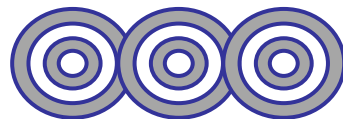




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Chapter 13

Properties of Solutions

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13.1

The Solution Process

Solutions

- **Solutions** are homogeneous mixtures of two or more pure substances.
- In a solution, the **solute** is dispersed uniformly throughout the **solvent**.
- Solutions may be solids, liquids or gases.

State of Solution	State of Solvent	State of Solute	Example
Gas	Gas	Gas	Air
Liquid	Liquid	Gas	Oxygen in water
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Liquid	Solid	Salt in water
Solid	Solid	Gas	Hydrogen in palladium
Solid	Solid	Liquid	Mercury in silver
Solid	Solid	Solid	Silver in gold

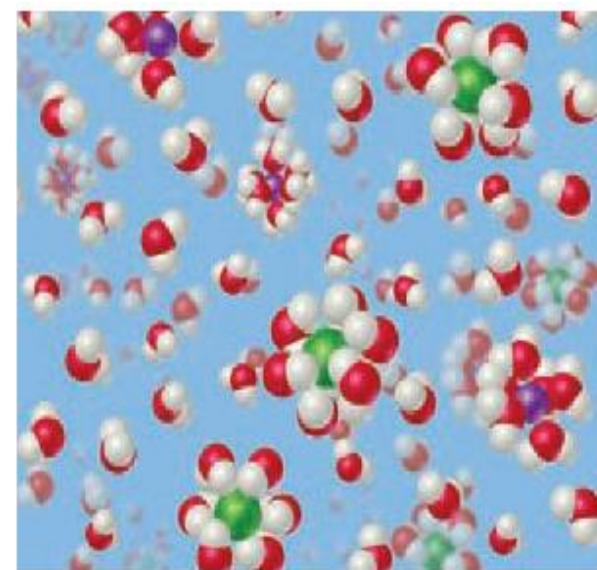
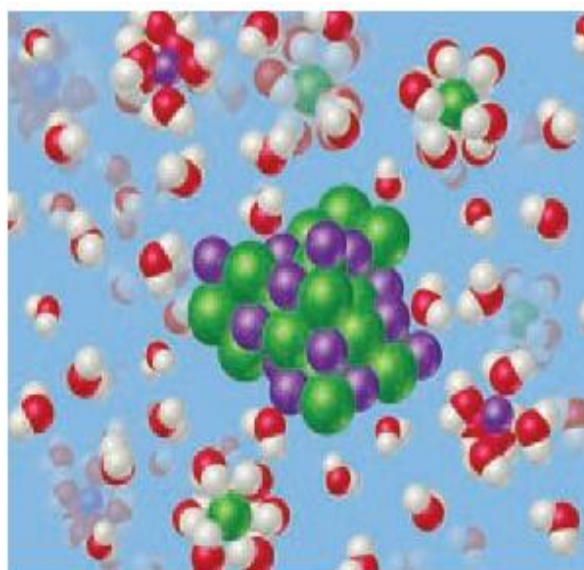
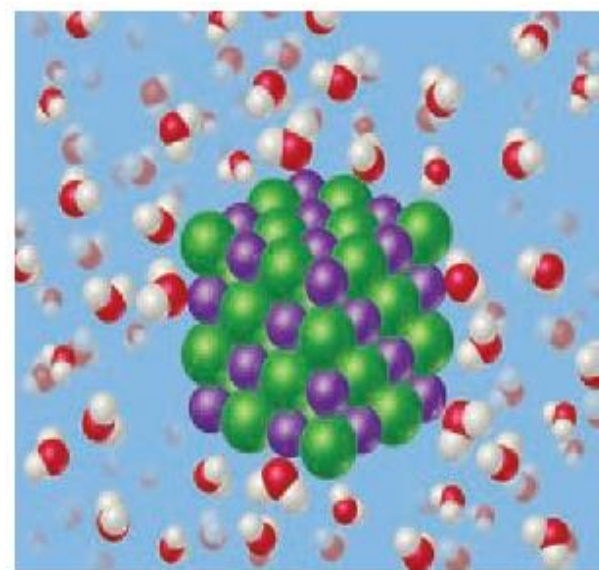
The ability of substances to form solutions depends on two general factors:

- (1) The types of intermolecular interactions involved in the solution process.
- (2) The natural tendency of substances to spread into larger volumes when not restrained in some way (like gases).

