involved. The hydrocarbon (nonpolar) end of a soap or detergent molecule dissolves in, or is attracted to, nonpolar substances such as oil, grease, or dirt particles. The ionic end is attracted by water (polar), illustrated in **Figure 8.34**. As a result, the soap or detergent molecules become oriented at the interface between the dirt particles and the water so they act as a kind of bridge between two different kinds of matter, nonpolar and polar. Molecules such as this are termed **amphiphilic** since they have both a hydrophobic ("water-fearing") part and a hydrophilic ("water-loving") part. As a consequence, dirt particles become suspended as colloidal particles and are readily washed away.

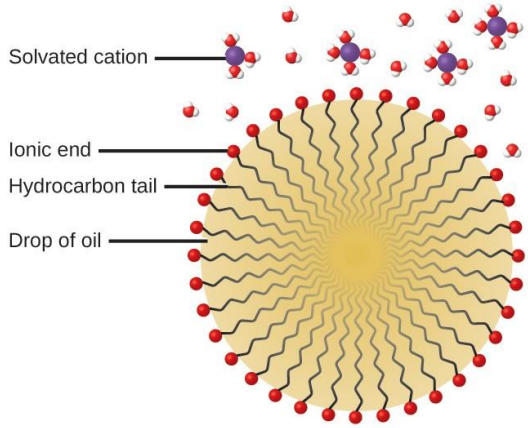


Figure 8.34 This diagrammatic cross section of an emulsified drop of oil in water shows how soap or detergent acts as an emulsifier.

Chemistry in Everyday Life: Deepwater Horizon Oil Spill

The blowout of the Deepwater Horizon oil drilling rig on April 20, 2010, in the Gulf of Mexico near Mississippi began the largest marine oil spill in the history of the petroleum industry. In the 87 days following the blowout, an estimated 4.9 million barrels (210 million gallons) of oil flowed from the ruptured well 5000 feet below the water's surface. The well was finally declared sealed on September 19, 2010. Crude oil is immiscible with and less dense than water, so the spilled oil rose to the surface of the water. Floating booms, skimmer ships, and controlled burns were used to remove oil from the water's surface in an attempt to protect beaches and wetlands along the Gulf coast. In addition to removal of the oil, attempts were also made to lessen its environmental impact by rendering it "soluble" (in the loose sense of the term) and thus allowing it to be diluted to hopefully less harmful levels by the vast volume of ocean water. This approach used 1.84 million gallons of the oil dispersant Corexit 9527, most of which was injected underwater at the site of the leak, with small amounts being sprayed on top of the spill. Corexit 9527 contains 2-butoxyethanol ($C_6H_{14}O_2$), an amphiphilic molecule whose polar and nonpolar ends are useful for emulsifying oil into small droplets, increasing the surface area of the oil and making it more available to marine bacteria for digestion (**Figure 8.35**). While this approach avoids many of the immediate hazards that bulk oil poses to marine and coastal ecosystems, it introduces the possibility of long-term effects resulting from the introduction of the complex and potential toxic components of petroleum into the ocean's food chain. A number of organizations are involved in monitoring the extended impact of this oil spill, including the National Oceanic and Atmospheric Administration (visit this **website (http://openstaxcollege.org/l/16gulfspill)** for additional details).

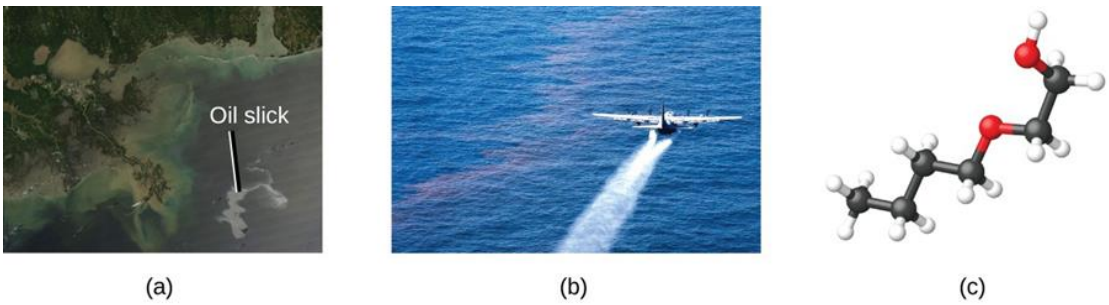


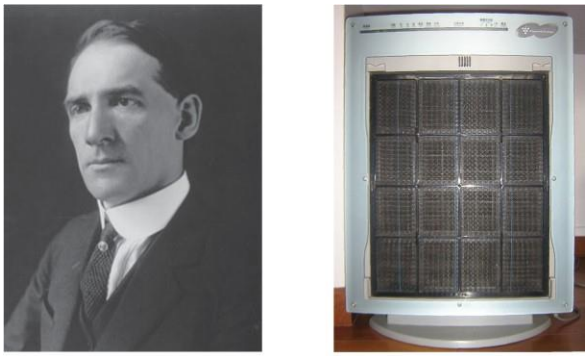
Figure 8.35 (a) This NASA satellite image shows the oil slick from the Deepwater Horizon spill. (b) A US Air Force plane sprays Corexit, a dispersant. (c) The molecular structure of 2-butoxyethanol is shown. (credit a: modification of work by "NASA, FT2, demis.nl"/Wikimedia Commons; credit b: modification of work by "NASA/MODIS Rapid Response Team"/Wikimedia Commons)

