

CHAPTER 14: SOLUTIONS

Enhanced Introductory College Chemistry

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Please visit the web version of Enhanced Introductory College Chemistry (<https://ecampusontario.pressbooks.pub/enhancedchemistry/>) to access the complete book, interactive activities and ancillary resources.

In this chapter, you will learn about

- What a solution is, how they form, their molecular properties, and how they can either absorb or produce heat

- The effects that temperature and pressure have on solubility
- Solution concentration calculations using the molarity equation, dilution calculations using the dilution equation, and the concentration units of mass percentage, volume percentage, mass-volume percentage, parts-per-million (ppm), and parts-per-billion (ppb)
- The effect of solute concentration on various solution properties such as vapour pressure, boiling point, freezing point, and osmotic pressure
- The process of distillation and its practical applications
- The process of osmosis and how it is applied industrially and in nature
- The composition and properties of colloidal dispersions, and applications of colloids

To better support your learning, you should be familiar with the following concepts before starting this chapter:

- Mathematical measurements, uncertainty, accuracy, and precision
- The periodic table, molecular formulas, and the difference between elements and compounds
- Nomenclature of molecular and ionic compounds, and acids
- The concept of the mole, molecular mass, and Avogadro's number
- Writing and balancing chemical equations
- Stoichiometry



Figure 14a: 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 Jim Maragos/USFWS – Pacific Region, CC BY-NC 2.0. / Cropped.)

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.

Attribution & References

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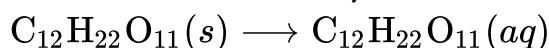
14.1 SOLUTIONS: AN INTRODUCTION

Learning Objectives

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

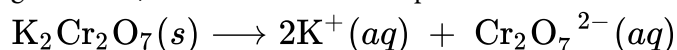
- Describe the basic properties of solutions and how they form
- Predict whether a given mixture will yield a solution based on molecular properties of its components
- Explain why some solutions either produce or absorb heat when they form

An earlier chapter of this text introduced **solutions**, 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:



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 colourless 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 14.1a), as indicated in this equation:



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 14.1a: When potassium dichromate ($K_2Cr_2O_7$) is mixed with water, it forms a homogeneous orange solution. (credit: modification of work by Mark Ott in *Chemistry (OpenStax)*, CC BY 4.0).

Exercise 14.1a

Practice using the following PhET simulation: Sugar and Salt Solutions

(<https://phet.colorado.edu/sims/cheerpj/sugar-and-salt-solutions/latest/sugar-and-salt-solutions.html?simulation=sugar-and-salt-solutions>)

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 14.1a gives examples of several different solutions and the phases of the solutes and solvents.

Table 14.1a: Different Types of Solutions

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	$Zn(s)$	$Cu(s)$

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 14.1a.
- 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. The topic of spontaneity is critically important to the study of chemical thermodynamics and is treated more thoroughly in a later chapter of this text. For purposes of this chapter's discussion, it will suffice to consider two criteria that *favour*, but do not guarantee, the spontaneous formation of a solution:

1. a decrease in the internal energy of the system (an exothermic change, as discussed in the previous chapter on thermochemistry)
2. an increase in the disorder in the system (which indicates an increase in the *entropy* of the system, as you will learn about in the later chapter on thermodynamics)

In the process of dissolution, an internal energy change often, but not always, occurs as heat is absorbed or evolved. An increase in disorder always results when a solution forms.

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 behaviour) 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 14.1b). 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.

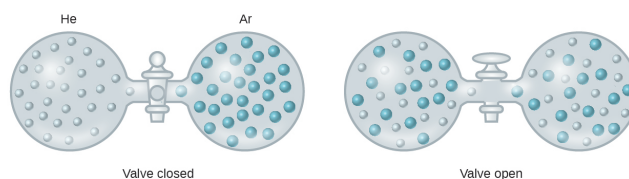


Figure 14.1b: Samples of helium and argon spontaneously mix to give a solution in which the disorder of the atoms of the two gases is increased (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Ideal solutions may also form when structurally similar liquids are mixed. For example, mixtures of the alcohols methanol (CH_3OH) and ethanol ($\text{C}_2\text{H}_5\text{OH}$) 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 14.1b 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 14.1c, 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.

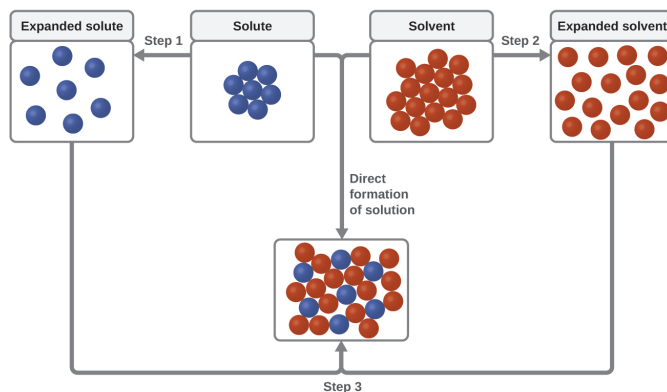


Figure 14.1c: This schematic representation of dissolution shows a stepwise process involving the endothermic separation of solute and solvent species (Steps 1 and 2) and exothermic solvation (Step 3) (credit: *Chemistry (OpenStax)*, CC BY 4.0).

For example, cooking oils and water will not mix to any appreciable extent to yield solutions (Figure 14.1d). 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.

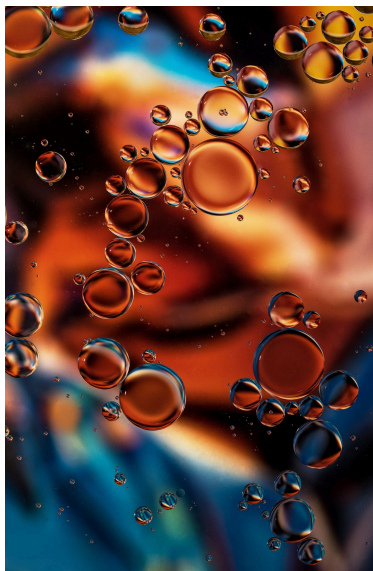


Figure 14.1d: A mixture of nonpolar cooking oil and polar water does not yield a solution. (credit: work by David Clode, Unsplash license).

On the other hand, a mixture of ethanol and water will mix in any proportion 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 favoured, 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 (Figure 14.1e) for treating injuries. 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 14.1e: An instant cold pack gets cold when certain salts, such as ammonium nitrate, dissolve in water—an endothermic process (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Watch Comparing Temperature Change from Heat of Solution (KCl vs NaOH) (2 mins)
(<https://www.youtube.com/watch?v=o7adWQqvDUU>)

Indigenous Perspective: Maple Syrup

For thousands of years maple syrup has been gathered from maple trees in Canada and the northern United States, a practice that has sustained Indigenous communities while also helping to connect them to the land. This is a practice carried out by the Anishinaabeg, Haudenosaunee and other First Nations. The physical and chemical properties of this unique solution give it the ability to act as not only a sweetener, but also as an anesthetic and a desiccant to help cure meats. Aside from containing many nutrients and carbohydrates, maple syrup was also used as an essential trade item. To learn more about the chemistry of maple syrup and its importance in Indigenous culture, please visit:



Figure 14.1f: Sap collection (credit: work by Oven Fresh, PD)

- The Indigenous Origins of Maple Syrup [New Tab] (<https://www.americanindianmagazine.org/Indigenous-origins-of-maple-syrup>)
- Food for Thought: The rich Indigenous history of maple syrup [New Tab] (<https://the-peak.ca/2021/06/food-for-thought-the-rich-indigenous-history-of-maple-syrup/>)

Attribution & References

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Notes

1. If bubbles of gas are observed within the liquid, the mixture is not homogeneous and, thus, not a solution.

14.2 SOLUBILITY

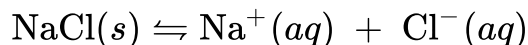
Learning Objectives

By the end of this section, 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 module 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:



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.

Exercise 14.2a

Practice using the following PhET simulation: Salts & Solubility (<https://phet.colorado.edu/sims/cheerpi/soluble-salts/latest/soluble-salts.html?simulation=soluble-salts>)

Solutions in which a solute concentration *exceeds* its solubility may be prepared. 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 CO₂ concentration will decrease until it reaches its equilibrium value.

Watch Crystal Growing – Cool Science Experiment (3 mins) (<https://www.youtube.com/watch?v=D1PDE5OawuI>)

Watch crystallization of sodium acetate (v2) (2 min) (https://www.youtube.com/watch?v=M_DYQ4C9fVc)

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, CHCl₃. Considering the role of the solvent’s chemical structure, note that the solubility of oxygen in the liquid hydrocarbon hexane, C₆H₁₄, 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 14.2a). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.

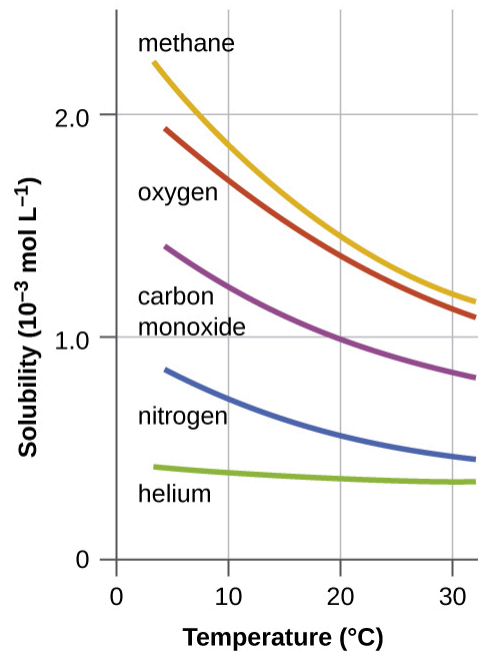


Figure 14.2a: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions (credit: *Chemistry (OpenStax)*, CC BY 4.0).

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 14.2b).

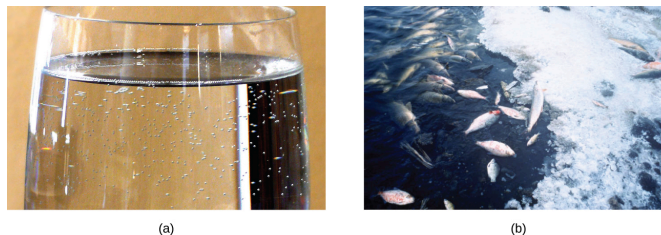


Figure 14.2b: Solubility of oxygen in water. (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (credit a: modification of work by Liz West, CC BY 2.0; credit b: modification of work by U.S. Fish and Wildlife Service, PD)

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

CO₂ 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 14.2c). 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.”

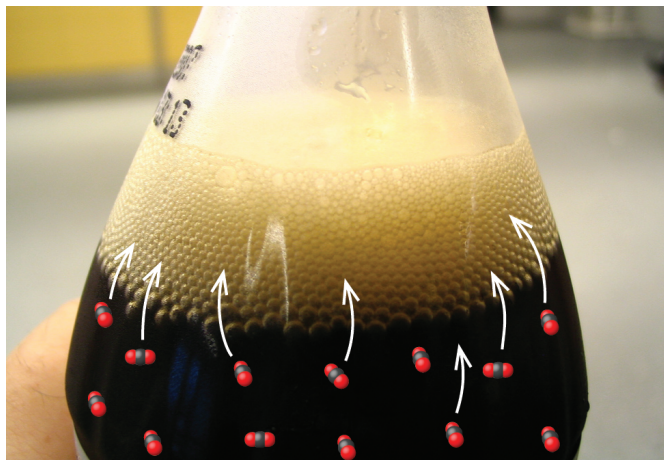


Figure 14.2c: Opening the bottle of carbonated beverage reduces the pressure of the *gaseous* carbon dioxide above the beverage. The solubility of CO₂ is thus lowered, and some *dissolved* carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Chiara Coetzee, CCO; in *Chemistry (OpenStax)*, CC BY 4.0).

For many gaseous solutes, the relationship 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 14.2a

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 $1.38 \times 10^{-3} \text{ mol L}^{-1}$. 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, C_g , of a gas ($1.38 \times 10^{-3} \text{ mol L}^{-1}$, in this case) is directly proportional to the pressure, P_g , of the undissolved gas above the solution (101.3 kPa, or 760 torr, in this case). Because we know both C_g and P_g , we can rearrange this expression to solve for k .

$$\begin{aligned}
 C_g &= kP_g \\
 k &= \frac{C_g}{P_g} \\
 &= \frac{1.38 \times 10^{-3} \text{ mol L}^{-1}}{101.3 \text{ kPa}} \\
 &= 1.36 \times 10^{-5} \text{ mol L}^{-1} \text{ kPa}^{-1} \\
 &= (1.82 \times 10^{-6} \text{ mol L}^{-1} \text{ torr}^{-1})
 \end{aligned}$$

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

$$\begin{aligned}
 C_g &= kP_g \\
 &= 1.36 \times 10^{-5} \text{ mol L}^{-1} \text{ kPa}^{-1} \times 20.7 \text{ kPa} \\
 &\text{(or } 1.82 \times 10^{-6} \text{ mol L}^{-1} \text{ torr}^{-1} \times 155 \text{ torr)} \\
 &= 2.82 \times 10^{-4} \text{ mol L}^{-1}
 \end{aligned}$$

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.

Exercise 14.2b

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).

Check Your Answer¹

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 become 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 decompression 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 14.2d).

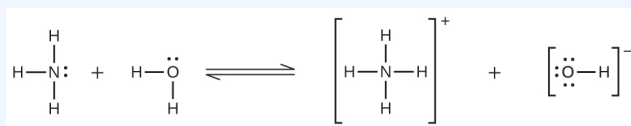


Figure 14.2d: Reaction of ammonia with water to form ammonium ion and hydroxide ion (credit: Chemistry (OpenStax), CC BY 4.0).

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.

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 14.2e) are examples of liquids that are completely miscible with water. Two-cycle motor oil is miscible with gasoline.



Figure 14.2e: Water and antifreeze are miscible; mixtures of the two are homogeneous in all proportions. (credit: work by dno1967, CC BY 2.0)

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 14.2f), 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 14.2f: 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: work by Yortw, CC BY 2.0)

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 discoloured 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 colour of the bromine layer (Figure 14.2g).



Figure 14.2g: 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 in *Chemistry (OpenStax)*, CC BY 4.0).

Solutions of Solids in Liquids

The dependence of solubility on temperature for a number of inorganic solids in water is shown by the solubility curves in Figure 14.2h. Reviewing these data indicate a general trend of increasing solubility with temperature, although there are exceptions, as illustrated by the ionic compound cerium sulfate.