The Effect of Intermolecular Forces

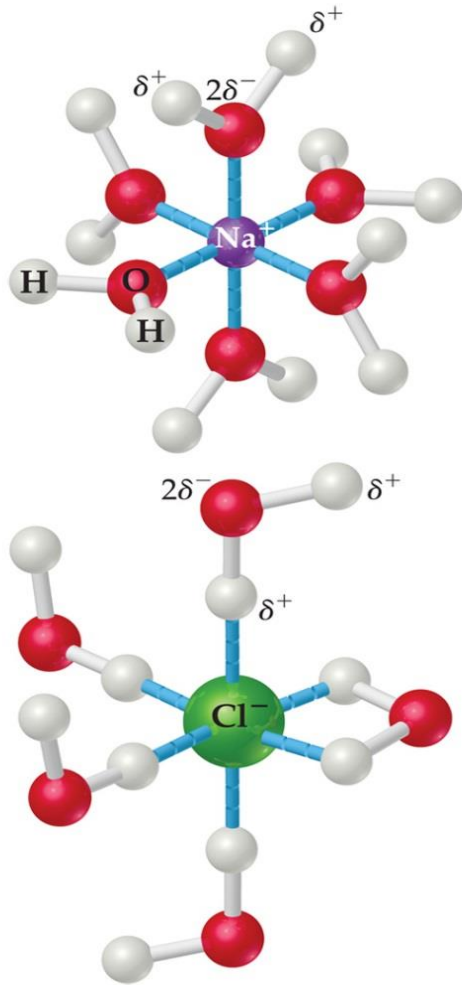
The extent to which one substance is able to dissolve in another depends on the relative magnitudes of the **solute-solvent**, **solute-solute**, and **solvent-solvent** interactions involved in the solution process.





The intermolecular forces between solute and solvent particles must be strong enough to compete with those between solute particles and those between solvent particles.

How Does a Solution Form?



When **NaCl** is added to **H₂O**, **H₂O** molecules orient themselves on the surface of the **NaCl** crystals.

The +ve end of **H₂O** dipole is oriented toward the **Cl⁻** ions, and the -ve end of the **H₂O** dipole is oriented toward the **Na⁺** ions.

The ion-dipole attractions between the ions and **H₂O** molecules are strong enough to pull the ions from their positions in the crystal.

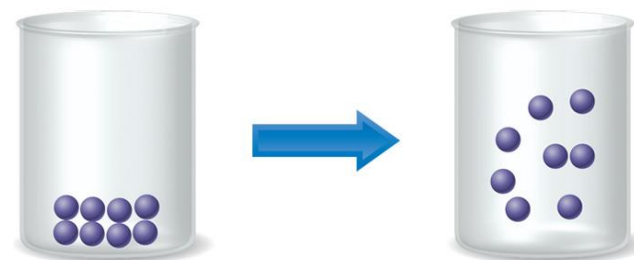
Once separated from the crystal, the **Na⁺** and **Cl⁻** ion are surrounded by **H₂O** molecules.

This type of interaction between solute and solvent molecules are known as **solvation**. When the solvent is water, the interaction are called **hydration**.

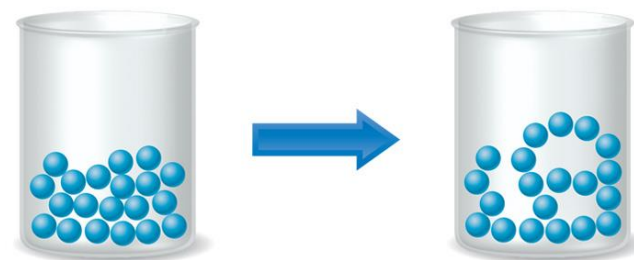
Energy Changes in Solution

Three processes affect the energetics of solution:

- separation of solute particles (breaking solute-solute interactions),
- separation of solvent particles (breaking solvent-solvent interactions),
- new interactions between solute and solvent (forming the solute-solvent interactions).



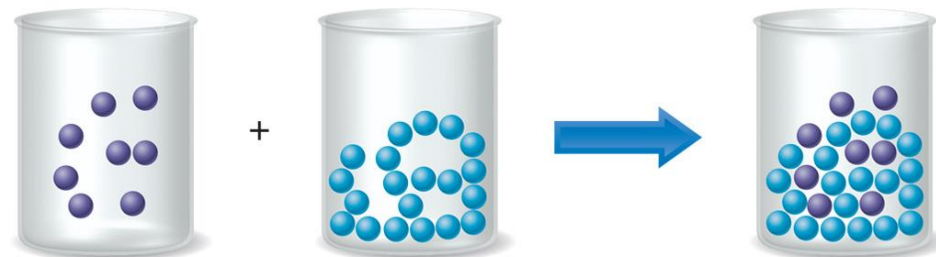
ΔH_1 : Separation of solute molecules



ΔH_2 : Separation of solvent molecules

The overall enthalpy change in forming a solution, ΔH_{soln} , is the sum of the three associated processes:

$$\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$



ΔH_3 : Formation of solute-solvent interactions

- Regardless of the particular solute being considered, separation of the solute particles from one another requires an input of energy to overcome their attractive interactions. The process is therefore endothermic ($\Delta H_1 > 0$).
- Separation of solvent molecules to accommodate the solute also always requires energy ($\Delta H_2 > 0$).
- The third component, which arises from the attractive interactions between solute and solvent, is always exothermic ($\Delta H_3 < 0$).

Depending on the relative magnitudes of the terms, the sum of the three components are either –ve or +ve value. Thus, the formation of a solution can be either **exothermic** or **endothermic**.

The solvent-solute interaction must be strong enough to make ΔH_3 comparable in magnitude to $\Delta H_1 + \Delta H_2$

Examples:

- Ionic substances such as **NaCl** dissolves readily in water.
- Nonpolar hydrocarbons such as **C₆H₁₄** dissolves in another nonpolar one like **CCl₄**.
- **NaCl** do not dissolve in nonpolar liquids such as gasoline.
- A polar liquid such as water does not form solutions with a nonpolar liquid such as octane (**C₈H₁₈**).