Electrical Properties of Colloidal Particles

Dispersed colloidal particles are often electrically charged. A colloidal particle of iron (III) hydroxide, for example, does not contain enough hydroxide ions to compensate exactly for the positive charges on the iron (III) ions. Thus, each individual colloidal particle bears a positive charge, and the colloidal dispersion consists of charged colloidal particles and some free hydroxide ions, which keep the dispersion electrically neutral. Most metal hydroxide colloids have positive charges, whereas most metals and metal sulfides form negatively charged dispersions. All colloidal particles in any one system have charges of the same sign. This helps keep them dispersed because particles containing like charges repel each other.

We can take advantage of the charge on colloidal particles to remove them from a variety of mixtures. If we place a colloidal dispersion in a container with charged electrodes, positively charged particles, such as iron(III) hydroxide particles, would move to the negative electrode. There, the colloidal particles lose their charge and coagulate as a precipitate.The carbon and dust particles in smoke are often colloidally dispersed and electrically charged. Frederick Cottrell, an American chemist, developed a process to remove these particles.

Portrait of a Chemist: Frederick Gardner Cottrell



(a) (b)

Figure 8.36 (a) Frederick Cottrell developed (b) the electrostatic precipitator, a device designed to curb air pollution by removing colloidal particles from air. (credit b: modification of work by "SpLot"/Wikimedia Commons)

Born in Oakland, CA in 1877, Frederick Cottrell devoured textbooks as if they were novels and graduated from high school at the age of 16. He then entered the University of California (UC), Berkeley, completing a bachelor's degree in three years. He saved money from his $1200 annual salary as a chemistry teacher at Oakland High School to fund his studies in chemistry in Berlin with Nobel prize winner Jacobus Henricus van't Hoff, and in Leipzig with Wilhelm Ostwald, another Nobel awardee. After earning his PhD in physical chemistry, he returned to the United States to teach at UC Berkeley. He also consulted for the DuPont Company, where he developed the electrostatic precipitator, a device designed to curb air pollution by removing colloidal particles from air. Cottrell used the proceeds from his invention to fund a nonprofit research corporation to finance scientific research.

The charged particles are attracted to highly charged electrodes, where they are neutralized and deposited as dust (**Figure 8.37**). This is one of the important methods used to clean up the smoke from a variety of industrial processes. The process is also important in the recovery of valuable products from the smoke and flue dust of smelters, furnaces, and kilns. There are also ionic air filters designed for home use to improve indoor air quality.

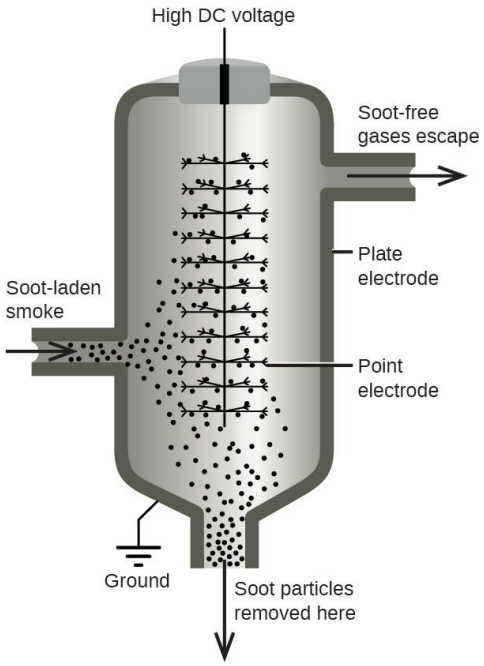


Figure 8.37 In a Cottrell precipitator, positively and negatively charged particles are attracted to highly charged electrodes, where they are neutralized and deposited as dust.

Gels

When we make gelatin, such as Jell-O, we are making a type of colloid (**Figure 8.38**). Gelatin sets on cooling because the hot aqueous mixture of gelatin coagulates as it cools and the whole mass, including the liquid, sets to an extremely viscous body known as a **gel**, a colloid in which the dispersing medium is a solid and the dispersed phase is a liquid. It appears that the fibers of the dispersing medium form a complex three-dimensional network, the interstices being filled with the liquid medium or a dilute solution of the dispersing medium. Because the formation of a gel is accompanied by the taking up of water or some other solvent, the gel is said to be hydrated or solvated.