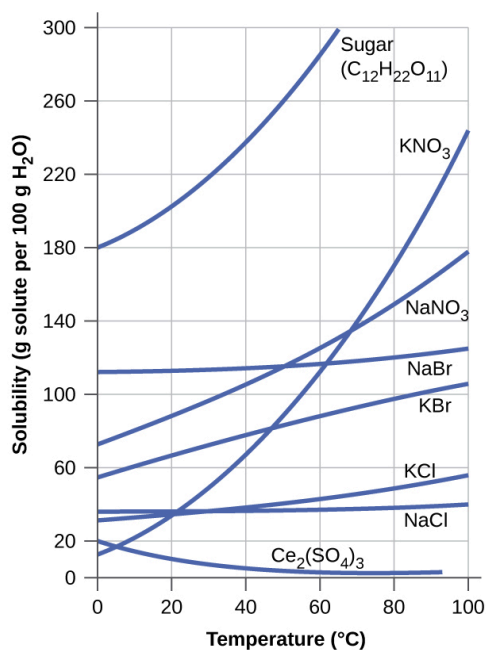


Figure 14.2h: This graph shows how the solubility of several solids changes with temperature (credit: *Chemistry (OpenStax)*, CC BY 4.0).

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 take advantage of this behaviour.

Watch Crystallization of the “Magic” Gel Hand Warmer bag (HD) (1 min)

(<https://www.youtube.com/watch?v=8J1igZvEK3c>)

Solubility of Ionic Compounds

The solubility of ionic compounds can be predicted based on the nature of the cations and anions present in the compound. Compounds that are soluble (along with their exceptions) can be seen in Table 14.2a, while those that are insoluble (along with their exceptions) can be seen in Table 14.2b.

Table 14.2a: Soluble Ionic Compounds

Compounds containing these ions	Exceptions
NH_4^+	none
Li^+	none
Na^+	none
K^+	none
Rb^+	none
Cs^+	none
Cl^-	Compounds with Ag^+ , Hg_2^{2+} , Pb^{2+}
Br^-	Compounds with Ag^+ , Hg_2^{2+} , Pb^{2+}
I^-	Compounds with Ag^+ , Hg_2^{2+} , Pb^{2+}
F^-	Compounds with group 2 metal cations Pb^{2+} , Fe^{2+}
$\text{C}_2\text{H}_3\text{O}_2^-$	None
HCO_3^-	None
ClO_3^-	None
SO_4^{2-}	Compounds with Ag^+ , Ba^{2+} , Ca^{2+} , Hg_2^{2+} , Pb^{2+} , Sr^{2+}

Source: “Table 14.2a” by Gregory Anderson is adapted from “4.2 Classifying Chemical Reactions” in *Chemistry (OpenStax)*, CC BY 4.0

Table 14.2b: Insoluble Ionic Compounds

Compounds containing these ions	Exceptions
CO_3^{2-}	Compounds with group 1 cations and NH_4^+
CrO_4^{2-}	Compounds with group 1 cations and NH_4^+
PO_4^{3-}	Compounds with group 1 cations and NH_4^+
S^{2-}	Compounds with group 1 cations and NH_4^+
OH^-	Compounds with group 1 cations and Ba^{2+}

Source: “Table 14.2b” by Gregory Anderson is adapted from “4.2 Classifying Chemical Reactions” in *Chemistry (OpenStax)*, CC BY 4.0

Links to Interactive Learning Tools

Explore Precipitation Reactions (<https://www.physicsclassroom.com/Concept-Builders/Chemistry/Precipitation-Reactions>) from the Physics Classroom (<https://www.physicsclassroom.com>).

Practice Solution Word Definitions (<https://h5pstudio.ecampusontario.ca/content/19791>) from eCampusOntario H5P Studio (<https://h5pstudio.ecampusontario.ca/>).

Key Equations

- $C_g = kP_g$

Attribution & References

Except where otherwise noted, this page is adapted by Gregory A. Anderson from “11.3 Solubility (<https://boisestate.pressbooks.pub/chemistry/chapter/11-3-solubility/>)” In *General Chemistry 1 & 2* by Rice University, a derivative of *Chemistry (Open Stax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson and is licensed under CC BY 4.0. Access for free at *Chemistry (OpenStax)* (<https://openstax.org/books/chemistry/pages/1-introduction>).

Notes

1. 7.25×10^{-3} in 100.0 mL or 0.0725 g/L

14.3 MOLARITY

Learning Objectives

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

- Describe the fundamental properties of solutions
- Calculate solution concentrations using molarity
- Perform dilution calculations using the dilution equation

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 14.3a). In this section, we will describe one of the most common ways in which the relative compositions of mixtures may be quantified.



Figure 14.3a: 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: Photo by Jane Whitney in *Chemistry (OpenStax)*, CC BY 4.0).

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 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 litre (1 L) of the solution:

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

Example 14.3a

Calculating Molarity

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:

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

Exercise 14.3a

Check Your Learning Exercise (Text Version)

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?

- a. 0.05 M
- b. 0.03 M
- c. 0.07 M
- d. 0.04 M

Check Your Answer¹

Source: “Exercise 14.3a” is adapted from “Example 6.3-1” in General Chemistry 1 & 2, a derivative of Chemistry (Open Stax) by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson, licensed under CC BY 4.0.

Example 14.3b

Deriving Moles and Volumes from Molarity

How much sugar (mol) is contained in a modest sip (~10 mL) of the soft drink from Example 14.3a?

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 14.3a, 0.375 *M*:

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

$$\text{mol solute} = M \times \text{L solution}$$

$$\text{mol solute} = 0.375 \frac{\text{mol sugar}}{\text{L}} \times (10 \text{ mL} \times \frac{1\text{L}}{1000 \text{ mL}}) = 0.004 \text{ mol sugar}$$

Exercise 14.3b

Check Your Learning Exercise (Text Version)

What volume (mL) of the sweetened tea (0.05*M*) described in Exercise 14.3a contains the same amount of sugar as 10 mL of the soft drink described in Example 14.3b?

- 80 mL
- 75 mL
- 70 mL
- 85 mL

Check Your Answer²

Source: “Exercise 14.3b” is adapted from “Example 6.3-2” in General Chemistry 1 & 2, a derivative of Chemistry (Open Stax) by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson, licensed under CC BY 4.0.

Example 14.3c

Calculating Molarity from the Mass of Solute

Distilled white vinegar (Figure 14.3b) is a solution of acetic acid, $\text{CH}_3\text{CO}_2\text{H}$, 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 14.3b: Distilled white vinegar is a solution of acetic acid in water. (credit: modification of work by HomeSpot HQ, CC BY 2.0)

Solution

As in previous textbox shaded, 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:

$$M = \frac{\text{mol solute}}{\text{L solution}} = \frac{25.2 \text{ g CH}_3\text{CO}_2\text{H} \times \frac{1 \text{ mol CH}_3\text{CO}_2\text{H}}{60.052 \text{ g CH}_3\text{CO}_2\text{H}}}{0.500 \text{ L solution}} = 0.839 M$$

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

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

Exercise 14.3c

Calculate the molarity of 6.52 g of CoCl_2 (128.9 g/mol) dissolved in an aqueous solution with a total volume of 75.0 mL.

Check Your Answer³

Example 14.3d

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 14.2b:

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

$$\text{mol solute} = M \times \text{L solution}$$

$$\text{mol solute} = 5.30 \frac{\text{mol NaCl}}{\text{L}} \times 0.250 \text{ L} = 1.325 \text{ mol NaCl}$$

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

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

Exercise 14.3d

How many grams of CaCl_2 (110.98 g/mol) are contained in 250.0 mL of a 0.200-*M* solution of calcium chloride?

Check Your Answer⁴

When performing stepwise calculations, as in Example 4, it is important to refrain from rounding any intermediate calculation results, which can lead to rounding errors in the final result. In Example 4, 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 14.3e). This eliminates intermediate steps so that only the final result is rounded.

Example 14.3e

Determining the Volume of Solution Containing a Given Mass of Solute

In Example 14.3c, 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:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g solute}} = \text{mol solute}$$

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

$$\text{mol solute} \times \frac{\text{L solution}}{\text{mol solute}} = \text{L solution}$$

Combining these two steps into one yields:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g solute}} \times \frac{\text{L solution}}{\text{mol solute}} = \text{L solution} \quad 75.6 \text{ g CH}_3\text{CO}_2\text{H} \left(\frac{\text{mol CH}_3\text{CO}_2\text{H}}{60.05 \text{ g}} \right) \left(\frac{\text{L solution}}{0.839 \text{ mol CH}_3\text{CO}_2\text{H}} \right) = 1.50 \text{ L solution}$$

Exercise 14.3e

Check Your Learning Exercise (Text Version)

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

- a. 0.370 L
- b. 0.250 L
- c. 0.983 L
- d. 0.137 L

Check Your Answer⁵

Source: “Exercise 14.3e” is adapted from “Example 6.3-5” in General Chemistry 1 & 2, a derivative of Chemistry (Open Stax) by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson, licensed under CC BY 4.0.

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 14.3c).

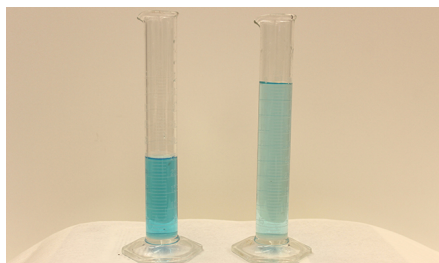


Figure 14.3c: Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott in *Chemistry (OpenStax)*, CC BY 4.0).

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 14.3d).



Figure 14.3d: 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 in *Chemistry (OpenStax)*, CC BY 4.0).

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 litres:

$$n = ML$$

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 litres 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.

Exercise 14.3f

Practice using the following PhET simulation: Concentration. (https://phet.colorado.edu/sims/html/concentration/latest/concentration_en.html)

Exercise 14.3g

Practice using the following PhET simulation: Molarity

Example 14.3f

Determining the Concentration of a Diluted Solution

If 0.850 L of a 5.00-*M* solution of copper nitrate, $\text{Cu}(\text{NO}_3)_2$, 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 :

$$C_1 V_1 = C_2 V_2$$
$$C_2 = \frac{C_1 V_1}{V_2}$$

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:

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

This result compares well to our ballpark estimate (it's a bit less than one-half the stock concentration, 5 M).

Exercise 14.3h

Check Your Learning Exercise (Text Version)

What is the concentration of the solution that results from diluting 25.0 mL of a 2.04-M solution of CH_3OH to 500.0 mL?

- 0.102 M CH_3OH
- 0.202 M CH_3OH
- 0.205 M CH_3OH
- 0.703 M CH_3OH

Check Your Answer⁶

Source: "Exercise 14.3h" is adapted from "Example 6.3-6" in General Chemistry 1 & 2, a derivative of Chemistry (Open Stax) by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson, licensed under CC BY 4.0.

Example 14.3g

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 :

$$C_1 V_1 = C_2 V_2$$

$$V_2 = \frac{C_1 V_1}{C_2}$$

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:

$$V_2 = \frac{(0.45 \text{ M})(0.011 \text{ L})}{0.12 \text{ M}}$$

$$V_2 = 0.041 \text{ L}$$

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.

Exercise 14.3i

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 ?

Check Your Answer⁷

Example 14.3h

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, C_1 , and the volume and concentration of the

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

$$C_1 V_1 = C_2 V_2$$

$$V_1 = \frac{C_2 V_2}{C_1}$$

Since the concentration of the diluted solution 0.100 M is roughly one-sixteenth that of the stock solution (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 litres. Substituting the given values and solving for the unknown volume yields:

$$V_1 = \frac{(0.100\text{ M})(5.00\text{ L})}{1.59\text{ M}}$$

$$V_1 = 0.314\text{ L}$$

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.

Exercise 14.3j

What volume of a 0.575-M solution of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, can be prepared from 50.00 mL of a 3.00-M glucose solution?

Check Your Answer⁸

Visualizing Different Solutes in Solution

When seeking to understand compounds in solution, we must differentiate between the behaviour of covalent compounds and ionic compounds. Assuming the solid dissolves, ionic compounds will dissociate into their ions. In contrast, for covalent compounds, the molecules of the compound remain intact.

So far in this section, we have been examining solutions of sugar ($\text{C}_6\text{H}_{12}\text{O}_6$). If we were to be able to see on the level of molecules, we might see a solution like that below (Figure 14.3e), where a triangle represents a molecule of glucose. While the molecules move around in the solution once they are dissolved in water, the atoms in the molecule do not break apart. The covalent bonds are intact.

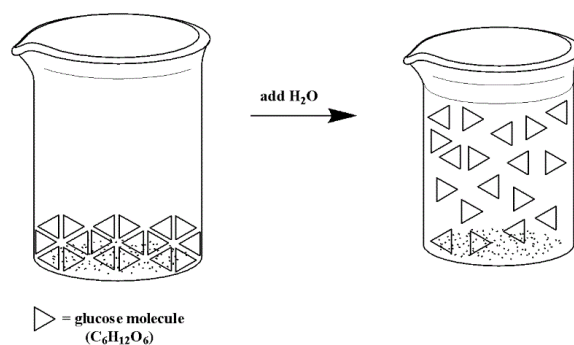


Figure 14.3e: Glucose, when placed into a solvent such as water, does not dissociate. The molecules maintain their covalent bonds, and thus float around in solution as complete glucose molecules (triangles above) (*General Chemistry I & II*, CC BY 4.0)

In contrast, when an ionic compound dissolves in solution, the cations and anions are separated (they dissociate) and each ion is surrounded by water molecules. We can depict the dissociation process for a compound like NaCl using the image below (Figure 14.3f):

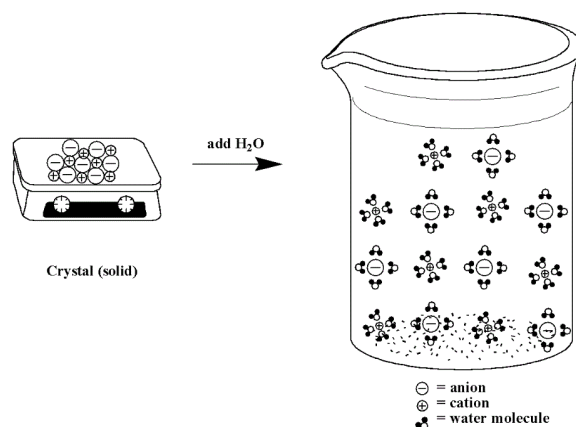


Figure 14.3f: Ionic solids, when placed into a polar solvent such as water, will undergo dissociation. Each ion is surrounded by water molecules due to dipole-dipole interactions, thus pulling the ionic compound apart into individual ions (*General Chemistry I & II*, CC BY 4.0)

If we leave out the water molecules (to make it easier to see the ions) below (Figure 14.3g) is an example of a sodium sulfate, Na_2SO_4 , solution. The sulfate anions will have a -2 charge and the sodium cations will have a +1 charge.