For example, when magnesium sulfate, MgSO_4 , is added to water, the resultant solution gets quite warm: $\Delta H_{\text{soln}} = -91.2 \text{ kJ/mol}$ (for hot packs).

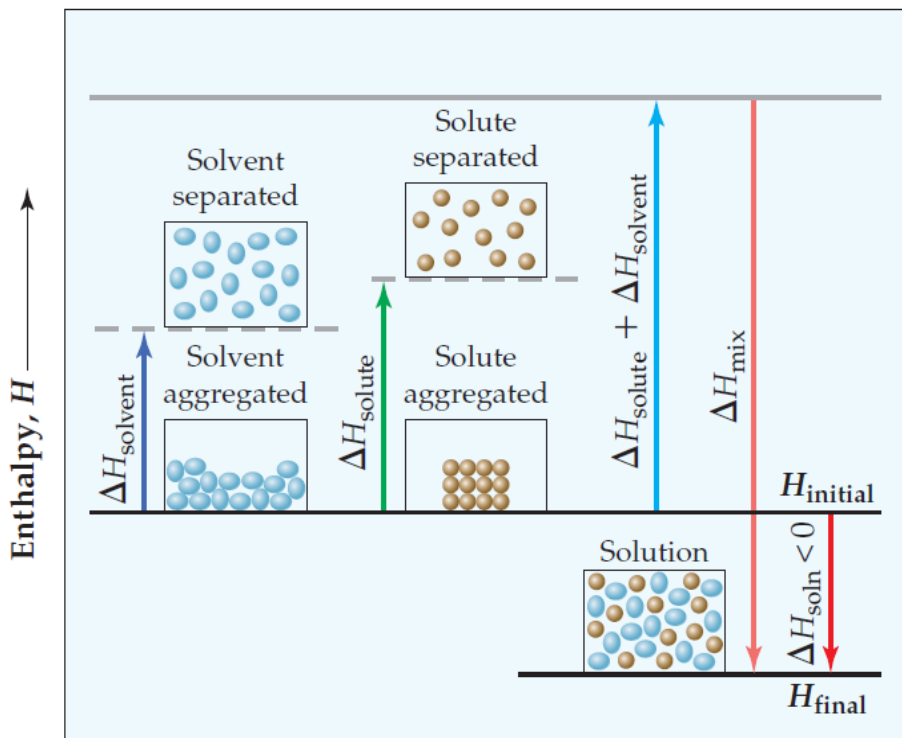
In contrast, the dissolution of ammonium nitrate (NH_4NO_3) is endothermic $\Delta H_{\text{soln}} = 26.4 \text{ kJ/mol}$ (for cold packs).

These particular substances have been used to make the instant heat packs and ice packs that are used to treat athletic injuries.



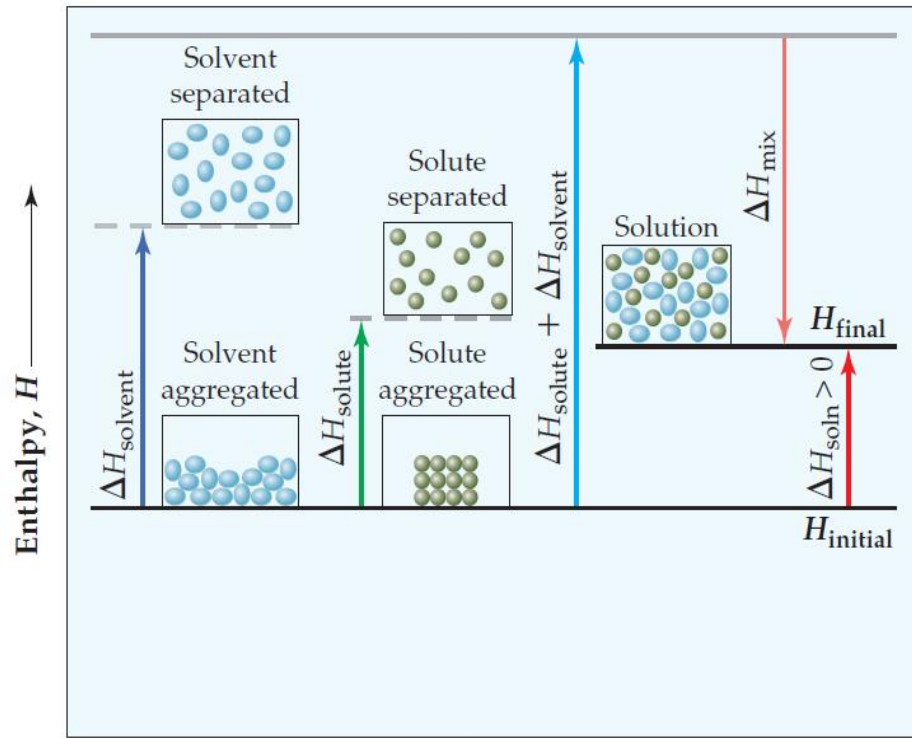
Enthalpy Changes in Solution

Depending on the relative magnitudes of the terms, the sum of the three components are either -ve or +ve value (exothermic or endothermic solution).