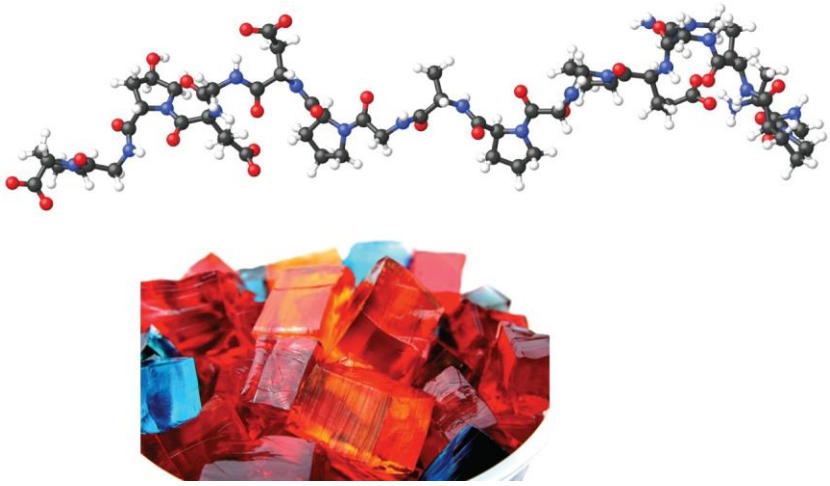


Figure 8.38 Gelatin desserts are colloids in which an aqueous solution of sweeteners and flavors is dispersed throughout a medium of solid proteins. (credit photo: modification of work by Steven Depolo)

Pectin, a carbohydrate from fruit juices, is a gel-forming substance important in jelly making. Silica gel, a colloidal dispersion of hydrated silicon dioxide, is formed when dilute hydrochloric acid is added to a dilute solution of sodium silicate. Canned Heat is a gel made by mixing alcohol and a saturated aqueous solution of calcium acetate.

Assessment 8.10:

A student assistant prepared two samples, a solution of sodium chloride in water and a colloid suspension in water that is very dilute. Before she could label the samples, she had to leave the building because of a fire alarm.

Being a conscientious individual, she came back as soon as possible to follow the protocol that all solutions must be labeled. To the naked eye, the solutions looked identical. How could she identify the two solutions?

Answer:

Strategy: The student assistant can apply her knowledge of the Tyndall Effect, wherein light is scattered by the colloid but not the solution. Using glass test tubes and a source of visible light, scattering would only be observed for the colloid!

Summary Assessment 8.11:

As promised in Assessment 8.1, we will analyze saline, used widely in clinical applications, which is prepared by dissolving sodium chloride in high purity water. When saline is made on an industrial scale, small portions of the solution are titrated to precipitate out the chloride ions as an insoluble salt, and the mass of the precipitate thus obtained is used to determine the concentration of sodium chloride.

In one such experiment, a 10.00 mL aliquot of saline was reacted with excess silver nitrate, and 0.214 g of silver chloride was obtained as a precipitate after considerable drying.

What is the concentration of sodium chloride in the saline solution?

Answer:

Strategy: First, we note that the precipitation is represented by the following net ionic equation, as seen earlier.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

It is convenient to use the molecular equation,

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

We note that 1 mole of precipitated AgCl corresponds to 1 mole of NaCl in the original saline solution, since the reactant $AgNO_3$ is in excess.

We apply the definition of molarity, given below.

Key equation:

$$M = \frac{mol\ solute}{L\ of\ solution}$$

Step 1: Calculate moles of AgCl, hence the moles of NaCl.

$$mol\ \text{AgCl} = 0.214 g\ \text{AgCl}\ \text{x}\ \frac{1\, mol\ \text{AgCl}}{143.321\, g\ \text{AgCl}} = 1.493\ \text{x}\ 10^{-3} mol$$

We keep an extra digit and will round to an appropriate number of sf in the next step.

Step 2: Calculate the molarity of NaCl. We note that the volume is 10.00 *mL*, or 0.01000 *L*.

$$M = \frac{mol\ solute}{L\ of\ solution} = \frac{1.493 \text{ x } 10^{-3} mol}{0.0100 L} = 0.1493\, M$$

We round this to 3 sf, based on the mass of the precipitate, 0.214 *g*. This is really close to the value anticipated, 0.150 *M*.

Key Terms

alloy solid mixture of a metallic element and one or more additional elements

amphiphilic molecules possessing both hydrophobic (nonpolar) and a hydrophilic (polar) parts

aqueous solution solution for which water is the solvent

boiling point elevation elevation of the boiling point of a liquid by addition of a solute

colligative property property of a solution that depends only on the concentration of a solute species

colloid (also, colloidal dispersion) mixture in which relatively large solid or liquid particles are dispersed uniformly throughout a gas, liquid, or solid

complete ionic equation chemical equation in which all dissolved ionic reactants and products, including spectator

concentrated qualitative term for a solution containing solute at a relatively high concentration

concentration quantitative measure of the relative amounts of solute and solvent present in a solution

crenation process whereby biological cells become shriveled due to loss of water by osmosis

dilute qualitative term for a solution containing solute at a relatively low concentration

dilution process of adding solvent to a solution in order to lower the concentration of solutes

dispersed phase substance present as relatively large solid or liquid particles in a colloid

dispersion medium solid, liquid, or gas in which colloidal particles are dispersed

dissociation physical process accompanying the dissolution of an ionic compound in which the compound's constituent ions are solvated and dispersed throughout the solution

dissolved describes the process by which solute components are dispersed in a solvent

electrolyte substance that produces ions when dissolved in water

emulsifying agent amphiphilic substance used to stabilize the particles of some emulsions

emulsion colloid formed from immiscible liquids

freezing point depression lowering of the freezing point of a liquid by addition of a solute

gel colloidal dispersion of a liquid in a solid

hemolysis rupture of red blood cells due to the accumulation of excess water by osmosis