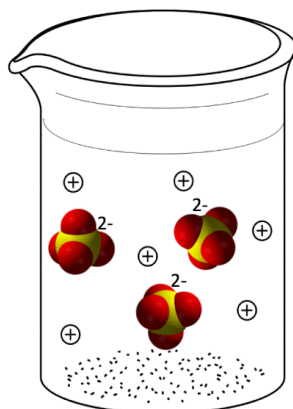


Figure 14.3g: Sodium sulfate (Na_2SO_4), when placed into a polar solvent such as water, will dissociate into individual Na^+ ions and SO_4^{2-} ions (*General Chemistry I & II*, CC BY 4.0)

Note that when the ionic compound dissociates, the covalent bonds of any polyatomic ions (like sulfate) remain intact. Also note that the ions in solution are present in the same ratio as in the solid compound. In this case, there are 2 Na^+ ions for every 1 SO_4^{2-} ion.

Example 14.3h

Determining Moles of Ions in Ionic Compounds:

For each of the following, if 1 mole of the compound were to dissolve and dissociate, how many moles of ions would be present in:

- magnesium chloride?
- ammonium phosphate?

Solution:

- The formula for magnesium chloride is MgCl_2 . It is made of Mg^{2+} ions and Cl^- ions. If a mole of MgCl_2 dissociates in solution, we'll get 1 mole of Mg^{2+} ions and 2 moles of Cl^- ions, for a total of 3 moles of ions.
- The formula for ammonium phosphate is $(\text{NH}_4)_3\text{PO}_4$. It is made of NH_4^+ ions and PO_4^{3-} ions. If a mole of $(\text{NH}_4)_3\text{PO}_4$ dissociates in solution, we'll get 3 moles of NH_4^+ ions and 1 mole of PO_4^{3-} ions, for a total of 4 moles of ions.

Exercise 14.3k

How many moles of *each* ion will be produced from 1 mole of ammonium sulfate?

Check Your Answer⁹

Because an ionic compound dissociates in solution, we can use our knowledge of molarity to describe the concentration of the compound in solution, OR the concentration of ions in solution.

Example 14.3i

Relating Concentrations of a Compound with Concentrations of its Ions in Solution

If you have a solution that is 0.5 M potassium sulfate, what is the concentration of potassium ions in solution? of sulfate ions in solution?

Solution

To answer the questions above, we want to know how many moles of each ion are present in a litre of solution. We are given that the solution is 0.5 M K_2SO_4 . Given the definition of molarity, we can express the following:

$$0.5 \text{ M } K_2SO_4 = \frac{0.5 \text{ mol } K_2SO_4}{1 \text{ L}}$$

We can then use the appropriate relationship we know from the formula of potassium sulfate to make the appropriate conversions:

$$\frac{0.5 \text{ mol } K_2SO_4}{1 \text{ L}} \times \frac{2 \text{ mol } K^+}{1 \text{ mol } K_2SO_4} = \frac{1 \text{ mol } K^+}{1 \text{ L}} = 1 \text{ M } K^+$$

$$\frac{0.5 \text{ mol } K_2SO_4}{1 \text{ L}} \times \frac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } K_2SO_4} = \frac{0.5 \text{ mol } SO_4^{2-}}{1 \text{ L}} = 0.5 \text{ M } SO_4^{2-}$$

Does this result make sense? We see that the concentration of potassium ions is twice that of the sulfate ions. Given that when the ions dissociate, there will be twice as many potassium ions present compared to sulfate ion, this result make sense.

Exercise 14.3l

If you have a 1.3 M solution of ammonium carbonate, $(\text{NH}_4)_2\text{CO}_3$, what is the concentration of each ion?

Check Your Answer¹⁰

Link to Interactive Learning Tools

Explore Dissociation (<https://www.physicsclassroom.com/Concept-Builders/Chemistry/Dissociation>) from the Physics Classroom (<https://www.physicsclassroom.com/>).

Explore Molarity Ranking Tasks (<https://www.physicsclassroom.com/Concept-Builders/Chemistry/Molarity-Ranking-Tasks>) from the Physics Classroom (<https://www.physicsclassroom.com/>).

Explore Molarity Calculations (<https://www.physicsclassroom.com/Concept-Builders/Chemistry/Molarity-Calculations>) from the Physics Classroom (<https://www.physicsclassroom.com/>).

Key Equations

- $M = \frac{\text{mol solute}}{\text{L solution}}$
- $C_1V_1 = C_2V_2$

Attribution & References

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Notes

1. 0.05 M
2. 80 mL
3. 0.674 M
4. 5.55 g CaCl₂
5. 0.370 L
6. 0.102 M CH₃OH
7. 3.76 L
8. 0.261 L
9. 2 moles of NH₄⁺, 1 mole of SO₄²⁻
10. 2.6 M NH₄⁺, 1.3 M CO₃²⁻

14.4 OTHER UNITS FOR SOLUTION CONCENTRATIONS

Learning Objectives

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

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:

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

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 14.4a) 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 14.4a: Liquid bleach is an aqueous solution of sodium hypochlorite (NaOCl). This brand has a concentration of 7.4% NaOCl by mass (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Example 14.4a

Calculation of Percent by Mass

A 5.0-g sample of spinal fluid contains 3.75 mg 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:

$$\% \text{ glucose} = \frac{3.75 \text{ mg glucose} \times \frac{1 \text{ g}}{1000 \text{ mg}}}{5.0 \text{ spinal fluid}} = 0.075\%$$

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.

Exercise 14.4a

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?

Check Your Answer¹

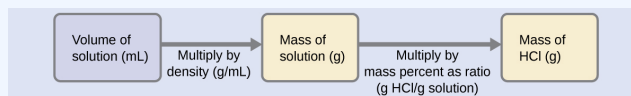
Example 14.4b

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 $5 \times 40 = 200$ 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:

$$500 \text{ mL solution} \left(\frac{1.19 \text{ g solution}}{\text{mL solution}} \right) \left(\frac{37.2 \text{ g HCl}}{100 \text{ g solution}} \right) = 221 \text{ g HCl}$$

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

Exercise 14.4b

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

Check Your Answer

2

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)%:

$$\text{volume percentage} = \frac{\text{volume solute}}{\text{volume solution}} \times 100\%$$

Example 14.3c

Calculations using Volume Percentage

Rubbing alcohol (isopropanol) is usually sold as a 70%vol aqueous solution. If the density of isopropyl

alcohol is 0.785 g/mL, how many grams of isopropyl alcohol are present in a 355 mL bottle of rubbing alcohol?

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:

$$(355 \text{ mL solution}) \left(\frac{70 \text{ isopropyl alcohol}}{100 \text{ mL solution}} \right) \left(\frac{0.785 \text{ g isopropyl alcohol}}{1 \text{ mL isopropyl alcohol}} \right) = 195 \text{ g isopropyl alcohol}$$

Exercise 14.4c

Wine is approximately 12% ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) 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?

Check Your Answer³

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 decilitre (100 mL) of blood (Figure 14.4b).

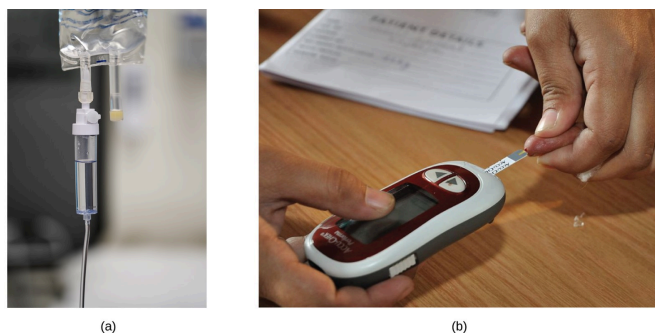


Figure 14.4b: “Mixed” mass-volume units are commonly encountered in medical settings. (a) The NaCl concentration of physiological saline is 0.9% (m/v). (b) This device measures glucose levels in a sample of blood. The normal range for glucose concentration in blood (fasting) is around 70–100 mg/dL. (credit a: work by Samuel Ramos, Unsplash license; credit b: modification of work by Biswarup Ganguly, CC BY 3.0)

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:

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

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

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 14.4c).

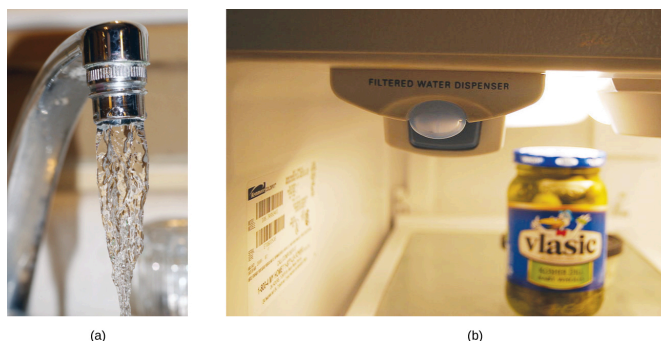


Figure 14.4c: Ppm and ppb are often used to describe very low solute concentrations in solutions. (a) In some areas, trace-level concentrations of contaminants can render unfiltered tap water unsafe for drinking and cooking. (b) Inline water filters reduce the concentration of solutes in tap water. (credit a: modification of work by Jenn Durfey, CC BY 2.0; credit b: modification of work by vastateparkstaff, CC BY 2.0)

Example 14.4d

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 ($1 \text{ ppm} = 10^3 \text{ ppb}$). Thus:

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

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 ($\sim 1.00 \text{ g/mL}$), 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:

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

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

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

Finally, convert this mass to the requested unit of micrograms:

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

Exercise 14.4d

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.

Check Your Answer⁴

Key Equations

- Percent by mass = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
- Volume percentage = $\frac{\text{volume solute}}{\text{volume solution}} \times 100\%$
- ppm = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 \text{ ppm}$
- ppb = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9 \text{ ppb}$

Attribution & References

Except where otherwise noted, this page is adapted by Gregory A. Anderson from “6.4 Other Units for Solution Concentrations (<https://boisestate.pressbooks.pub/chemistry/chapter/6-4-other-units-for-solution-concentrations/>)” In *General Chemistry 1 & 2* by Rice University, a derivative of *Chemistry (OpenStax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson and is licensed under CC BY 4.0. Access for free at *Chemistry (OpenStax)* (<https://openstax.org/books/chemistry/pages/1-introduction>).

Notes

1. 14.8%
2. 282 mL
3. 1.5 mol ethanol
4. 9.6 ppm, 9600 ppb

14.5 COLLIGATIVE PROPERTIES AND OSMOSIS

Learning Objectives

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

- Express concentrations of solution components using mole fraction and molality
- Describe the effect of solute concentration on various solution properties (vapour pressure, boiling point, freezing point, and osmotic pressure)
- Perform calculations using the mathematical equations that describe these various colligative effects
- Describe the process of distillation and its practical applications
- 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 denser, 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 vapour 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.

Mole Fraction and Molality

Several units commonly used to express the concentrations of solution components were introduced in an earlier chapter of this text, each providing certain benefits for use in different applications. For example, molarity (M) is a convenient unit for use in stoichiometric calculations, since it is defined in terms of the molar amounts of solute species:

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

Because solution volumes vary with temperature, molar concentrations will likewise vary. When expressed as molarity, the concentration of a solution with identical numbers of solute and solvent species will be different at different temperatures, due to the contraction/expansion of the solution. More appropriate for calculations involving many colligative properties are mole-based concentration units whose values are not dependent on temperature. Two such units are *mole fraction* (introduced in the previous chapter on gases) and *molality*.

The mole fraction, X , of a component is the ratio of its molar amount to the total number of moles of all solution components:

$$X_A = \frac{\text{mol A}}{\text{total mol of all components}}$$

Molality is a concentration unit defined as the ratio of the numbers of moles of solute to the mass of the solvent in kilograms:

$$m = \frac{\text{mol solute}}{\text{kg solvent}}$$

Since these units are computed using only masses and molar amounts, they do not vary with temperature and, thus, are better suited for applications requiring temperature-independent concentrations, including several colligative properties, as will be described in this chapter module.

Example 14.5a

Calculating Mole Fraction and Molality

The antifreeze in most automobile radiators is a mixture of equal volumes of ethylene glycol and water, with minor amounts of other additives that prevent corrosion. What are the (a) mole fraction and (b) molality of ethylene glycol, $\text{C}_2\text{H}_4(\text{OH})_2$, in a solution prepared from 2.22×10^3 g of ethylene glycol and 2.00×10^3 g of water (approximately 2 L of glycol and 2 L of water)?

Solution

(a) The mole fraction of ethylene glycol may be computed by first deriving molar amounts of both solution components and then substituting these amounts into the unit definition.

$$\begin{aligned} \text{mol C}_2\text{H}_4(\text{OH})_2 &= 2220 \text{ g} \times \frac{1 \text{ mol C}_2\text{H}_4(\text{OH})_2}{62.07 \text{ g C}_2\text{H}_4(\text{OH})_2} = 35.8 \text{ mol C}_2\text{H}_4(\text{OH})_2 \\ \text{mol H}_2\text{O} &= 2000 \text{ g} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 11.1 \text{ mol H}_2\text{O} \end{aligned}$$

$$X_{\text{ethylene glycol}} = \frac{35.8 \text{ mol C}_2\text{H}_4(\text{OH})_2}{(35.8 + 11.1) \text{ mol total}} = 0.763$$

Notice that mole fraction is a dimensionless property, being the ratio of properties with identical units (moles).

(b) To find molality, we need to know the moles of the solute and the mass of the solvent (in kg).

First, use the given mass of ethylene glycol and its molar mass to find the moles of solute:

$$2220 \text{ g C}_2\text{H}_4(\text{OH})_2 \left(\frac{\text{mol C}_2\text{H}_4(\text{OH})_2}{62.07 \text{ g}} \right) = 35.8 \text{ mol C}_2\text{H}_4(\text{OH})_2$$

Then, convert the mass of the water from grams to kilograms:

$$2000 \text{ g H}_2\text{O} \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) = 2 \text{ kg H}_2\text{O}$$

Finally, calculate molality per its definition:

$$\text{molality} = \frac{\text{mol solute}}{\text{kg solvent}}$$

$$\text{molality} = \frac{35.8 \text{ mol C}_2\text{H}_4(\text{OH})_2}{2 \text{ kg H}_2\text{O}}$$

$$\text{molality} = 17.9 \text{ } m$$

Exercise 14.5a

What are the mole fraction and molality of a solution that contains 0.850 g of ammonia, NH_3 , dissolved in 125 g of water?