Exothermic solution process

a net exothermic process ($\Delta H_{\text{soln}} < 0$)



Endothermic solution process

a net endothermic process ($\Delta H_{\text{soln}} > 0$)

Solution Formation, Spontaneity and Entropy

Enthalpy is only part of the picture

The other factor is the distribution of each component into a larger volume – the tendency in nature for substances to mix and spread out into larger volumes.

Increasing the disorder or randomness (**entropy**) of a system tends to lower the energy of the system. Processes in which the energy content of the system decreases tend to occur spontaneously.

So even though enthalpy may increase, the overall energy of the system can still decrease if the system becomes more disordered.

Most spontaneous processes tend to be exothermic.



(a)

Formation of a homogeneous solution has increased the degree of dispersal, or randomness, because the molecules of each substance are now mixed and distributed in a volume twice as large as that which they occupied individually before mixing.



(b)

The degree of randomness in the system, sometimes referred to as disorder, is given by a thermodynamic quantity called **entropy**.

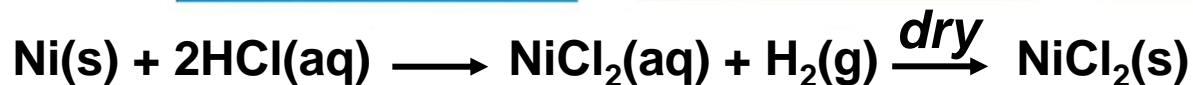
Thus, processes occurring at a constant temperature in which the randomness or dispersal in space (entropy) of the system increases tend to occur spontaneously.

Solution Formation and Chemical Reactions

We must be careful to distinguish the physical process of solution formation from chemical reactions that lead to a solution.

e.g., nickel metal is dissolved on contact with **HCl** solution because the following reaction occurs:

**Dissolution
vs.
Reaction**

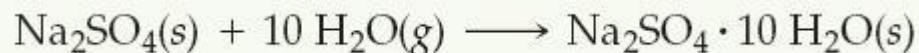


Dissolution is a physical change - you can get back the original solute by evaporating the solvent.

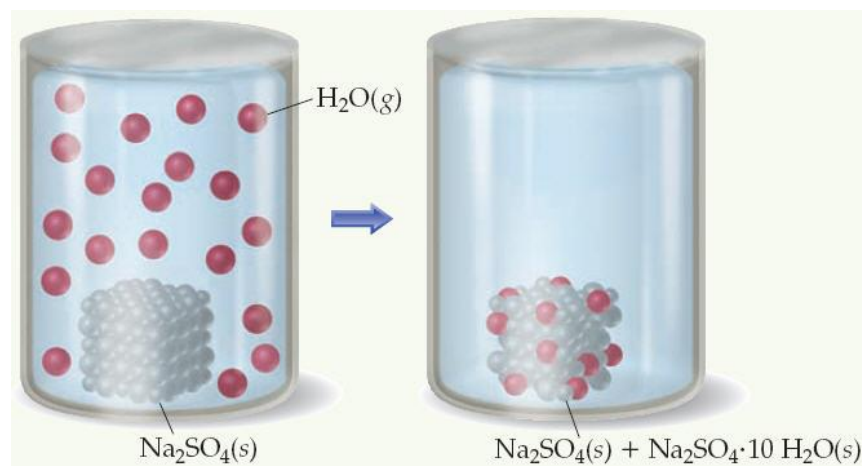
If you can't, the substance didn't dissolve, it reacted.

Sample Exercise Assessing Entropy Change

In the process illustrated below, water vapor reacts with excess solid sodium sulfate to form the hydrated form of the salt. The chemical reaction is



Essentially all of the water vapor in the closed container is consumed in this reaction. If we consider our system to consist initially of $\text{Na}_2\text{SO}_4(\text{s})$ and $10 \text{H}_2\text{O}(\text{g})$, **(a)** does the system become more or less ordered in this process, and **(b)** does the entropy of the system increase or decrease?



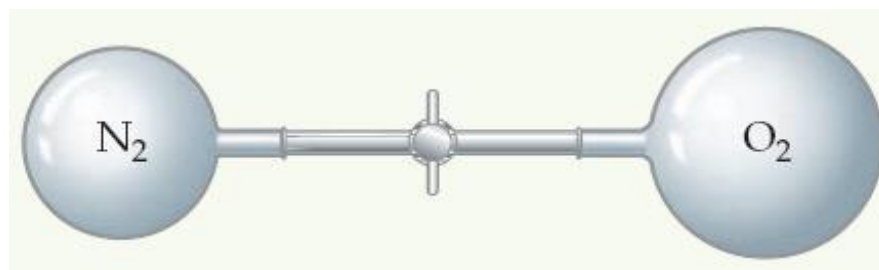
Solution

(a) In the course of forming the hydrate of $\text{Na}_2\text{SO}_4(\text{s})$, the water vapor moves from the vapor state, in which it is dispersed throughout the entire volume of the container, to the solid state, where it is confined to the lattice. This means that the water vapor becomes less dispersed (more ordered, or less random).

(b) When a system becomes less dispersed, or more ordered, its entropy is decreased.

Practice Exercise

Does the entropy of the system increase or decrease when the stopcock is opened to allow mixing of the two gases in this apparatus?



Answer: The entropy increases because each gas eventually becomes dispersed in twice the volume it originally occupied.