Henry's law law stating the proportional relationship between the concentration of dissolved gas in a solution and the partial pressure of the gas in contact with the solution

hypertonic of greater osmotic pressure

hypotonic of less osmotic pressure

ideal solution solution that forms with no accompanying energy change

immiscible of negligible mutual solubility; typically refers to liquid substances

insoluble of relatively low solubility; dissolving only to a slight extent

ion pair solvated anion/cation pair held together by moderate electrostatic attraction

ion-dipole attraction electrostatic attraction between an ion and a polar molecule ion, are explicitly represented by formulas for their dissociated ions

isotonic of equal osmotic pressure

mass percentage ratio of solute-to-solution mass expressed as a percentage

mass-volume percent ratio of solute mass to solution volume, expressed as a percentage

miscible mutually soluble in all proportions; typically refers to liquid substances

molarity (*M*) unit of concentration, defined as the number of moles of solute dissolved in 1 liter of solution

molecular equation chemical equation in which all reactants and products are represented as neutral substances

net ionic equation chemical equation in which only those dissolved ionic reactants and products that undergo a chemical or physical change are represented (excludes spectator ions)

nonelectrolyte substance that does not produce ions when dissolved in water

osmosis diffusion of solvent molecules through a semipermeable membrane

osmotic pressure (Π) opposing pressure required to prevent bulk transfer of solvent molecules through a semipermeable membrane

partially miscible of moderate mutual solubility; typically refers to liquid substances

parts per billion (ppb) ratio of solute-to-solution mass multiplied by 10^9

parts per million (ppm) ratio of solute-to-solution mass multiplied by 10^6

percent composition percentage by mass of the various elements in a compound

precipitate insoluble product that forms from reaction of soluble reactants

precipitation reaction reaction that produces one or more insoluble products; when reactants are ionic compounds, sometimes called double-displacement or metathesis

Raoult's law the partial pressure exerted by a solution component is equal to the product of the component's mole fraction in the solution and its equilibrium vapor pressure in the pure state

saturated of concentration equal to solubility; containing the maximum concentration of solute possible for a given temperature and pressure

semipermeable membrane a membrane that selectively permits passage of certain ions or molecules

solubility extent to which a solute may be dissolved in water, or any solvent

solubility the extent to which a substance may be dissolved in water, or any solvent

soluble of relatively high solubility; dissolving to a relatively large extent

solute solution component present in a concentration less than that of the solvent

solvation exothermic process in which intermolecular attractive forces between the solute and solvent in a solution are established

solvent solution component present in a concentration that is higher relative to other components

spectator ion ion that does not undergo a chemical or physical change during a reaction, but its presence is required to maintain charge neutrality

spontaneous process physical or chemical change that occurs without the addition of energy from an external source

strong electrolyte substance that dissociates or ionizes completely when dissolved in water

supersaturated of concentration that exceeds solubility; a nonequilibrium state

Tyndall effect scattering of visible light by a colloidal dispersion

unsaturated of concentration less than solubility

volume percentage ratio of solute-to-solution volume expressed as a percentage

weak electrolyte substance that ionizes only partially when dissolved in water

Key Equations

- $\boldsymbol{M = \frac{mol\ solute}{L\ solution}}$
- $Percent\ by\ mass = \frac{mass\ of\ solute}{mass\ of\ solution} * 100$
- $ppm = \frac{mass\ solute}{mass\ solution} * 10^6 ppm$
- $ppb = \frac{mass\ solute}{mass\ solution} * 10^9 ppb$
- $N = MRT$

Summary

8.1 The Dissolution Process

A solution forms when two or more substances combine physically to yield a mixture that is homogeneous at the molecular level. The solvent is the most concentrated component and determines the physical state of the solution. The solutes are the other components typically present at concentrations less than that of the solvent. Solutions may form endothermically or exothermically, depending upon the relative magnitudes of solute and solvent intermolecular attractive forces. Ideal solutions form with no appreciable change in energy.

8.2 Electrolytes

Substances that dissolve in water to yield ions are called electrolytes. Electrolytes may be covalent compounds that chemically react with water to produce ions (for example, acids and bases), or they may be ionic compounds that dissociate to yield their constituent cations and anions, when dissolved. Dissolution of an ionic compound is facilitated by ion-dipole attractions between the ions of the compound and the polar water molecules. Soluble ionic substances and strong acids ionize completely and are strong electrolytes, while weak acids and bases ionize to only a small extent and are weak electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water.