Check Your Answer¹

Example 14.5b

Converting Mole Fraction and Molal Concentrations

Calculate the mole fraction of solute and solvent in a 3.0 m solution of sodium chloride.

Solution

Converting from one concentration unit to another is accomplished by first comparing the two unit definitions. In this case, both units have the same numerator (moles of solute) but different denominators. The provided molal concentration may be written as:

$$\frac{3.0 \text{ mol NaCl}}{1.0 \text{ kg H}_2\text{O}}$$

The numerator for this solution's mole fraction is, therefore, 3.0 mol NaCl. The denominator may be computed by deriving the molar amount of water corresponding to 1.0 kg

$$1.0 \text{ kg H}_2\text{O} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \left(\frac{\text{mol H}_2\text{O}}{18.02 \text{ g}} \right) = 55 \text{ mol H}_2\text{O}$$

and then substituting these molar amounts into the definition for mole fraction.

$$X_{\text{H}_2\text{O}} = \frac{\text{mol H}_2\text{O}}{\text{mol NaCl} + \text{mol H}_2\text{O}}$$

$$X_{\text{H}_2\text{O}} = \frac{55 \text{ mol H}_2\text{O}}{3.0 \text{ mol NaCl} + 55 \text{ mol H}_2\text{O}}$$

$$X_{\text{H}_2\text{O}} = 0.95$$

$$X_{\text{NaCl}} = \frac{\text{mol NaCl}}{\text{mol NaCl} + \text{mol H}_2\text{O}}$$

$$X_{\text{NaCl}} = \frac{3.0 \text{ mol NaCl}}{3.0 \text{ mol NaCl} + 55 \text{ mol H}_2\text{O}}$$

$$X_{\text{NaCl}} = 0.052$$

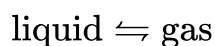
Exercise 14.5b

The mole fraction of iodine, I_2 , dissolved in dichloromethane, CH_2Cl_2 , is 0.115. What is the molal concentration, m , of iodine in this solution?

Check Your Answer²

Vapour Pressure Lowering

As described in the chapter on liquids and solids, the equilibrium vapour pressure of a liquid is the pressure exerted by its gaseous phase when vaporization and condensation are occurring at equal rates:



Dissolving a nonvolatile substance in volatile liquid results in a lowering of the liquid's vapour pressure. This phenomenon can be rationalized by considering the effect of added solute molecules on the liquid's vaporization and condensation processes. To vaporize, solvent molecules must be present at the surface of the solution. The presence of solute decreases the surface area available to solvent molecules and thereby reduces the rate of solvent vaporization. Since the rate of condensation is unaffected by the presence of solute, the net result is that the vaporization-condensation equilibrium is achieved with fewer solvent molecules in the vapour phase (i.e., at a lower vapour pressure) (Figure 14.5a). While this kinetic interpretation is useful, it does not account for several important aspects of the colligative nature of vapour pressure lowering. A more rigorous explanation involves the property of *entropy*, a topic of discussion in a later chapter on thermodynamics. For purposes of understanding the lowering of a liquid's vapour pressure, it is adequate to note that the greater entropy of a solution in comparison to its separate solvent and solute serves to effectively stabilize the solvent molecules and hinder their vaporization. A lower vapour pressure result and a correspondingly higher boiling point is described in the next section of this module.

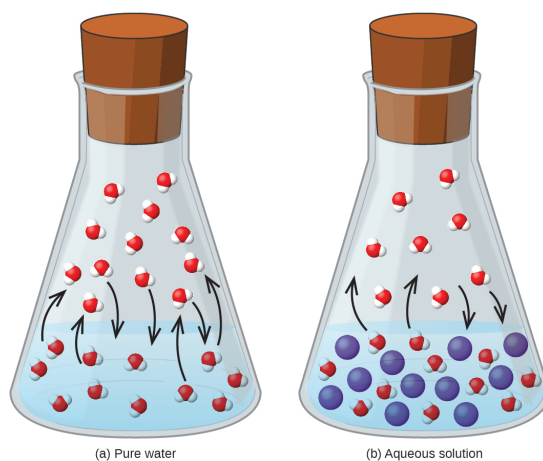


Figure 14.5a: The presence of nonvolatile solutes lowers the vapour pressure of a solution by impeding the evaporation of solvent molecules (credit: *Chemistry (OpenStax)*, CC BY 4.0).

The relationship between the vapour pressures of solution components and the concentrations of those components is described by **Raoult's law**: *The partial pressure exerted by any component of an ideal solution is equal to the vapour pressure of the pure component multiplied by its mole fraction in the solution.*

$$P_A = X_A P_A^\circ$$

where P_A is the partial pressure exerted by component A in the solution, P_A° is the vapour pressure of pure A, and X_A is the mole fraction of A in the solution. (Mole fraction is a concentration unit introduced in the chapter on gases.)

Recalling that the total pressure of a gaseous mixture is equal to the sum of partial pressures for all its components (Dalton's law of partial pressures), the total vapour pressure exerted by a solution containing i components is

$$P_{\text{solution}} = \sum_i P_i = \sum_i X_i P_i^\circ$$

A nonvolatile substance is one whose vapour pressure is negligible ($P^\circ \approx 0$), and so the vapour pressure above a solution containing only nonvolatile solutes is due only to the solvent:

$$P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}}^\circ$$

Example 14.5c

Calculation of a Vapour Pressure

Compute the vapour pressure of an ideal solution containing 92.1 g of glycerin, $\text{C}_3\text{H}_5(\text{OH})_3$, and 184.4 g of ethanol, $\text{C}_2\text{H}_5\text{OH}$, at 40 °C. The vapour pressure of pure ethanol is 0.178 atm at 40 °C. Glycerin is essentially nonvolatile at this temperature.

Solution

Since the solvent is the only volatile component of this solution, its vapour pressure may be computed per Raoult's law as:

$$P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}}^\circ$$

First, calculate the molar amounts of each solution component using the provided mass data.

$$92.1 \text{ g } \text{C}_3\text{H}_5(\text{OH})_3 \times \frac{1 \text{ mol } \text{C}_3\text{H}_5(\text{OH})_3}{92.094 \text{ g } \text{C}_3\text{H}_5(\text{OH})_3} = 1.00 \text{ mol } \text{C}_3\text{H}_5(\text{OH})_3$$

$$184.4 \text{ g } \text{C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol } \text{C}_2\text{H}_5\text{OH}}{46.069 \text{ g } \text{C}_2\text{H}_5\text{OH}} = 4.000 \text{ mol } \text{C}_2\text{H}_5\text{OH}$$

Next, calculate the mole fraction of the solvent (ethanol) and use Raoult's law to compute the solution's vapour pressure.

$$X_{\text{C}_2\text{H}_5\text{OH}} = \frac{4.000 \text{ mol}}{(1.00 \text{ mol} + 4.000 \text{ mol})} = 0.800$$

$$P_{\text{soln}} = X_{\text{soln}} P_{\text{soln}}^\circ = 0.800 \times 0.178 \text{ atm} = 0.142 \text{ atm}$$

Exercise 14.5c

A solution contains 5.00 g of urea, $\text{CO}(\text{NH}_2)_2$ (a nonvolatile solute) and 0.100 kg of water. If the vapour pressure of pure water at 25 °C is 23.7 torr, what is the vapour pressure of the solution?

Check Your Answer³

Elevation of the Boiling Point of a Solvent

As described in the chapter on liquids and solids, the *boiling point* of a liquid is the temperature at which its vapour pressure is equal to ambient atmospheric pressure. Since the vapour pressure of a solution is lowered due to the presence of nonvolatile solutes, it stands to reason that the solution's boiling point will subsequently be increased. Compared to pure solvent, a solution, therefore, will require a higher temperature to achieve any given vapour pressure, including one equivalent to that of the surrounding atmosphere. The increase in boiling point observed when nonvolatile solute is dissolved in a solvent, ΔT_b , is called **boiling point elevation** and is directly proportional to the molal concentration of solute species:

$$\Delta T_b = K_b m$$

where K_b is the **boiling point elevation constant**, or the *ebullioscopic constant*, and m is the molal concentration (molality) of all solute species.

Boiling point elevation constants are characteristic properties that depend on the identity of the solvent. Values of K_b for several solvents are listed in Table 14.5a.

Table 14.5a: Boiling Point Elevation and Freezing Point Depression Constants for Several Solvents

Solvent	Boiling Point (°C at 1 atm)	K_b (Cm^{-1})	Freezing Point (°C at 1 atm)	K_f (Cm^{-1})
water	100.0	0.512	0.0	1.86
hydrogen acetate	118.1	3.07	16.6	3.9
benzene	80.1	2.53	5.5	5.12
chloroform	61.26	3.63	-63.5	4.68
nitrobenzene	210.9	5.24	5.67	8.1

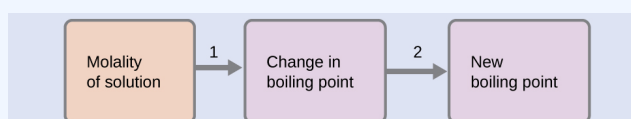
The extent to which the vapour pressure of a solvent is lowered and the boiling point is elevated depends on

the total number of solute particles present in a given amount of solvent, not on the mass or size or chemical identities of the particles. A 1 *m* aqueous solution of sucrose (342 g/mol) and a 1 *m* aqueous solution of ethylene glycol (62 g/mol) will exhibit the same boiling point because each solution has one mole of solute particles (molecules) per kilogram of solvent.

Example 14.5d

Calculating the Boiling Point of a Solution

What is the boiling point of a 0.33 *m* solution of a nonvolatile solute in benzene?



Solution

Use the equation relating boiling point elevation to solute molality to solve this problem in two steps.

1. Calculate the change in boiling point.

$$\Delta T_b = K_b m = 2.53 \text{ } ^\circ\text{C}m^{-1} \times 0.33 \text{ } m = 0.83 \text{ } ^\circ\text{C}$$

2. Add the boiling point elevation to the pure solvent's boiling point.

$$\text{Boiling temperature} = 80.1 \text{ } ^\circ\text{C} + 0.83 \text{ } ^\circ\text{C} = 80.9 \text{ } ^\circ\text{C}$$

Exercise 14.5d

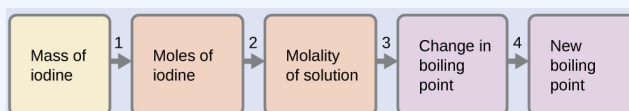
What is the boiling point of the antifreeze with a molality of 17.9 m in Example 14.5a (Hint: Water is the solvent).

Check Your Answer⁴

Example 14.5e

The Boiling Point of an Iodine Solution

Find the boiling point of a solution of 92.1 g of iodine, I_2 , in 800.0 g of chloroform, CHCl_3 , assuming that the iodine is nonvolatile and that the solution is ideal.



Solution

We can solve this problem using four steps.

1. Convert from grams to moles of I_2 using the molar mass of I_2 in the unit conversion factor.
Result: 0.363 mol
 2. Determine the molality of the solution from the number of moles of solute and the mass of solvent, in kilograms.
Result: 0.454 m
 3. Use the direct proportionality between the change in boiling point and molal concentration to determine how much the boiling point changes.
Result: $1.65\text{ }^\circ\text{C}$
 4. Determine the new boiling point from the boiling point of the pure solvent and the change.
Result: $62.91\text{ }^\circ\text{C}$
- Check each result as a self-assessment.

Exercise 14.5e

What is the boiling point of a solution of 1.0 g of glycerin, $C_3H_5(OH)_3$, in 47.8 g of water? Assume an ideal solution.

Check Your Answer⁵

Distillation of Solutions

Distillation is a technique for separating the components of mixtures that is widely applied both in the laboratory and in industrial settings. It is used to refine petroleum, to isolate fermentation products, and to purify water. This separation technique involves the controlled heating of a sample mixture to selectively vaporize, condense, and collect one or more components of interest. A typical apparatus for laboratory-scale distillations is shown in Figure 14.5b.

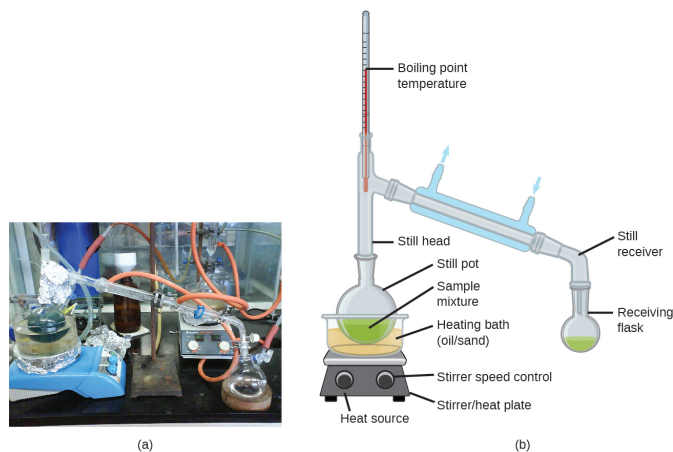


Figure 14.5b: A typical laboratory distillation unit is shown in (a) a photograph and (b) a schematic diagram of the components. (credit a: modification of work by Rifleman82, PD; credit b: modification of work adapted by Slashme, PD in *Chemistry (OpenStax)*, CC BY 4.0).

Oil refineries use large-scale *fractional distillation* to separate the components of crude oil. The crude oil is heated to high temperatures at the base of a tall *fractionating column*, vaporizing many of the components that rise within the column. As vaporized components reach adequately cool zones during their ascent, they condense and are collected. The collected liquids are simpler mixtures of hydrocarbons and other petroleum

compounds that are of appropriate composition for various applications (e.g., diesel fuel, kerosene, gasoline), as depicted in Figure 14.5c.

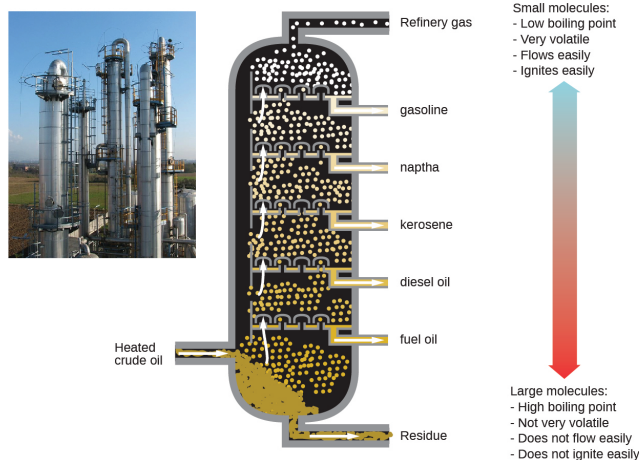


Figure 14.5c: Crude oil is a complex mixture that is separated by large-scale fractional distillation to isolate various simpler mixtures. (credit: modification of work by Luigi Chiesa, CC BY 3.0 in *Chemistry (OpenStax)*, CC BY 4.0.)