13.3

Factors Affecting Solubility

Factors Affecting Solubility

The extent to which one substance dissolves in another depends on:

- **Solute – Solvent Interactions**
nature of both solute and solvent.
- **Pressure Effects**
at least for gases.
- **Temperature Effects.**

Solute – Solvent Interactions

Chemists use the axiom “**like dissolves like**”.

Polar substances tend to dissolve in polar solvents.

Nonpolar substances tend to dissolve in nonpolar solvents.

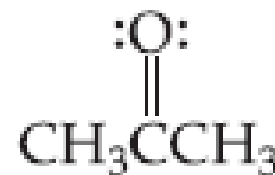
The more similar the intermolecular attractions, the more likely one substance is to be soluble in another.

The stronger the intermolecular attractions between solute and solvent, the more likely the solute will dissolve.

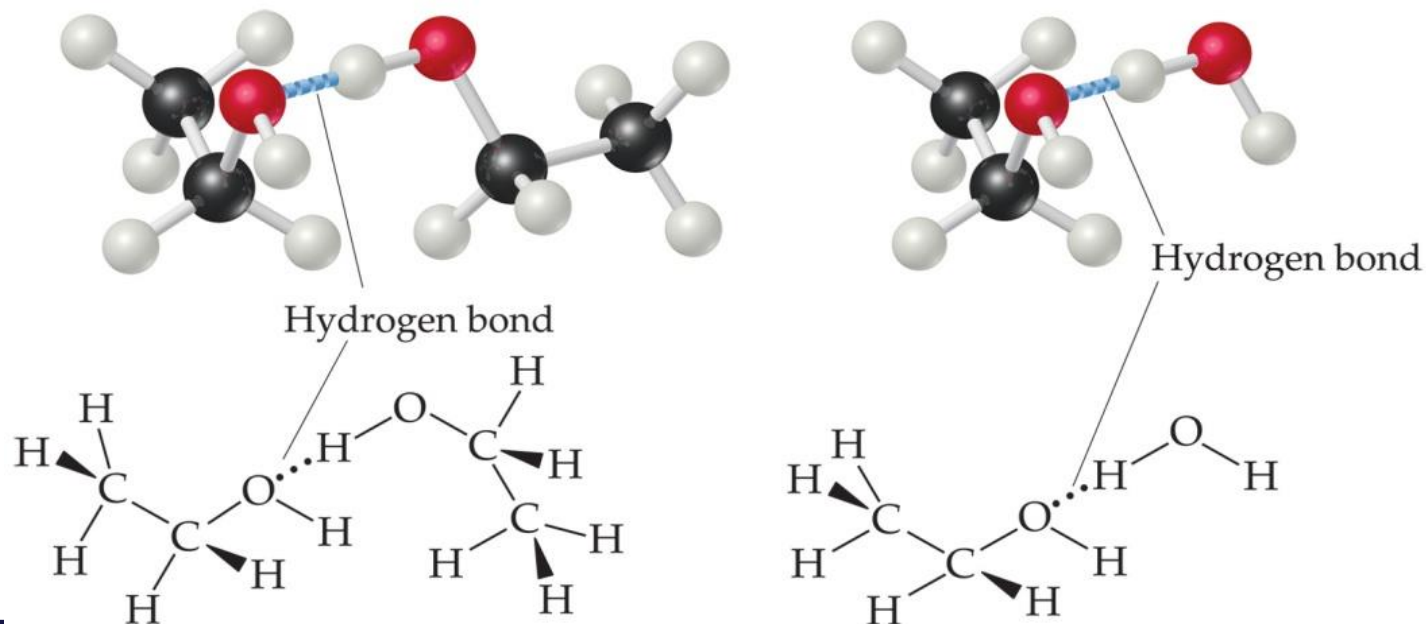
Polar liquids tend to dissolve readily in polar solvents because of favorable dipole-dipole attractions between solvent molecules and solute molecules.

Water (H_2O) is both polar and able to form hydrogen bonds. Thus, polar molecules, especially those that can form hydrogen bonds with water molecules, tend to be soluble in water.

For example, acetone has a strongly polar $\text{C}=\text{O}$ bond and pairs of nonbonding electrons on the O atom that can form hydrogen bonds with water.



Example
ethanol in water



Nonpolar liquids tend to be insoluble in polar liquids

Gasoline, which is a mixture of hydrocarbons, is immiscible with water.

Hydrocarbons are nonpolar substances because of several factors: the **C-C** bonds are nonpolar, the **C-H** bonds are nearly nonpolar, and the shapes of the molecules are symmetrical enough to cancel much of the weak **C-H** bond dipoles.

The attraction between the polar water molecules and the nonpolar hydrocarbon molecules is not sufficiently strong to allow the formation of a solution.



Hexane is insoluble in water (hexane is in the top layer because it is less dense than water).

Pairs of liquids that mix in all proportions are **miscible**, whereas those that do not dissolve in one another are **immiscible**.