8.3 Solubility

The extent to which one substance will dissolve in another is determined by several factors, including the types and relative strengths of intermolecular attractive forces that may exist between the substances' atoms, ions, or molecules. This tendency to dissolve is quantified as a substance's solubility, its maximum concentration in a solution at equilibrium under specified conditions. A saturated solution contains solute at a concentration equal to its solubility. A supersaturated solution is one in which a solute's concentration exceeds its solubility—a nonequilibrium (unstable) condition that will result in solute precipitation when the solution is appropriately perturbed. Miscible liquids are soluble in all proportions, and immiscible liquids exhibit very low mutual solubility. Solubilities for gaseous solutes decrease with increasing temperature, while those for most, but not all, solid solutes increase with temperature. The concentration of a gaseous solute in a solution is proportional to the partial pressure of the gas to which the solution is exposed, a relation known as Henry's law.

8.4 Molarity

Solutions are homogeneous mixtures. Many solutions contain one component, called the solvent, in which other components, called solutes, are dissolved. An aqueous solution is one for which the solvent is water. The concentration of a solution is a measure of the relative amount of solute in a given amount of solution. Concentrations may be measured using various units, with one very useful unit being molarity, defined as the number of moles of solute per liter of solution. The solute concentration of a solution may be decreased by adding solvent, a process referred to as dilution. The dilution equation is a simple relation between concentrations and volumes of a solution before and after dilution.

8.5 Other Units for Solution Concentrations

In addition to molarity, a number of other solution concentration units are used in various applications. Percentage concentrations based on the solution components' masses, volumes, or both are useful for expressing relatively high concentrations, whereas lower concentrations are conveniently expressed using ppm or ppb units. These units are popular in environmental, medical, and other fields where mole-based units such as molarity are not as commonly used.

8.6 Colligative Properties

Properties of a solution that depend only on the concentration of solute particles are called colligative properties. They include changes in the vapor pressure, boiling point, and freezing point of the solvent in the solution. The magnitudes of these properties depend only on the total concentration of solute particles in solution, not on the type of particles. The total concentration of solute particles in a solution also determines its osmotic pressure. This is the pressure that must be applied to the solution to prevent diffusion of molecules of pure solvent through a semipermeable membrane into the solution.

8.7 Colloids

Colloids are mixtures in which one or more substances are dispersed as relatively large solid particles or liquid droplets throughout a solid, liquid, or gaseous medium. The particles of a colloid remain dispersed and do not settle due to gravity, and they are often electrically charged. Colloids are widespread in nature and are involved in many technological applications.

Exercises

8.1 The Dissolution Process

1. How do solutions differ from compounds? From other mixtures?

2. Which of the principal characteristics of solutions can we see in the solutions of $K_2Cr_2O_7$ shown in **Figure 8.2**?

3. When KNO_3 is dissolved in water, the resulting solution is significantly colder than the water was originally.

(a) Is the dissolution of KNO_3 an endothermic or an exothermic process?

(b) What conclusions can you draw about the intermolecular attractions involved in the process?

(c) Is the resulting solution an ideal solution?

4. Give an example of each of the following types of solutions:

(a) a gas in a liquid

(b) a gas in a gas

(c) a solid in a solid

5. Indicate the most important types of intermolecular attractions in each of the following solutions:

(a) The solution in **Figure 8.2**.

(b) NO(*l*) in CO(*l*)

(c) Cl_2(*g*) in Br_2(*l*)

(d) HCl(*g*) in benzene C_6H_6(*l*)

(e) Methanol CH_3OH(*l*) in H_2O(*l*)

6. Predict whether each of the following substances would be more soluble in water (polar solvent) or in a hydrocarbon such as heptane (C_7H_{16}, nonpolar solvent):

(a) vegetable oil (nonpolar)

(b) isopropyl alcohol (polar)

(c) potassium bromide (ionic)

7. Heat is released when some solutions form; heat is absorbed when other solutions form. Provide a molecular explanation for the difference between these two types of spontaneous processes.

8. Solutions of hydrogen in palladium may be formed by exposing Pd metal to H_2 gas. The concentration of hydrogen in the palladium depends on the pressure of H_2 gas applied, but in a more complex fashion than can be described by Henry's law. Under certain conditions, 0.94 g of hydrogen gas is dissolved in 215 g of palladium metal (solution density = 10.8 g cm^3).

(a) Determine the molarity of this solution.

(b) Determine the molality of this solution.

(c) Determine the percent by mass of hydrogen atoms in this solution.