Depression of the Freezing Point of a Solvent

Solutions freeze at lower temperatures than pure liquids. This phenomenon is exploited in “de-icing” schemes that use salt (Figure 14.5d), calcium chloride, or urea to melt ice on roads and sidewalks, and in the use of ethylene glycol as an “antifreeze” in automobile radiators. Seawater freezes at a lower temperature than freshwater, and so the Arctic and Antarctic oceans remain unfrozen even at temperatures below 0 °C (as do the body fluids of fish and other cold-blooded sea animals that live in these oceans).



Figure 14.5d: Rock salt (NaCl), calcium chloride (CaCl₂), or a mixture of the two are used to melt ice. (credit: modification of work by Eddie Welker, CC BY 2.0)

The decrease in the freezing point of a dilute solution compared to that of the pure solvent, ΔT_f , is called the **freezing point depression** and is directly proportional to the molal concentration of the solute

$$\Delta T_f = K_f m$$

where m is the molal concentration of the solute in the solvent and K_f is called the **freezing point depression constant** (or *cryoscopic constant*). Just as for boiling point elevation constants, these are characteristic properties whose values depend on the chemical identity of the solvent. Values of K_f for several solvents are listed in Table 14.5a.

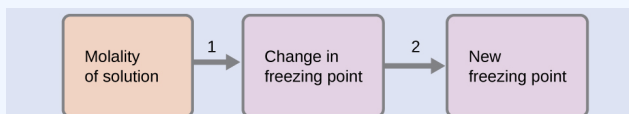
Example 14.5f

Calculation of the Freezing Point of a Solution

What is the freezing point of the 0.33 m solution of a nonvolatile nonelectrolyte solute in benzene described in Example 14.5b?

Solution

Use the equation relating freezing point depression to solute molality to solve this problem in two steps.



1. Calculate the change in freezing point.

$$\Delta T_f = K_f m = 5.12 \text{ }^\circ\text{C}m^{-1} \times 0.33m = 1.7 \text{ }^\circ\text{C}$$

2. Subtract the freezing point change observed from the pure solvent's freezing point.

$$\text{Freezing temperature} = 5.5 \text{ }^\circ\text{C} - 1.7 \text{ }^\circ\text{C} = 3.8 \text{ }^\circ\text{C}$$

Exercise 14.5f

What is the freezing point of a 1.85 *m* solution of a nonvolatile nonelectrolyte solute in nitrobenzene?

Check Your Answer⁶

Colligative Properties and De-Icing

Sodium chloride and its group 2 analogs calcium and magnesium chloride are often used to de-ice roadways and sidewalks, due to the fact that a solution of any one of these salts will have a freezing point lower than 0 °C, the freezing point of pure water. The group 2 metal salts are frequently mixed with the cheaper and more readily available sodium chloride (“rock salt”) for use on roads, since they tend to be somewhat less corrosive than the NaCl, and they provide a larger depression of the freezing point, since they dissociate to yield three particles per formula unit, rather than two particles like the sodium chloride.

Because these ionic compounds tend to hasten the corrosion of metal, they would not be a wise choice to use in antifreeze for the radiator in your car or to de-ice a plane prior to takeoff. For these applications, covalent compounds, such as ethylene or propylene glycol, are often used. The glycols used in radiator fluid not only lower the freezing point of the liquid, but they elevate the boiling point, making the fluid useful in both winter and summer. Heated glycols are often sprayed onto the surface of airplanes prior to takeoff in inclement weather in the winter to remove ice that has already formed and prevent the formation of more ice, which would be particularly dangerous if formed on the control surfaces of the aircraft (Figure 14.5e).

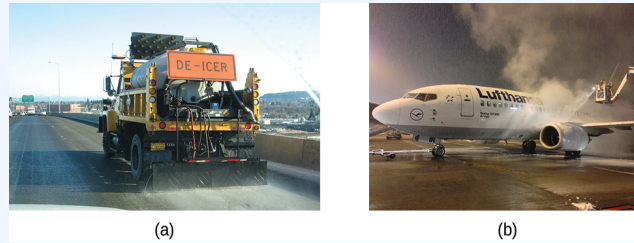


Figure 14.5e: Freezing point depression is exploited to remove ice from (a) roadways and (b) the control surfaces of aircraft. (credit a: work by Oregon Department of Transportation, CC BY 2.0; credit b: work by Paolo Cerutti (<https://www.flickr.com/people/76926545@N05>), CC BY 2.0)

Phase Diagram for an Aqueous Solution of a Nonelectrolyte

The colligative effects on vapour pressure, boiling point, and freezing point described in the previous section are conveniently summarized by comparing the phase diagrams for a pure liquid and a solution derived from that liquid. Phase diagrams for water and an aqueous solution are shown in Figure 14.5f.

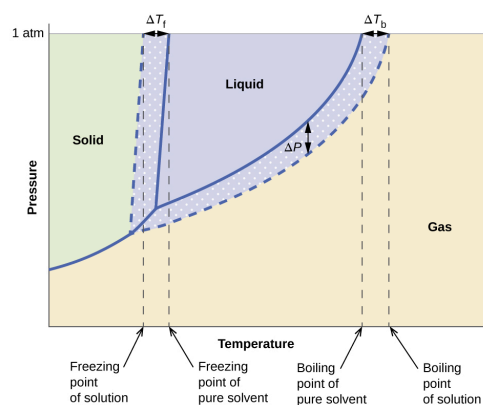


Figure 14.5f: These phase diagrams show water (solid curves) and an aqueous solution of nonelectrolyte (dashed curves) (credit: *Chemistry (OpenStax)*, CC BY 4.0).

The liquid-vapour curve for the solution is located *beneath* the corresponding curve for the solvent, depicting the vapour pressure *lowering*, ΔP , that results from the dissolution of nonvolatile solute. Consequently, at any given pressure, the solution's boiling point is observed at a higher temperature than that for the pure solvent, reflecting the boiling point elevation, ΔT_b , associated with the presence of nonvolatile solute. The solid-liquid curve for the solution is displaced left of that for the pure solvent, representing the freezing point depression, ΔT_b , that accompanies solution formation. Finally, notice that the solid-gas curves for the solvent and its solution are identical. This is the case for many solutions comprising liquid solvents and nonvolatile solutes.

Just as for vaporization, when a solution of this sort is frozen, it is actually just the *solvent* molecules that undergo the liquid-to-solid transition, forming pure solid solvent that excludes solute species. The solid and gaseous phases, therefore, are composed solvent only, and so transitions between these phases are not subject to colligative effects.

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 14.5g, 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 in 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**.

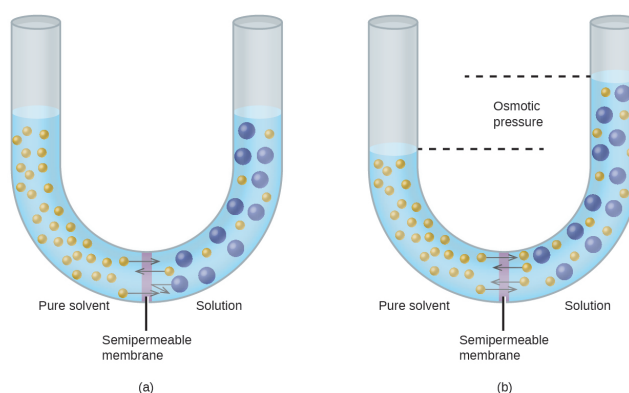


Figure 14.5g Osmosis results in the transfer of solvent molecules from a sample of low (or zero) solute concentration to a sample of higher solute concentration. (a) A semi-permeable membrane is placed between a sample of pure solvent and another solution which contains large molecules that cannot permeate through the membrane. (b) After some time, the small solvent molecules are able to permeate back-and-forth through the membrane until equilibrium is reached, while the solution molecules remain on the one side. (credit: *Chemistry (OpenStax)*, CC BY 4.0).

When osmosis is carried out in an apparatus like that shown in Figure 14.5g, 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

$$\Pi = MRT$$

where R is the universal gas constant.

Example 14.5g

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 $\Pi = MRT$, 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).

$$\begin{aligned}\Pi &= MRT \\ &= 0.30 \text{ mol/L} \times 0.08206 \text{ L atm/mol K} \times 310 \text{ K} \\ &= 7.6 \text{ atm}\end{aligned}$$

Exercise 14.5g

What is the osmotic pressure (atm) a solution with a volume of 0.750 L that contains 5.0 g of methanol, CH_3OH , in water at 37 °C

Check Your Answer⁷

If a solution is placed in an apparatus like the one shown in Figure 14.5h, 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.

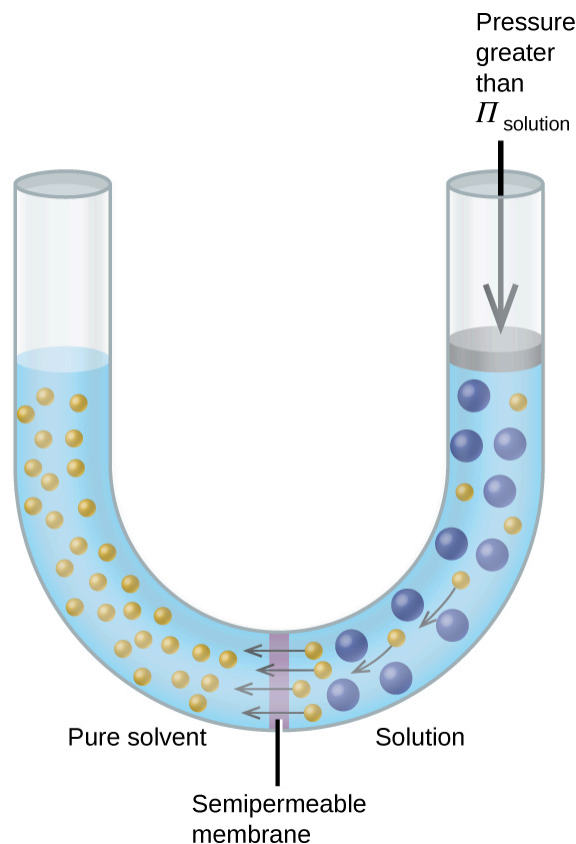


Figure 14.5h Applying a pressure greater than the osmotic pressure of a solution will result in reverse osmosis. Solvent molecules from the solution are pushed into the pure solvent side of the mechanism, leaving the solution molecules behind. (credit: *Chemistry (OpenStax)*, CC BY 4.0).

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, 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 transported in vehicles to remote locations.

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 14.5i.

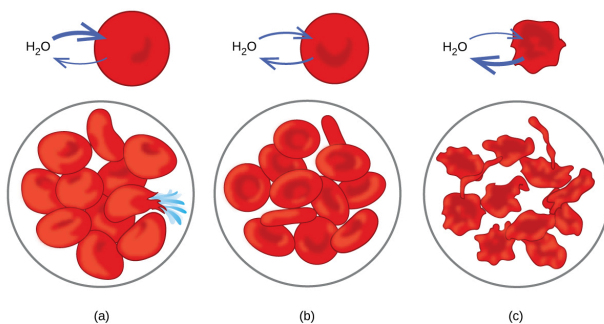


Figure 14.5i: Red blood cell membranes are water permeable and will (a) swell and possibly rupture in a hypotonic solution; (b) maintain normal volume and shape in an isotonic solution; and (c) shrivel and possibly die in a hypertonic solution. (credit a/b/c: modifications of work by LadyofHats, PD in *Chemistry (OpenStax)*, CC BY 4.0).

Determination of Molar Masses

Osmotic pressure and changes in freezing point, boiling point, and vapour 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.

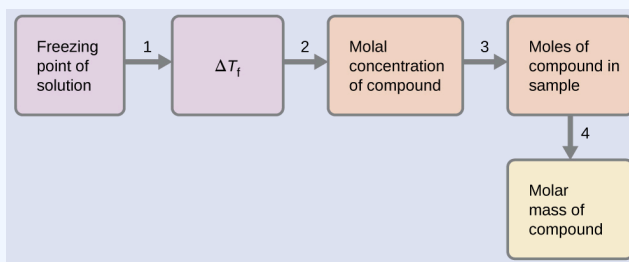
Example 14.5h

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.



1. Determine the change in freezing point from the observed freezing point and the freezing point of pure benzene (Table 14.5a).

$$\Delta T_f = 5.5 \text{ }^\circ\text{C} - 2.32 \text{ }^\circ\text{C} = 3.2 \text{ }^\circ\text{C}$$

2. Determine the molal concentration from K_f , the freezing point depression constant for benzene (Table 14.5a), and ΔT_f .

$$\Delta T_f = K_f m$$

$$m = \frac{\Delta T_f}{K_f} = \frac{3.2 \text{ }^\circ\text{C}}{5.12 \text{ }^\circ\text{C}m^{-1}} = 0.63m$$

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.

$$\text{Moles of solute} = \frac{0.62 \text{ mol solute}}{1.00 \text{ kg solvent}} \times 0.0550 \text{ kg solvent} = 0.035 \text{ mol}$$

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

$$\text{Molar mass} = \frac{4.00 \text{ g}}{0.034 \text{ mol}} = 1.2 \times 10^2 \text{ g/mol}$$

Exercise 14.5h

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?

Check Your Answer⁸

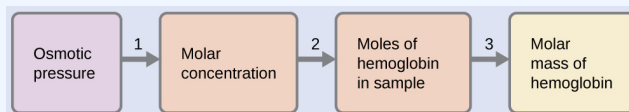
Example 14.5i

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:



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

$$\Pi = \frac{5.9 \text{ torr} \times 1 \text{ atm}}{760 \text{ torr}} = 7.8 \times 10^{-3} \text{ atm}$$

$$\Pi = MRT$$

$$M = \frac{\Pi}{RT} = \frac{7.8 \times 10^{-3} \text{ atm}}{(0.08206 \text{ L atm/mol K})(295 \text{ K})} = 3.2 \times 10^{-4} \text{ M}$$

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

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

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

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

Exercise 14.5i

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

Check Your Answer⁹

Colligative Properties of Electrolytes

As noted previously in this module, the colligative properties of a solution depend only on the number, not on the kind, of solute species dissolved. For example, 1 mole of any nonelectrolyte dissolved in 1 kilogram of solvent produces the same lowering of the freezing point as does 1 mole of any other nonelectrolyte. However, 1 mole of sodium chloride (an electrolyte) forms *2 moles* of ions when dissolved in solution. Each individual ion produces the same effect on the freezing point as a single molecule does.