TABLE 13.3 ■ Solubilities of Some Alcohols in Water and in Hexane*

Alcohol	Solubility in H ₂ O	Solubility in C ₆ H ₁₄
CH ₃ OH (methanol)	∞	0.12
CH ₃ CH ₂ OH (ethanol)	∞	∞
CH ₃ CH ₂ CH ₂ OH (propanol)	∞	∞
CH ₃ CH ₂ CH ₂ CH ₂ OH (butanol)	0.11	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ OH (pentanol)	0.030	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ OH (hexanol)	0.0058	∞

*Expressed in mol alcohol/100 g solvent at 20 °C. The infinity symbol (∞) indicates that the alcohol is completely miscible with the solvent.

The number of carbon atoms in an alcohol affects its solubility in water.

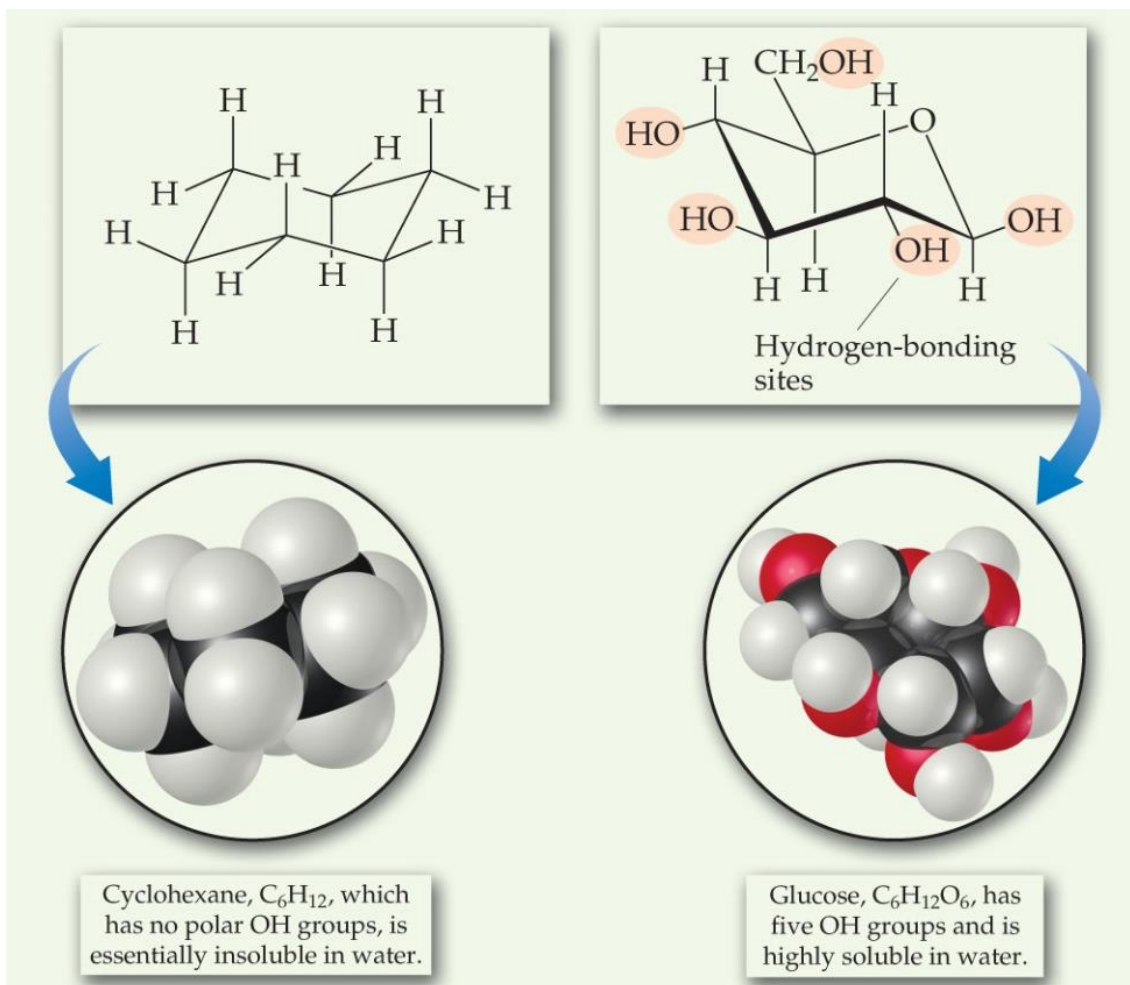
- As the length of the carbon chain increases, the polar **OH** becomes smaller and the molecule behaves more like a hydrocarbon. The solubility of the alcohol in water decreases.

- The solubility of the alcohols in a nonpolar solvent like hexane (**C₆H₁₄**) increases as the nonpolar hydrocarbon chain increases in length.

One way to enhance the solubility of a substance in water is to increase the number of polar groups it contains.