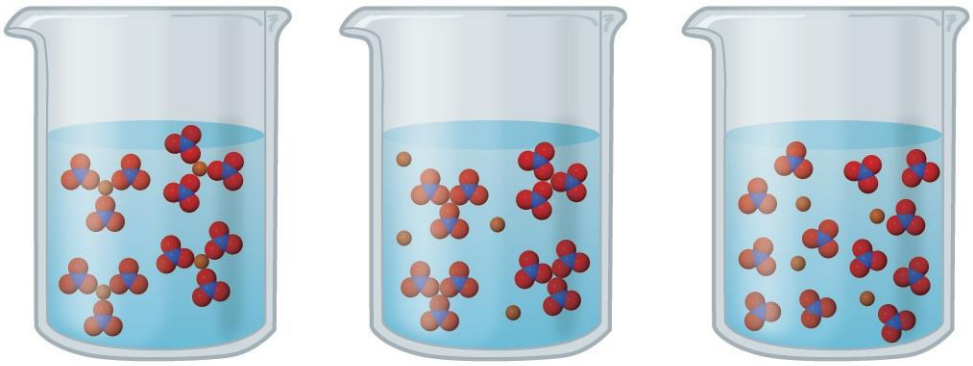
8.2 Electrolytes

9. Explain why the ions Na^+ and Cl^- are strongly solvated in water but not in hexane, a solvent composed of nonpolar molecules.

10. Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.

11. Consider the solutions presented:

(a) Which of the following sketches best represents the ions in a solution of $Fe(NO_3)_3(aq)$?



(x) (y) (z)

Write a balanced chemical equation showing the products of the dissolution of $Fe(NO_3)_3$.

12. Compare the processes that occur when methanol (CH_3OH), hydrogen chloride (HCl), and sodium hydroxide (NaOH) dissolve in water. Write equations and prepare sketches showing the form in which each of these compounds is present in its respective solution.

13. What is the expected electrical conductivity of the following solutions?

(a) $NaOH(aq)$

(b) $HCl(aq)$

(c) $C_6H_{12}O_6(aq)$ (glucose)

(d) $NH_3(aq)$

14. Why are most *solid* ionic compounds electrically nonconductive, whereas aqueous solutions of ionic compounds are good conductors? Would you expect a *liquid* (molten) ionic compound to be electrically conductive or nonconductive? Explain.

15. Indicate the most important type of intermolecular attraction responsible for solvation in each of the following solutions:

(a) the solutions in **Figure 8.8**

(b) methanol, CH_3OH, dissolved in ethanol, C_2H_5OH

(c) methane, CH_4, dissolved in benzene, C_6H_6

(d) the polar halocarbon CF_2Cl_2 dissolved in the polar halocarbon $CF_2ClCFCl_2$

(e) $O_2(l)$ in $N_2(l)$

8.3 Solubility

16. Suppose you are presented with a clear solution of sodium thiosulfate, $Na_2S_2O_3$. How could you determine whether the solution is unsaturated, saturated, or supersaturated?

17. Supersaturated solutions of most solids in water are prepared by cooling saturated solutions. Supersaturated solutions of most gases in water are prepared by heating saturated solutions. Explain the reasons for the difference in the two procedures.

2 From the balanced molecular equations, write the complete ionic and net ionic equations for the following:

(a) $K_2C_2O_4(aq) + Ba(OH)_2(aq) \longrightarrow 2KOH(aq) + BaC_2O_4(s)$

(b) $Pb(NO_3)_2(aq) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + 2HNO_3(aq)$

(c) $CaCO_3(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$

3 Write the molecular, total ionic, and net ionic equations for the following reactions:

(a) $Ca(OH)_2(aq) + HC_2H_3O_2(aq) \longrightarrow$

(b) $H_3PO_4(aq) + CaCl_2(aq) \longrightarrow$

18. Suggest an explanation for the observations that ethanol, C_2H_5OH, is completely miscible with water and that ethanethiol, C_2H_5SH, is soluble only to the extent of 1.5 g per 100 mL of water.

19. Calculate the percent by mass of KBr in a saturated solution of KBr in water at 10 °C. See **Figure 11.17** for useful data, and report the computed percentage to one significant digit.

20. Which of the following gases is expected to be most soluble in water? Explain your reasoning.

(a) CH_4

(b) CCl_4

(c) $CHCl_3$

21. At 0 °C and 1.00 atm, as much as 0.70 g of O_2 can dissolve in 1 L of water. At 0 °C and 4.00 atm, how many grams of O_2 dissolve in 1 L of water?

22. Refer to **Figure 8.11**.

(a) How did the concentration of dissolved CO_2 in the beverage change when the bottle was opened?

(b) What caused this change?

(c) Is the beverage unsaturated, saturated, or supersaturated with CO_2?

23. The Henry's law constant for CO_2 is 3.4×10^{-2} *M*/atm at 25 °C. What pressure of carbon dioxide is needed to maintain a CO_2 concentration of 0.10 *M* in a can of lemon-lime soda?

24. The Henry's law constant for O_2 is 1.3×10^{-3} *M*/atm at 25 °C. What mass of oxygen would be dissolved in a 40-L aquarium at 25 °C, assuming an atmospheric pressure of 1.00 atm, and that the partial pressure of O_2 is 0.21 atm?