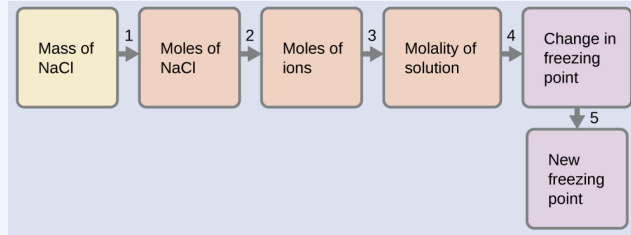
Example 14.5j

The Freezing Point of a Solution of an Electrolyte

The concentration of ions in seawater is approximately the same as that in a solution containing 4.2 g of NaCl dissolved in 125 g of water. Assume that each of the ions in the NaCl solution has the same effect on the freezing point of water as a nonelectrolyte molecule, and determine the freezing temperature the solution (which is approximately equal to the freezing temperature of seawater).

Solution

We can solve this problem using the following series of steps.



1. Convert from grams to moles of NaCl using the molar mass of NaCl in the unit conversion factor.
Result: 0.072 mol NaCl
 2. Determine the number of moles of ions present in the solution using the number of moles of ions in 1 mole of NaCl as the conversion factor (2 mol ions/1 mol NaCl).
Result: 0.14 mol ions
 3. Determine the molality of the ions in the solution from the number of moles of ions and the mass of solvent, in kilograms.
Result: 1.1 m
 4. Use the direct proportionality between the change in freezing point and molal concentration to determine how much the freezing point changes.
Result: 2.0 °C
 5. Determine the new freezing point from the freezing point of the pure solvent and the change.
Result: -2.0 °C
- Check each result as a self-assessment.

Exercise 14.5j

Assume that each of the ions in calcium chloride, CaCl_2 , has the same effect on the freezing point of water as a nonelectrolyte molecule. Calculate the freezing point of a solution of 0.724 g of CaCl_2 in 175 g of water.

Check Your Answer¹⁰

Assuming complete dissociation, a 1.0 m aqueous solution of NaCl contains 1.0 mole of ions (1.0 mol Na^+ and 1.0 mol Cl^-) per each kilogram of water, and its freezing point depression is expected to be

$$\Delta T_f = 2.0 \text{ mol ions/kg water} \times 1.86 \text{ }^\circ\text{C kg water/mol ion} = 3.7 \text{ }^\circ\text{C}$$

When this solution is actually prepared and its freezing point depression measured, however, a value of 3.4 °C is obtained. Similar discrepancies are observed for other ionic compounds, and the differences between the measured and expected colligative property values typically become more significant as solute concentrations increase. These observations suggest that the ions of sodium chloride (and other strong electrolytes) are not completely dissociated in solution.

To account for this and avoid the errors accompanying the assumption of total dissociation, an experimentally measured parameter named in honour of Nobel Prize-winning German chemist Jacobus Henricus van't Hoff is used. The **van't Hoff factor (i)** is defined as the ratio of solute particles in solution to the number of formula units dissolved:

$$i = \frac{\text{moles of particles in solution}}{\text{moles of formula units dissolved}}$$

Values for measured van't Hoff factors for several solutes, along with predicted values assuming complete dissociation, are shown in Table 14.5b.

Table 14.5b: Expected and Observed van't Hoff Factors for Several 0.050 *m* Aqueous Electrolyte Solutions

Electrolyte	Particles in Solution	i (Predicted)	i (Measured)
HCl	H ⁺ , Cl ⁻	2	1.9
NaCl	Na ⁺ , Cl ⁻	2	1.9
MgSO ₄	Mg ²⁺ , SO ₄ ²⁻	2	1.3
MgCl ₂	Mg ²⁺ , 2Cl ⁻	3	2.7
FeCl ₃	Fe ³⁺ , 3Cl ⁻	4	3.4
glucose ¹¹	C ₁₂ H ₂₂ O ₁₁	1	1.0

In 1923, the chemists Peter Debye and Erich Hückel proposed a theory to explain the apparent incomplete ionization of strong electrolytes. They suggested that although interionic attraction in an aqueous solution is very greatly reduced by solvation of the ions and the insulating action of the polar solvent, it is not completely nullified. The residual attractions prevent the ions from behaving as totally independent particles (Figure 14.5j). In some cases, a positive and negative ion may actually touch, giving a solvated unit called an ion pair. Thus, the activity, or the effective concentration, of any particular kind of ion is less than that indicated by the actual concentration. Ions become more and more widely separated the more dilute the solution, and the residual interionic attractions become less and less. Thus, in extremely dilute solutions, the effective concentrations of the ions (their activities) are essentially equal to the actual concentrations. Note that the van't Hoff factors for the electrolytes in Table 14.5b are for 0.05 *m* solutions, at which concentration the value of i for NaCl is 1.9, as opposed to an ideal value of 2.

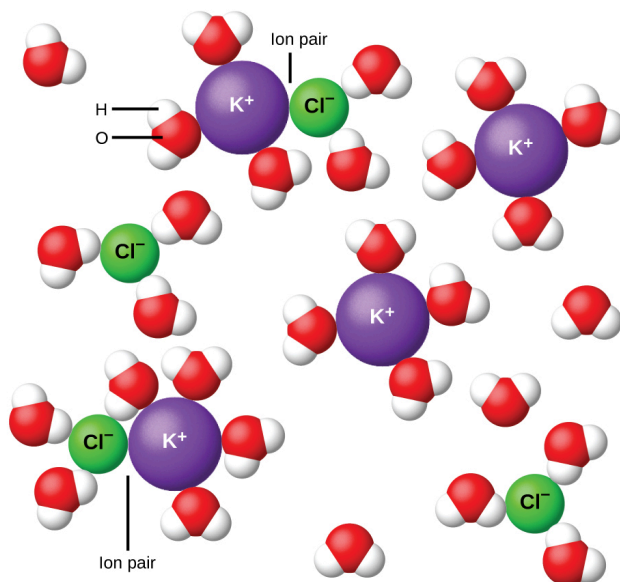


Figure 14.5j: Ions become more and more widely separated the more dilute the solution, and the residual interionic attractions become less (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Key Equations

- $(P_A = X_A P_A^\circ)$
- $P_{\text{solution}} = \sum_i P_i = \sum_i X_i P_i^\circ$
- $P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}}^\circ$
- $\Delta T_b = K_b m$
- $\Delta T_f = K_f m$
- $\Pi = MRT$

Attribution & References

Except where otherwise noted, this page is adapted by Gregory A. Anderson from “11.4 Colligative Properties (<https://boisestate.pressbooks.pub/chemistry/chapter/11-4-colligative-properties/>)” In *General Chemistry 1 & 2* by Rice University, a derivative of *Chemistry (Open Stax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson and is licensed under CC BY 4.0. Access for free at *Chemistry (OpenStax)* (<https://openstax.org/books/chemistry/pages/1-introduction>).

Notes

1. 7.14×10^{-3} ; $0.399\ m$
2. $1.50\ m$
3. $23.4\ \text{torr}$
4. $109.2\ ^\circ\text{C}$
5. $100.12\ ^\circ\text{C}$
6. $-9.3\ ^\circ\text{C}$
7. $5.3\ \text{atm}$
8. $1.8 \times 10^2\ \text{g/mol}$
9. $2.7 \times 10^4\ \text{g/mol}$
10. $-0.208\ ^\circ\text{C}$
11. A nonelectrolyte shown for comparison.

14.6 COLLOIDS

Learning Objectives

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 behaviour from suspensions. A solution may be coloured, 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 14.6a). 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.

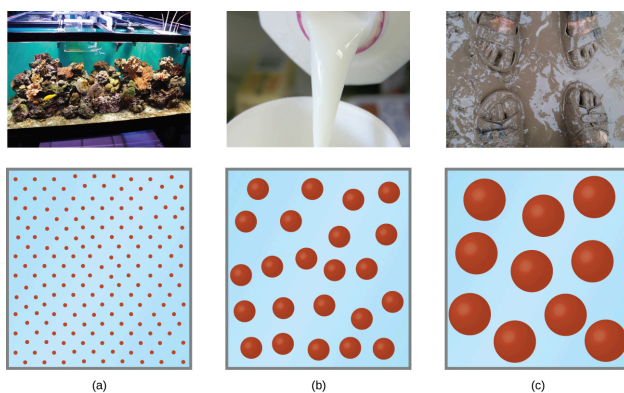


Figure 14.6a: (a) A solution is a homogeneous mixture that appears clear, such as the saltwater in this aquarium. (b) In a colloid, such as milk, the particles are much larger but remain dispersed and do not settle. (c) A suspension, such as mud, is a heterogeneous mixture of suspended particles that appears cloudy and in which the particles can settle. (credit a photo: modification of work by Adam Wimsatt, CC BY 2.0; credit b photo: modification of work by Melissa Wiese, CC BY 2.0; credit c photo: modification of work by Peter Burgess, CC BY 2.0; in *Chemistry (OpenStax)*, CC BY 4.0).

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 14.6b. 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 14.6b: The paths of searchlight beams are made visible when light is scattered by colloidal-size particles in the air (fog, smoke, etc.). (credit: work by Bahman, CC BY 2.0)

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 14.6a.

Table 14.6a: Examples of Colloidal Systems

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 coloured gems, some alloys	—
liquid	gas	clouds, fogs, mists, sprays	aerosol
liquid	liquid	milk, mayonnaise, butter	emulsion
liquid	solid	jellies, gels, pearl, opal (H ₂ O in SiO ₂)	gel
gas	liquid	foams, whipped cream, beaten egg whites	foam
gas	solid	pumice, floating soaps	—

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:

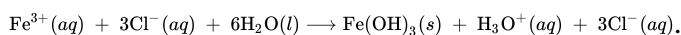
1. Dispersion methods: that is, by breaking down larger particles. For example, paint pigments are produced by dispersing large particles by grinding in special mills.
2. Condensation methods: that is, growth from smaller units, such as molecules or ions. For example, clouds form when water molecules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

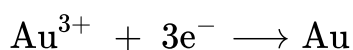
We can prepare an **emulsion** by shaking together or blending two immiscible liquids. This breaks one liquid into droplets of colloidal size, which then disperse throughout the other liquid. Oil spills in the ocean may be difficult to clean up, partly because wave action can cause the oil and water 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:



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:



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 $\text{C}_{17}\text{H}_{35}\text{CO}_2\text{Na}$ and contains an uncharged nonpolar hydrocarbon chain, the $\text{C}_{17}\text{H}_{35}$ — unit, and an ionic carboxylate group, the $-\text{CO}_2^{-}$ unit (Figure 14.6c).

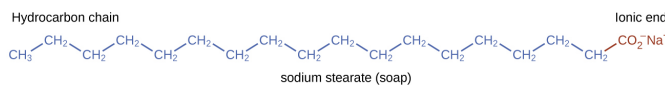


Figure 14.6c: Soaps contain a nonpolar hydrocarbon end (blue) and an ionic end (red). An ionic end is a carboxylate group. The length of the hydrocarbon end can vary from soap to soap (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Detergents (soap substitutes) also contain nonpolar hydrocarbon chains, such as $\text{C}_{12}\text{H}_{25}$ —, and an ionic group, such as a sulfate— OSO_3^{-} , or a sulfonate— SO_3^{-} (Figure 14.6d). Soaps form insoluble calcium and magnesium compounds in hard water; detergents form water-soluble products – a definite advantage for detergents.

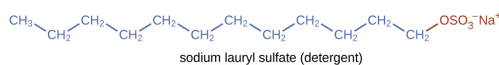


Figure 14.6d: Detergents contain a nonpolar hydrocarbon end (blue) and an ionic end (red). The ionic end can be either a sulfate or a sulfonate. The length of the hydrocarbon end can vary from detergent to detergent (credit: *Chemistry (OpenStax)*, CC BY 4.0).

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 14.6e. 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.

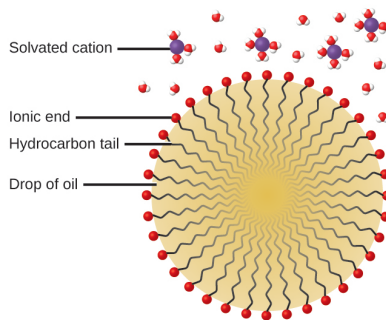


Figure 14.6e: This diagrammatic cross-section of an emulsified drop of oil in water shows how soap or detergent acts as an emulsifier (credit: *Chemistry (OpenStax)*, CC BY 4.0).

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 petroleum. 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 the 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. 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 potentially 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 (Gulf Spill Restoration (<https://www.gulfspillrestoration.noaa.gov/assessment/>)).

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 colloiddally dispersed and electrically charged. Frederick Cottrell, an American chemist, developed a process to remove these particles (Figure 14.6f).

Frederick Gardner Cottrell

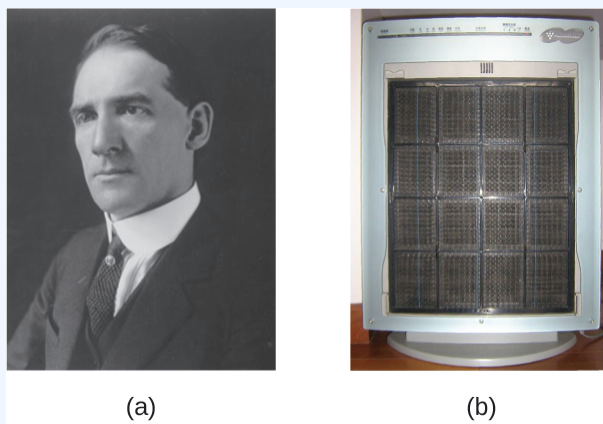


Figure 14.6f: (a) Frederick Cottrell developed (b) the electrostatic precipitator, a device designed to curb air pollution by removing colloidal particles from air. (credit a: work by unknown, PD; credit b: modification of work by SpLot, CC BY-SA 3.0)

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 14.6g). 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.