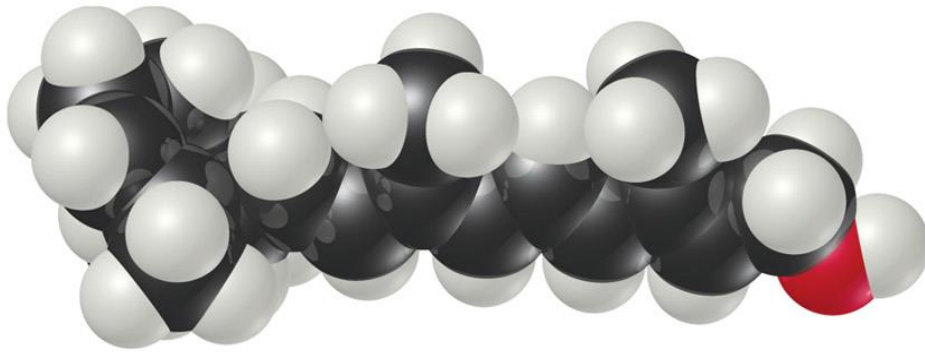
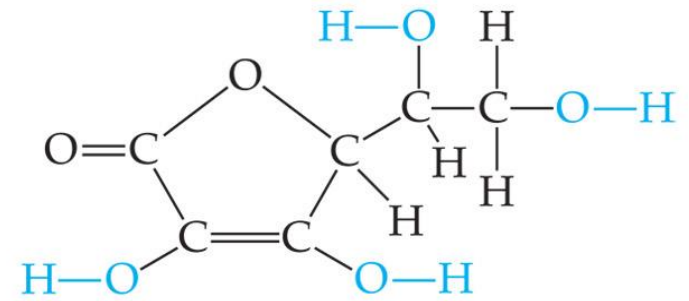
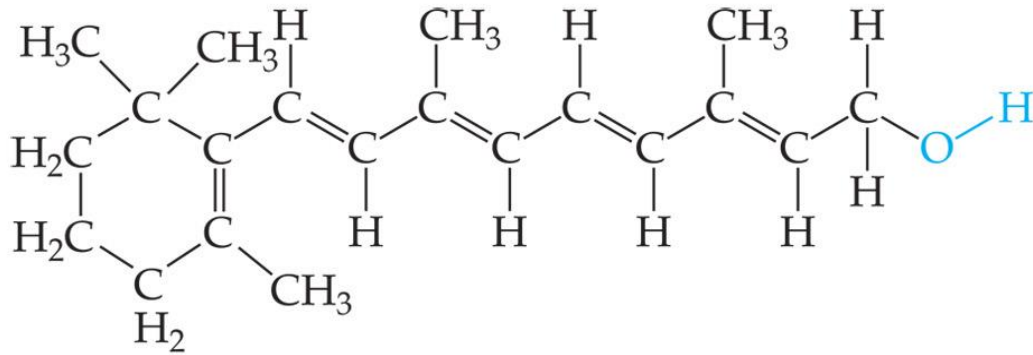
For example, increasing the number of **OH** groups along a carbon chain of a solute increases the extent of hydrogen bonding between that solute and water, thereby increasing solubility.

Glucose (which has hydrogen bonding) is very soluble in water, while **cyclohexane** (which only has dispersion forces) is not.

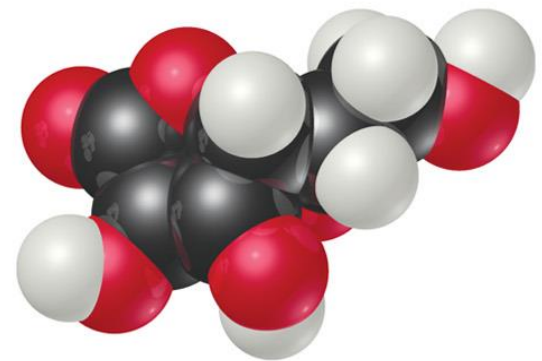


Vitamin A is soluble in nonpolar compounds (fat soluble vitamin).

Vitamin C is soluble in water (water soluble vitamin).



Vitamin A



Vitamin C

Sample Exercise Predicting Solubility Patterns

Predict whether each of the following substances is more likely to dissolve in the nonpolar solvent carbon tetrachloride (CCl_4) or in water: C_7H_{16} , Na_2SO_4 , HCl , and I_2 .

Solution

- C_7H_{16} : hydrocarbon, so it is molecular and nonpolar.
- Na_2SO_4 : compound containing a metal and nonmetals, is ionic.
- HCl : containing two nonmetals that differ in electronegativity, is polar.
- I_2 : diatomic molecule with atoms of equal electronegativity, is nonpolar.

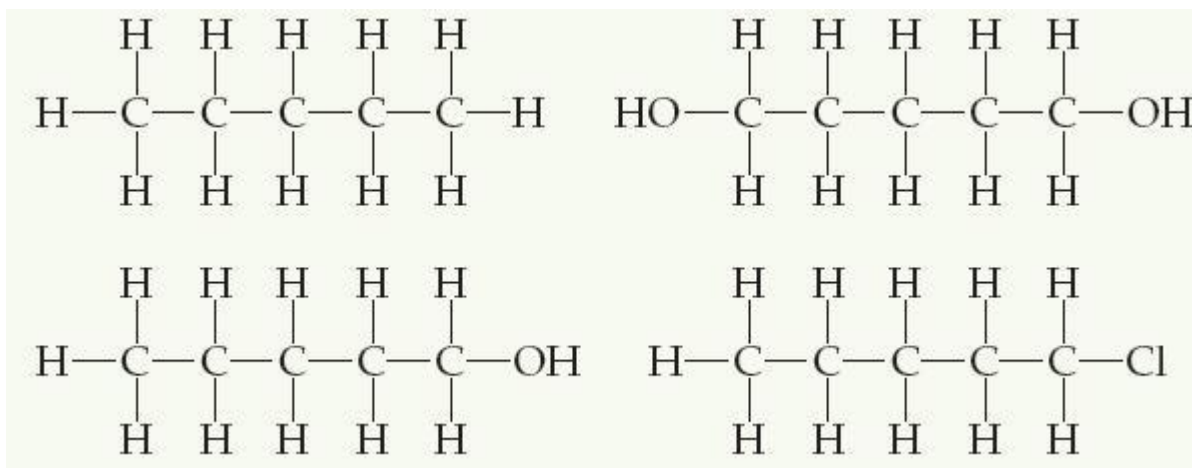
Therefore,

C_7H_{16} and I_2 (the nonpolar solutes) would be more soluble in the nonpolar CCl_4 than in polar H_2O ,