25. How many liters of HCl gas, measured at 30.0 °C and 745 torr, are required to prepare 1.25 L of a 3.20-*M* solution of hydrochloric acid?

8.4 Molarity

7. Explain what changes and what stays the same when 1.00 L of a solution of NaCl is diluted to 1.80 L.

8. What information do we need to calculate the molarity of a sulfuric acid solution?

9. What does it mean when we say that a 200-mL sample and a 400-mL sample of a solution of salt have the same molarity? In what ways are the two samples identical? In what ways are these two samples different?

10. Determine the molarity for each of the following solutions:

(a) 0.444 mol of $CoCl_2$ in 0.654 L of solution

(b) 98.0 g of phosphoric acid, H_3PO_4, in 1.00 L of solution

(c) 0.2074 g of calcium hydroxide, $Ca(OH)_2$, in 40.00 mL of solution

(d) 10.5 kg of $Na_2SO_4 \cdot 10H_2O$ in 18.60 L of solution

(e) 7.0×10^{-3} mol of I_2 in 100.0 mL of solution

(f) 1.8×10^4 mg of HCl in 0.075 L of solution

11. Determine the molarity of each of the following solutions:

(a) 1.457 mol KCl in 1.500 L of solution

(b) 0.515 g of H_2SO_4 in 1.00 L of solution

(c) 20.54 g of $Al(NO_3)_3$ in 1575 mL of solution

(d) 2.76 kg of $CuSO_4 \cdot 5H_2O$ in 1.45 L of solution

(e) 0.005653 mol of Br_2 in 10.00 mL of solution

(f) 0.000889 g of glycine, $C_2H_5NO_2$, in 1.05 mL of solution

12. Consider this question: What is the mass of the solute in 0.500 L of 0.30 *M* glucose, $C_6H_{12}O_6$, used for intravenous injection?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

13. Consider this question: What is the mass of solute in 200.0 L of a 1.556-*M* solution of KBr?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

14. Calculate the number of moles and the mass of the solute in each of the following solutions:

(a) 2.00 L of 18.5 *M* H_2SO_4, concentrated sulfuric acid

(b) 100.0 mL of 3.8×10^{-5} *M* NaCN, the minimum lethal concentration of sodium cyanide in blood serum

(c) 5.50 L of 13.3 *M* H_2CO, the formaldehyde used to "fix" tissue samples

(d) 325 mL of 1.8×10^{-6} *M* $FeSO_4$, the minimum concentration of iron sulfate detectable by taste in drinking water

15. Calculate the number of moles and the mass of the solute in each of the following solutions:

(a) 325 mL of 8.23×10^{-5} *M* KI, a source of iodine in the diet

(b) 75.0 mL of 2.2×10^{-5} *M* H_2SO_4, a sample of acid rain

(c) 0.2500 L of 0.1135 *M* K_2CrO_4, an analytical reagent used in iron assays

(d) 10.5 L of 3.716 *M* $(NH_4)_2SO_4$, a liquid fertilizer

16. Consider this question: What is the molarity of $KMnO_4$ in a solution of 0.0908 g of $KMnO_4$ in 0.500 L of solution?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

17. Consider this question: What is the molarity of HCl if 35.23 mL of a solution of HCl contain 0.3366 g of

HCl?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

18. Calculate the molarity of each of the following solutions:

(a) 0.195 g of cholesterol, $C_{27}H_{46}O$, in 0.100 L of serum, the average concentration of cholesterol in human serum

(b) 4.25 g of NH_3 in 0.500 L of solution, the concentration of NH_3 in household ammonia

(c) 1.49 kg of isopropyl alcohol, C_3H_7OH, in 2.50 L of solution, the concentration of isopropyl alcohol in rubbing alcohol

(d) 0.029 g of I_2 in 0.100 L of solution, the solubility of I_2 in water at 20 °C

19. Calculate the molarity of each of the following solutions:

(a) 293 g HCl in 666 mL of solution, a concentrated HCl solution

(b) 2.026 g $FeCl_3$ in 0.1250 L of a solution used as an unknown in general chemistry laboratories

(c) 0.001 mg Cd^{2+} in 0.100 L, the maximum permissible concentration of cadmium in drinking water

(d) 0.0079 g $C_7H_5SNO_3$ in one ounce (29.6 mL), the concentration of saccharin in a diet soft drink.

20. There is about 1.0 g of calcium, as Ca^{2+}, in 1.0 L of milk. What is the molarity of Ca^{2+} in milk?

21. What volume of a 1.00-*M* $Fe(NO_3)_3$ solution can be diluted to prepare 1.00 L of a solution with a concentration of 0.250 *M*?

22. If 0.1718 L of a 0.3556-*M* C_3H_7OH solution is diluted to a concentration of 0.1222 *M*, what is the volume of the resulting solution?

23. If 4.12 L of a 0.850 *M*-H_3PO_4 solution is be diluted to a volume of 10.00 L, what is the concentration of the resulting solution?