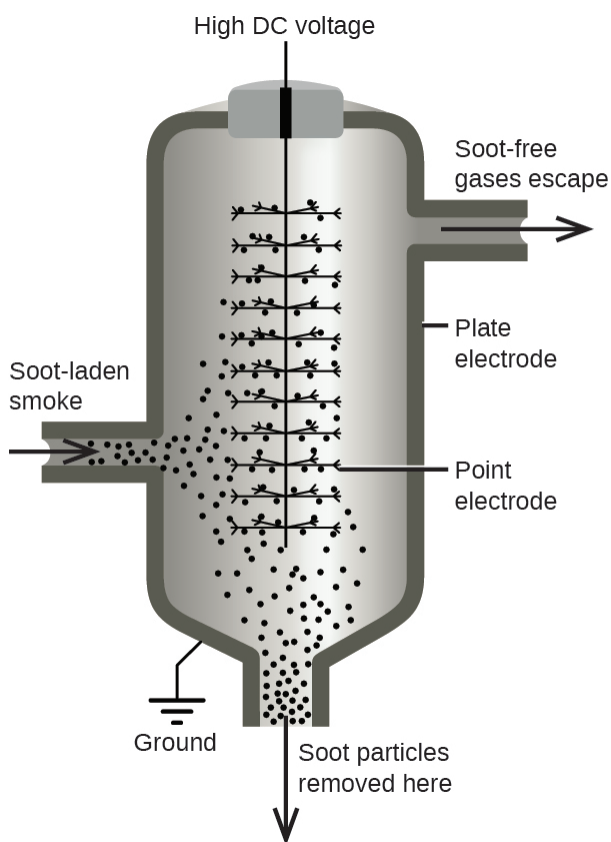


Figure 14.6g: In a Cottrell precipitator, positively and negatively charged particles are attracted to highly charged electrodes, where they are neutralized and deposited as dust (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Gels

When we make gelatin, such as Jell-O, we are making a type of colloid (Figure 14.6h). 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 fibres 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.

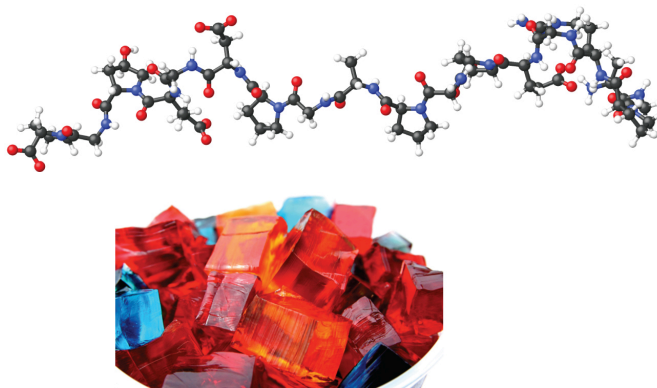


Figure 14.6h: Gelatin desserts are colloids in which an aqueous solution of sweeteners and flavours is dispersed throughout a medium of solid proteins. (credit photo: modification of work by Steven Depolo, CC BY 2.0; in *Chemistry (OpenStax)*, CC BY 4.0.)

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.

Attribution & References

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CHAPTER 14 - SUMMARY

14.1 – Solutions: An introduction

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.

14.2 – 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 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.

14.3 – 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 litre 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.

14.4 – 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.

14.5 – Colligative Properties and Osmosis

Properties of a solution that depend only on the concentration of solute particles are called colligative properties. They include changes in the vapour 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. Ionic compounds may not completely dissociate in solution due to activity effects, in which case observed colligative effects may be less than predicted.

14.6 – 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.

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CHAPTER 14 - REVIEW

14.1 – Solutions: An introduction

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

Check Answer: ¹

2. Which of the principal characteristics of solutions can we see in the solutions of $\text{K}_2\text{Cr}_2\text{O}_7$ shown in Figure 14.1a?
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?

Check Answer: ²

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 14.1a.
 - b. $\text{NO}(l)$ in $\text{CO}(l)$
 - c. $\text{Cl}_2(g)$ in $\text{Br}_2(l)$
 - d. $\text{HCl}(aq)$ in benzene $\text{C}_6\text{H}_6(l)$
 - e. Methanol $\text{CH}_3\text{OH}(l)$ in $\text{H}_2\text{O}(l)$

Check Answer: ³

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. **Check Answer:** ⁴

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.

- Determine the molarity of this solution (solution density = 1.8 g/cm^3).
- Determine the molality of this solution (solution density = 1.8 g/cm^3).
- Determine the percent by mass of hydrogen atoms in this solution (solution density = 1.8 g/cm^3).

14.2 – Solubility

- Suppose you are presented with a clear solution of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$. How could you determine whether the solution is unsaturated, saturated, or supersaturated?
- 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.
Check Answer: ⁵
- Suggest an explanation for the observations that ethanol, $\text{C}_2\text{H}_5\text{OH}$, is completely miscible with water and that ethanethiol, $\text{C}_2\text{H}_5\text{SH}$, is soluble only to the extent of 1.5 g per 100 mL of water.
- Calculate the percent by mass of KBr in a saturated solution of KBr in water at 10°C . See Figure 14.2h for useful data, and report the computed percentage to one significant digit. **Check Answer:** ⁶
- Which of the following gases is expected to be most soluble in water? Explain your reasoning.
 - CH_4
 - CCl_4
 - CHCl_3
- 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? **Check Answer:** ⁷
- Refer to Figure 14.2c.
 - How did the concentration of dissolved CO_2 in the beverage change when the bottle was opened?
 - What caused this change?
 - Is the beverage unsaturated, saturated, or supersaturated with CO_2 ?
- The Henry's law constant for CO_2 is $3.4 \times 10^{-2} \text{ M/atm}$ 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? **Check Answer:** ⁸
- The Henry's law constant for O_2 is $1.3 \times 10^{-3} \text{ M/atm}$ 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?
- How many litres 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? **Check Answer:** ⁹

14.3 – Molarity

1. Explain what changes and what stays the same when 1.00 L of a solution of NaCl is diluted to 1.80 L.
2. What information do we need to calculate the molarity of a sulfuric acid solution?

Check Answer:¹⁰

3. 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?
4. 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, $\text{Ca}(\text{OH})_2$, in 40.00 mL of solution
 - d. 10.5 kg of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ 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

Check Answer:¹¹

5. 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 $\text{Al}(\text{NO}_3)_3$ in 1575 mL of solution
 - d. 2.76 kg of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ 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, $\text{C}_2\text{H}_5\text{NO}_2$, in 1.05 mL of solution
6. Consider this question: What is the mass of the solute in 0.500 L of 0.30 M glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, used for intravenous injection?
 - a. Outline the steps necessary to answer the question.
 - b. Answer the question.

Check Answer:¹²

7. 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.
8. 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

Check Answer: ¹³

9. Calculate the number of moles and the mass of the solute in each of the following solutions:
- 325 mL of $8.23 \times 10^{-5} M$ KI, a source of iodine in the diet
 - 75.0 mL of $2.2 \times 10^{-5} M$ H_2SO_4 , a sample of acid rain
 - 0.2500 L of $0.1135 M$ K_2CrO_4 , an analytical reagent used in iron assays
 - 10.5 L of $3.716 M$ $(NH_4)_2SO_4$, a liquid fertilizer
10. 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?

- Outline the steps necessary to answer the question.
- Answer the question.

Check Answer: ¹⁴

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

- Outline the steps necessary to answer the question.
- Answer the question.

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

- 0.195 g of cholesterol, $C_{27}H_{46}O$, in 0.100 L of serum, the average concentration of cholesterol in human serum
- 4.25 g of NH_3 in 0.500 L of solution, the concentration of NH_3 in household ammonia
- 1.49 kg of isopropyl alcohol, C_3H_7OH , in 2.50 L of solution, the concentration of isopropyl alcohol in rubbing alcohol
- 0.029 g of I_2 in 0.100 L of solution, the solubility of I_2 in water at $20^\circ C$

Check Answer: ¹⁵

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

- 293 g HCl in 666 mL of solution, a concentrated HCl solution
- 2.026 g $FeCl_3$ in 0.1250 L of a solution used as an unknown in general chemistry laboratories
- 0.001 mg Cd^{2+} in 0.100 L, the maximum permissible concentration of cadmium in drinking water
- 0.0079 g $C_7H_5SNO_3$ in one ounce (29.6 mL), the concentration of saccharin in a diet soft drink.

14. 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?

Check Answer: ¹⁶

15. 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 ?

16. 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?

Check Answer: ¹⁷

17. 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?

18. 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 ?

Check Answer:¹⁸

19. 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?
20. What is the molarity of the diluted solution when each of the following solutions is diluted to the given final volume?
- 1.00 L of a 0.250- M solution of $Fe(NO_3)_3$ is diluted to a final volume of 2.00 L
 - 0.5000 L of a 0.1222- M solution of C_3H_7OH is diluted to a final volume of 1.250 L
 - 2.35 L of a 0.350- M solution of H_3PO_4 is diluted to a final volume of 4.00 L
 - 22.50 mL of a 0.025- M solution of $C_{12}H_{22}O_{11}$ is diluted to 100.0 mL

Check Answer:¹⁹

21. 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?
22. 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? **Check Answer:**²⁰
23. 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?
24. What volume of a 0.20- M K_2SO_4 solution contains 57 g of K_2SO_4 ? **Check Answer:**²¹
25. 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?

14.4 – Other Units for Solution Concentrations

- 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?
 - Outline the steps necessary to answer the question.
 - Answer the question.

Check Answer:²²

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

Check Answer:²³

- 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?
- 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 litre 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 ?
Check Answer:²⁴
- 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.
- In Canada and the United Kingdom, devices that measure blood glucose levels provide a reading in millimoles per litre. If a measurement of 5.3 mM is observed, what is the concentration of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in mg/dL?
Check Answer:²⁵
- A throat spray is 1.40% by mass phenol, $\text{C}_6\text{H}_5\text{OH}$, in water. If the solution has a density of 0.9956 g/mL , calculate the molarity of the solution.
- 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?
Check Answer:²⁶
- A cough syrup contains 5.0% ethyl alcohol, $\text{C}_2\text{H}_5\text{OH}$, by mass. If the density of the solution is 0.9928 g/mL , determine the molarity of the alcohol in the cough syrup.
- D5W is a solution used as an intravenous fluid. It is a 5.0% by mass solution of dextrose ($\text{C}_6\text{H}_{12}\text{O}_6$) in water. If the density of D5W is 1.029 g/mL , calculate the molarity of dextrose in the solution.
Check Answer:²⁷
- 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 .

14.5 – Colligative Properties and Osmosis

- 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?
- What is the microscopic explanation for the macroscopic behaviour illustrated in Figure 14.2f in Chapter 14.2 Solubility?
Check Answer:²⁸
- Sketch a qualitative graph of the pressure versus time for water vapour above a sample of pure water and a sugar solution, as the liquids evaporate to half their original volume.
- A solution of potassium nitrate, an electrolyte, and a solution of glycerin ($\text{C}_3\text{H}_5(\text{OH})_3$), a

nonelectrolyte, both boil at 100.3 °C. What other physical properties of the two solutions are identical?

Check Answer:²⁹

5. What are the mole fractions of H_3PO_4 and water in a solution of 14.5 g of H_3PO_4 in 125 g of water?
 - a. Outline the steps necessary to answer the question.
 - b. Answer the question.
6. What are the mole fractions of HNO_3 and water in a concentrated solution of nitric acid (68.0% HNO_3 by mass)?
 - a. Outline the steps necessary to answer the question.
 - b. Answer the question.

Check Answer:³⁰

7. Calculate the mole fraction of each solute and solvent:
 - a. 583 g of H_2SO_4 in 1.50 kg of water—the acid solution used in an automobile battery
 - b. 0.86 g of NaCl in 1.00×10^2 g of water—a solution of sodium chloride for intravenous injection
 - c. 46.85 g of codeine, $\text{C}_{18}\text{H}_{21}\text{NO}_3$, in 125.5 g of ethanol, $\text{C}_2\text{H}_5\text{OH}$
 - d. 25 g of I_2 in 125 g of ethanol, $\text{C}_2\text{H}_5\text{OH}$
8. Calculate the mole fraction of each solute and solvent:
 - a. 0.710 kg of sodium carbonate (washing soda), Na_2CO_3 , in 10.0 kg of water—a saturated solution at 0 °C
 - b. 125 g of NH_4NO_3 in 275 g of water—a mixture used to make an instant ice pack
 - c. 25 g of Cl_2 in 125 g of dichloromethane, CH_2Cl_2
 - d. 0.372 g of histamine, $\text{C}_5\text{H}_9\text{N}$, in 125 g of chloroform, CHCl_3

Check Answer:³¹

9. Calculate the mole fractions of methanol, CH_3OH ; ethanol, $\text{C}_2\text{H}_5\text{OH}$; and water in a solution that is 40% methanol, 40% ethanol, and 20% water by mass. (Assume the data are good to two significant figures.)
10. What is the difference between a 1 *M* solution and a 1 *m* solution?

Check Answer:³²

11. What is the molality of phosphoric acid, H_3PO_4 , in a solution of 14.5 g of H_3PO_4 in 125 g of water?
 - a. Outline the steps necessary to answer the question.
 - b. Answer the question.
12. What is the molality of nitric acid in a concentrated solution of nitric acid (68.0% HNO_3 by mass)?
 - a. Outline the steps necessary to answer the question.
 - b. Answer the question.

Check Answer:³³

13. Calculate the molality of each of the following solutions:
 - a. 583 g of H_2SO_4 in 1.50 kg of water—the acid solution used in an automobile battery
 - b. 0.86 g of NaCl in 1.00×10^2 g of water—a solution of sodium chloride for intravenous injection

- c. 46.85 g of codeine, $C_{18}H_{21}NO_3$, in 125.5 g of ethanol, C_2H_5OH
- d. 25 g of I_2 in 125 g of ethanol, C_2H_5OH

14. Calculate the molality of each of the following solutions:

- a. 0.710 kg of sodium carbonate (washing soda), Na_2CO_3 , in 10.0 kg of water—a saturated solution at $0^\circ C$
- b. 125 g of NH_4NO_3 in 275 g of water—a mixture used to make an instant ice pack
- c. 25 g of Cl_2 in 125 g of dichloromethane, CH_2Cl_2
- d. 0.372 g of histamine, C_5H_9N , in 125 g of chloroform, $CHCl_3$

Check Answer:³⁴

15. The concentration of glucose, $C_6H_{12}O_6$, in normal spinal fluid is $\frac{75 \text{ mg}}{100 \text{ g}}$. What is the molality of the solution?

16. A 13.0% solution of K_2CO_3 by mass has a density of 1.09 g/cm^3 . Calculate the molality of the solution.

Check Answer:³⁵

17. Why does 1 mol of sodium chloride depress the freezing point of 1 kg of water almost twice as much as 1 mol of glycerin?

18. What is the boiling point of a solution of 115.0 g of sucrose, $C_{12}H_{22}O_{11}$, in 350.0 g of water?

- a. Outline the steps necessary to answer the question
- b. Answer the question

Check Answer:³⁶

19. What is the boiling point of a solution of 9.04 g of I_2 in 75.5 g of benzene, assuming the I_2 is nonvolatile?

- a. Outline the steps necessary to answer the question.
- b. Answer the question.

20. What is the freezing temperature of a solution of 115.0 g of sucrose, $C_{12}H_{22}O_{11}$, in 350.0 g of water, which freezes at $0.0^\circ C$ when pure?

- a. Outline the steps necessary to answer the question.
- b. Answer the question.

Check Answer:³⁷

21. What is the freezing point of a solution of 9.04 g of I_2 in 75.5 g of benzene?

- a. Outline the steps necessary to answer the following question.
- b. Answer the question.

22. What is the osmotic pressure of an aqueous solution of 1.64 g of $Ca(NO_3)_2$ in water at $25^\circ C$? The volume of the solution is 275 mL.

- a. Outline the steps necessary to answer the question.
- b. Answer the question.

Check Answer:³⁸

23. What is osmotic pressure of a solution of bovine insulin (molar mass, 5700 g mol^{-1}) at 18°C if 100.0 mL of the solution contains 0.103 g of the insulin?
- Outline the steps necessary to answer the question.
 - Answer the question.
24. What is the molar mass of a solution of 5.00 g of a compound in 25.00 g of carbon tetrachloride (bp 76.8°C ; $K_b = 5.02^\circ\text{C}/m$) that boils at 81.5°C at 1 atm?
- Outline the steps necessary to answer the question.
 - Solve the problem.
- Check Answer:** ³⁹
25. A sample of an organic compound (a nonelectrolyte) weighing 1.35 g lowered the freezing point of 10.0 g of benzene by 3.66°C . Calculate the molar mass of the compound.
26. A 1.0 *m* solution of HCl in benzene has a freezing point of 0.4°C . Is HCl an electrolyte in benzene? Explain.
- Check Answer:** ⁴⁰
27. A solution contains 5.00 g of urea, $\text{CO}(\text{NH}_2)_2$, a nonvolatile compound, dissolved in 0.100 kg of water. If the vapour pressure of pure water at 25°C is 23.7 torr, what is the vapour pressure of the solution?
28. A 12.0-g sample of a nonelectrolyte is dissolved in 80.0 g of water. The solution freezes at -1.94°C . Calculate the molar mass of the substance.
- Check Answer:** ⁴¹
29. Arrange the following solutions in order by their decreasing freezing points: 0.1 *m* Na_3PO_4 , 0.1 *m* $\text{C}_2\text{H}_5\text{OH}$, 0.01 *m* CO_2 , 0.15 *m* NaCl , and 0.2 *m* CaCl_2 .
30. Calculate the boiling point elevation of 0.100 kg of water containing 0.010 mol of NaCl , 0.020 mol of Na_2SO_4 , and 0.030 mol of MgCl_2 , assuming complete dissociation of these electrolytes.
- Check Answer:** ⁴²
31. How could you prepare a 3.08 *m* aqueous solution of glycerin, $\text{C}_3\text{H}_8\text{O}_3$? What is the freezing point of this solution?
32. A sample of sulfur weighing 0.210 g was dissolved in 17.8 g of carbon disulfide, CS_2 ($K_b = 2.43^\circ\text{C}/m$). If the boiling point elevation was 0.107°C , what is the formula of a sulfur molecule in carbon disulfide?
- Check Answer:** ⁴³
33. In a significant experiment performed many years ago, 5.6977 g of cadmium iodide in 44.69 g of water raised the boiling point 0.181°C . What does this suggest about the nature of a solution of CdI_2 ?
34. Lysozyme is an enzyme that cleaves cell walls. A 0.100-L sample of a solution of lysozyme that contains 0.0750 g of the enzyme exhibits an osmotic pressure of 1.32×10^{-3} atm at 25°C . What is the molar mass of lysozyme? **Check Answer:** ⁴⁴
35. The osmotic pressure of a solution containing 7.0 g of insulin per litre is 23 torr at 25°C . What is the molar mass of insulin?
36. The osmotic pressure of human blood is 7.6 atm at 37°C . What mass of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, is required

to make 1.00 L of aqueous solution for intravenous feeding if the solution must have the same osmotic pressure as blood at body temperature, 37 °C?

Check Answer: ⁴⁵

37. What is the freezing point of a solution of dibromobenzene, $C_6H_4Br_2$, in 0.250 kg of benzene, if the solution boils at 83.5 °C?

38. What is the boiling point of a solution of NaCl in water if the solution freezes at -0.93 °C?

Check Answer: ⁴⁶

39. The sugar fructose contains 40.0% C, 6.7% H, and 53.3% O by mass. A solution of 11.7 g of fructose in 325 g of ethanol has a boiling point of 78.59 °C. The boiling point of ethanol is 78.35 °C, and K_b for ethanol is 1.20 °C/ m . What is the molecular formula of fructose?

40. The vapour pressure of methanol, CH_3OH , is 94 torr at 20 °C. The vapour pressure of ethanol, C_2H_5OH , is 44 torr at the same temperature.

- Calculate the mole fraction of methanol and of ethanol in a solution of 50.0 g of methanol and 50.0 g of ethanol.
- Ethanol and methanol form a solution that behaves like an ideal solution. Calculate the vapour pressure of methanol and of ethanol above the solution at 20 °C.
- Calculate the mole fraction of methanol and of ethanol in the vapour above the solution.

Check Answer: ⁴⁷

41. The triple point of air-free water is defined as 273.15 K. Why is it important that the water be free of air?

42. Meat can be classified as fresh (not frozen) even though it is stored at -1 °C. Why wouldn't meat freeze at this temperature?

Check Answer: ⁴⁸

43. An organic compound has a composition of 93.46% C and 6.54% H by mass. A solution of 0.090 g of this compound in 1.10 g of camphor melts at 158.4 °C. The melting point of pure camphor is 178.4 °C. K_f for camphor is 37.7 °C/ m . What is the molecular formula of the solute? Show your calculations.

44. A sample of $HgCl_2$ weighing 9.41 g is dissolved in 32.75 g of ethanol, C_2H_5OH ($K_b = 1.20$ °C/ m). The boiling point elevation of the solution is 1.27 °C. Is $HgCl_2$ an electrolyte in ethanol? Show your calculations. **Check Answer:** ⁴⁹

45. A salt is known to be an alkali metal fluoride. A quick approximate determination of freezing point indicates that 4 g of the salt dissolved in 100 g of water produces a solution that freezes at about -1.4 °C. What is the formula of the salt? Show your calculations.

14.6 – Colloids

- 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.

Check Answer: ⁵⁰

2. Distinguish between dispersion methods and condensation methods for preparing colloidal systems.
3. How do colloids differ from solutions with regard to dispersed particle size and homogeneity?

Check Answer: ⁵¹

4. Explain the cleansing action of soap.
5. How can it be demonstrated that colloidal particles are electrically charged?

Check Answer: ⁵²

Attribution & References

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Notes

1. A solution can vary in composition, while a compound cannot vary in composition. Solutions are homogeneous at the molecular level, while other mixtures are heterogeneous.
2. (a) The process is endothermic as the solution is consuming heat. (b) Attraction between the K^+ and NO_3^- ions is stronger than between the ions and water molecules (the ion-ion interactions have a lower, more negative energy). Therefore, the dissolution process increases the energy of the molecular interactions, and it consumes the thermal energy of the solution to make up for the difference. (c) No, an ideal solution is formed with no appreciable heat release or consumption.
3. (a) ion-dipole forces; (b) dipole-dipole forces; (c) dispersion forces; (d) dispersion forces; (e) hydrogen bonding
4. Heat is released when the total intermolecular forces (IMFs) between the solute and solvent molecules are stronger than the total IMFs in the pure solute and in the pure solvent: Breaking weaker IMFs and forming stronger IMFs releases heat. Heat is absorbed when the total IMFs in the solution are weaker than the total of those in the pure solute and in the pure solvent: Breaking stronger IMFs and forming weaker IMFs absorbs heat.
5. The solubility of solids usually decreases upon cooling a solution, while the solubility of gases usually decreases upon heating.
6. 40%
7. 2.80 g

8. 2.9 atm
9. 102 L HCl
10. We need to know the number of moles of sulfuric acid dissolved in the solution and the volume of the solution.
11. (a) 0.679 *M*; (b) 1.00 *M*; (c) 0.06998 *M*; (d) 1.75 *M*; (e) 0.070 *M*; (f) 6.6 *M*
12. (a) determine the number of moles of glucose in 0.500 L of solution; determine the molar mass of glucose; determine the mass of glucose from the number of moles and its molar mass; (b) 27 g
13. (a) 37.0 mol H₂SO₄; 3.63 × 10³ g H₂SO₄; (b) 3.8 × 10⁻⁶ mol NaCN; 1.9 × 10⁻⁴ g NaCN; (c) 73.2 mol H₂CO; 2.20 kg H₂CO; (d) 5.9 × 10⁻⁷ mol FeSO₄; 8.9 × 10⁻⁵ g FeSO₄
14. (a) Determine the molar mass of KMnO₄; determine the number of moles of KMnO₄ in the solution; from the number of moles and the volume of solution, determine the molarity; (b) 1.15 × 10⁻³ *M*
15. (a) 5.04 × 10⁻³ *M*; (b) 0.499 *M*; (c) 9.92 *M*; (d) 1.1 × 10⁻³ *M*
16. 0.025 *M*
17. 0.5000 L
18. 1.9 mL
19. (a) 0.125 *M*; (b) 0.04888 *M*; (c) 0.206 *M*; (e) 0.0056 *M*
20. 11.9 *M*
21. 1.6 L
22. (a) The dilution equation can be used, appropriately modified to accommodate mass-based concentration units: $\% \text{mass}_1 \times \text{mass}_1 = \% \text{mass}_2 \times \text{mass}_2$ This equation can be rearranged to isolate mass₁ and the given quantities substituted into this equation. (b) 58.8 g
23. 114 g
24. 1.75 × 10⁻³ *M*
25. 95 mg/dL
26. 2.38 × 10⁻⁴ mol
27. 0.29 mol
28. The strength of the bonds between like molecules is stronger than the strength between unlike molecules. Therefore, some regions will exist in which the water molecules will exclude oil molecules and other regions will exist in which oil molecules will exclude water molecules, forming a heterogeneous region.
29. Both form homogeneous solutions; their boiling point elevations are the same, as are their lowering of vapour pressures. Osmotic pressure and the lowering of the freezing point are also the same for both solutions.
30. (a) Find number of moles of HNO₃ and H₂O in 100 g of the solution. Find the mole fractions for the components. (b) The mole fraction of HNO₃ is 0.378. The mole fraction of H₂O is 0.622.
31. (a) $X_{\text{Na}_2\text{CO}_3} = 0.0119$; $X_{\text{H}_2\text{O}} = 0.988$; (b) $X_{\text{NH}_4\text{NO}_3} = 0.9927$; $X_{\text{H}_2\text{O}} = 0.907$; (c) $X_{\text{Cl}_2} = 0.192$; $X_{\text{CH}_2\text{Cl}_2} = 0.808$; (d) $X_{\text{C}_5\text{H}_9\text{N}} = 0.00426$; $X_{\text{CHCl}_3} = 0.997$
32. In a 1 *M* solution, the mole is contained in exactly 1 L of solution. In a 1 *m* solution, the mole is contained in exactly 1 kg of solvent.
33. (a) Determine the molar mass of HNO₃. Determine the number of moles of acid in the solution. From the number of moles and the mass of solvent, determine the molality. (b) 33.7 *m*
34. (a) 6.70 × 10⁻¹ *m*; (b) 5.67 *m*; (c) 2.8 *m*; (d) 0.0358 *m*
35. 1.08 *m*
36. (a) Determine the molar mass of sucrose; determine the number of moles of sucrose in the solution; convert the mass of

- solvent to units of kilograms; from the number of moles and the mass of solvent, determine the molality; determine the difference between the boiling point of water and the boiling point of the solution; determine the new boiling point. (b) 100.5 °C
37. (a) Determine the molar mass of sucrose; determine the number of moles of sucrose in the solution; convert the mass of solvent to units of kilograms; from the number of moles and the mass of solvent, determine the molality; determine the difference between the freezing temperature of water and the freezing temperature of the solution; determine the new freezing temperature. (b) -1.8 °C
38. (a) Determine the molar mass of $\text{Ca}(\text{NO}_3)_2$; determine the number of moles of $\text{Ca}(\text{NO}_3)_2$ in the solution; determine the number of moles of ions in the solution; determine the molarity of ions, then the osmotic pressure. (b) 2.67 atm
39. (a) Determine the molal concentration from the change in boiling point and K_b ; determine the moles of solute in the solution from the molal concentration and mass of solvent; determine the molar mass from the number of moles and the mass of solute. (b) $2.1 \times 10^2 \text{ g mol}^{-1}$
40. No. Pure benzene freezes at 5.5 °C, and so the observed freezing point of this solution is depressed by $\Delta T_f = 5.5 - 0.4 = 5.1$ °C. The value computed, assuming no ionization of HCl, is $\Delta T_f = (1.0 \text{ m})(5.14 \text{ °C}/m) = 5.1$ °C. Agreement of these values supports the assumption that HCl is not ionized.
41. 144 g mol^{-1}
42. 0.870 °C
43. S_8
44. $1.39 \times 10^4 \text{ g mol}^{-1}$
45. 54 g
46. 100.26 °C
47. (a) $X_{\text{CH}_3\text{OH}} = 0.590$; $X_{\text{C}_2\text{H}_5\text{OH}} = 0.410$; (b) Vapor pressures are: CH_3OH : 55 torr; $\text{C}_2\text{H}_5\text{OH}$: 18 torr; (c) CH_3OH : 0.75; $\text{C}_2\text{H}_5\text{OH}$: 0.25
48. The ions and compounds present in the water in the beef lower the freezing point of the beef below -1 °C.
49. $\Delta bp = K_b m = (1.20 \text{ °C}/m) \left(\frac{9.41 \text{ g} \times \frac{1 \text{ mol HgCl}_2}{271.496 \text{ g}}}{0.03275 \text{ kg}} \right) = 1.27 \text{ °C}$ The observed change equals the theoretical change; therefore, no dissociation occurs.

50.	Colloidal System	Dispersed Phase	Dispersion Medium
	starch dispersion	starch	water
	smoke	solid particles	air
	fog	water	air
	pearl	water	calcium carbonate (CaCO_3)
	whipped cream	air	cream
	floating soap	air	soap
	jelly	fruit juice	pectin gel
	milk	butterfat	water
	ruby	chromium(III) oxide (Cr_2O_3)	aluminum oxide (Al_2O_3)

51. Colloidal dispersions consist of particles that are much bigger than the solutes of typical solutions. Colloidal particles are either very large molecules or aggregates of smaller species that usually are big enough to scatter light. Colloids are homogeneous on a macroscopic (visual) scale, while solutions are homogeneous on a microscopic (molecular) scale.

52. If they are placed in an electrolytic cell, dispersed particles will move toward the electrode that carries a charge opposite to their own charge. At this electrode, the charged particles will be neutralized and will coagulate as a precipitate.