24. What volume of a 0.33-*M* $C_{12}H_{22}O_{11}$ solution can be diluted to prepare 25 mL of a solution with a concentration of 0.025 *M*?

25. What is the concentration of the NaCl solution that results when 0.150 L of a 0.556-*M* solution is allowed to evaporate until the volume is reduced to 0.105 L?

26. What is the molarity of the diluted solution when each of the following solutions is diluted to the given final volume?

(a) 1.00 L of a 0.250-*M* solution of $Fe(NO_3)_3$ is diluted to a final volume of 2.00 L

(b) 0.5000 L of a 0.1222-*M* solution of C_3H_7OH is diluted to a final volume of 1.250 L

(c) 2.35 L of a 0.350-*M* solution of H_3PO_4 is diluted to a final volume of 4.00 L

(d) 22.50 mL of a 0.025-*M* solution of $C_{12}H_{22}O_{11}$ is diluted to 100.0 mL

27. What is the final concentration of the solution produced when 225.5 mL of a 0.09988-*M* solution of Na_2CO_3 is allowed to evaporate until the solution volume is reduced to 45.00 mL?

28. A 2.00-L bottle of a solution of concentrated HCl was purchased for the general chemistry laboratory. The solution contained 868.8 g of HCl. What is the molarity of the solution?

29. An experiment in a general chemistry laboratory calls for a 2.00-*M* solution of HCl. How many mL of

11.9 *M*

HCl would be required to make 250 mL of 2.00 *M* HCl?

30. What volume of a 0.20-*M* K_2SO_4 solution contains 57 g of K_2SO_4?

31. The US Environmental Protection Agency (EPA) places limits on the quantities of toxic substances that may be discharged into the sewer system. Limits have been established for a variety of substances, including hexavalent chromium, which is limited to 0.50 mg/L. If an industry is discharging hexavalent chromium as potassium dichromate ($K_2Cr_2O_7$), what is the maximum permissible molarity of that substance?

8.5 Other Units for Solution Concentrations

32. Consider this question: What mass of a concentrated solution of nitric acid (68.0% HNO_3 by mass) is needed to prepare 400.0 g of a 10.0% solution of HNO_3 by mass?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

33. What mass of a 4.00% NaOH solution by mass contains 15.0 g of NaOH?

34. What mass of solid NaOH (97.0% NaOH by mass) is required to prepare 1.00 L of a 10.0% solution of NaOH by mass? The density of the 10.0% solution is 1.109 g/mL.

35. What mass of HCl is contained in 45.0 mL of an aqueous HCl solution that has a density of 1.19 g cm^{-3} and contains 37.21% HCl by mass?

36. The hardness of water (hardness count) is usually expressed in parts per million (by mass) of $CaCO_3$, which is equivalent to milligrams of $CaCO_3$ per liter of water. What is the molar concentration of Ca^{2+} ions in a water sample with a hardness count of 175 mg $CaCO_3$/L?

37. The level of mercury in a stream was suspected to be above the minimum considered safe (1 part per billion by weight). An analysis indicated that the concentration was 0.68 parts per billion. Assume a density of 1.0 g/mL and calculate the molarity of mercury in the stream.

38. In Canada and the United Kingdom, devices that measure blood glucose levels provide a reading in millimoles per liter. If a measurement of 5.3 m*M* is observed, what is the concentration of glucose ($C_6H_{12}O_6$) in mg/dL?

39. A throat spray is 1.40% by mass phenol, C_6H_5OH, in water. If the solution has a density of 0.9956 g/mL, calculate the molarity of the solution.

40. Copper(I) iodide (CuI) is often added to table salt as a dietary source of iodine. How many moles of CuI are contained in 1.00 lb (454 g) of table salt containing 0.0100% CuI by mass?

41. A cough syrup contains 5.0% ethyl alcohol, C_2H_5OH, by mass. If the density of the solution is 0.9928 g/mL, determine the molarity of the alcohol in the cough syrup.

42. D5W is a solution used as an intravenous fluid. It is a 5.0% by mass solution of dextrose ($C_6H_{12}O_6$) in water. If the density of D5W is 1.029 g/mL, calculate the molarity of dextrose in the solution.

43. Find the molarity of a 40.0% by mass aqueous solution of sulfuric acid, H_2SO_4, for which the density is 1.3057 g/mL.

8.6 Colligative Properties

26. Which is/are part of the macroscopic domain of solutions and which is/are part of the microscopic domain: boiling point elevation, Henry's law, hydrogen bond, ion-dipole attraction, molarity, nonelectrolyte, nonstoichiometric compound, osmosis, solvated ion?

8.7 Colloids

27. Identify the dispersed phase and the dispersion medium in each of the following colloidal systems: starch dispersion, smoke, fog, pearl, whipped cream, floating soap, jelly, milk, and ruby.

28. Distinguish between dispersion methods and condensation methods for preparing colloidal systems.

29. How do colloids differ from solutions with regard to dispersed particle size and homogeneity?

30. Explain the cleansing action of soap.

31. How can it be demonstrated that colloidal particles are electrically charged?