Whereas water would be the better solvent for Na_2SO_4 and HCl (the ionic and polar covalent solutes).

Practice Exercise

Arrange the following substances in order of increasing solubility in water:



Answer: $\text{C}_5\text{H}_{12} < \text{C}_5\text{H}_{11}\text{Cl} < \text{C}_5\text{H}_{11}\text{OH} < \text{C}_5\text{H}_{10}(\text{OH})_2$ (in order of increasing polarity and hydrogen-bonding ability)

Pressure Effects

The solubility of solids and liquids are not appreciably affected by pressure.

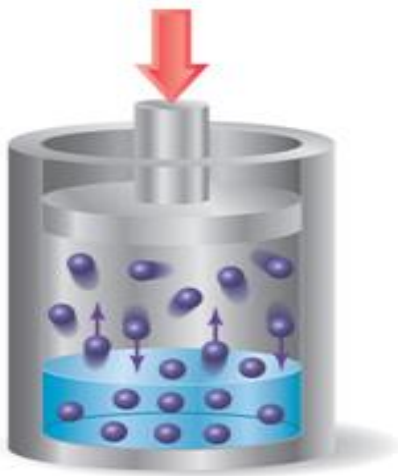
The solubility of a gas in any solvent is increased as the pressure over the solvent increases.

TABLE 13.2 ■ Solubilities of Gases in Water at 20 °C, with 1 atm Gas Pressure

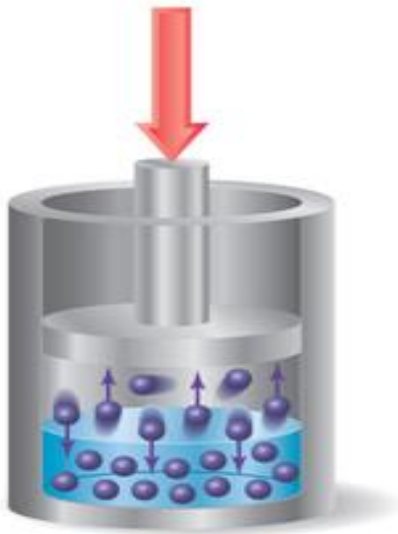
Gas	Solubility (<i>M</i>)
N ₂	0.69×10^{-3}
CO	1.04×10^{-3}
O ₂	1.38×10^{-3}
Ar	1.50×10^{-3}
Kr	2.79×10^{-3}

The attractive forces between the gas and solvent molecules are mainly of the **London dispersion** type, which increase with increasing size and mass of the gas molecules.

The solubility of gases in water increase as the attraction between the solute (gas) and solvent (water) increases.



The solubility of a gas in a liquid is directly proportional to its pressure (increasing pressure above solution forces more gas to dissolve).



As a result, the solubility of the gas in the solution would increase until equilibrium is again established, that is, solubility increases until the rate at which gas molecules enter the solution equals the rate at which solute molecules escape from the solution.

When the pressure is increased, the rate at which gas molecules enter the solution increases. The concentration of solute molecules at equilibrium increases in proportion to the pressure.

Henry's Law

The relationship between pressure and the solubility of a gas is expressed by a simple equation known as **Henry's law**:

$$S_g = kP_g$$

Where:

S_g is the solubility of the gas (usually expressed as molarity).

k is a proportionality constant known as the Henry's Law constant for that gas in that solvent, its different for each solute-solvent pair. It also varies with temperature.

P_g is the partial pressure of the gas over the solution.

Example, the solubility of N_2 gas in water at 25 °C and 0.78 atm pressure is 5.3×10^{-4} M. The Henry's law constant for N_2 in water at 25 °C is thus given by

$$5.3 \times 10^{-4} \text{ mol/L} / 0.78 \text{ atm} = 6.8 \times 10^{-4} \text{ mol/L-atm.}$$

If the partial pressure of N_2 is doubled, Henry's law predicts that the solubility in water at 25 °C will also double, to 1.36×10^{-3} M.

Bottles use the effect of pressure on solubility in producing carbonated beverages such as soft drinks.

These are bottled under a carbon dioxide pressure greater than 1 atm.

When the bottles are opened to the air, the partial pressure of CO_2 above the solution decreases.

Hence, the solubility of CO_2 decreases, and $\text{CO}_2(\text{g})$ escapes from the solution as bubbles.



Solubility decreases as pressure decreases. CO_2 bubbles out of solution when a carbonated beverage is opened, because the CO_2 partial pressure above the solution is reduced.

Sample Exercise A Henry's Law Calculation

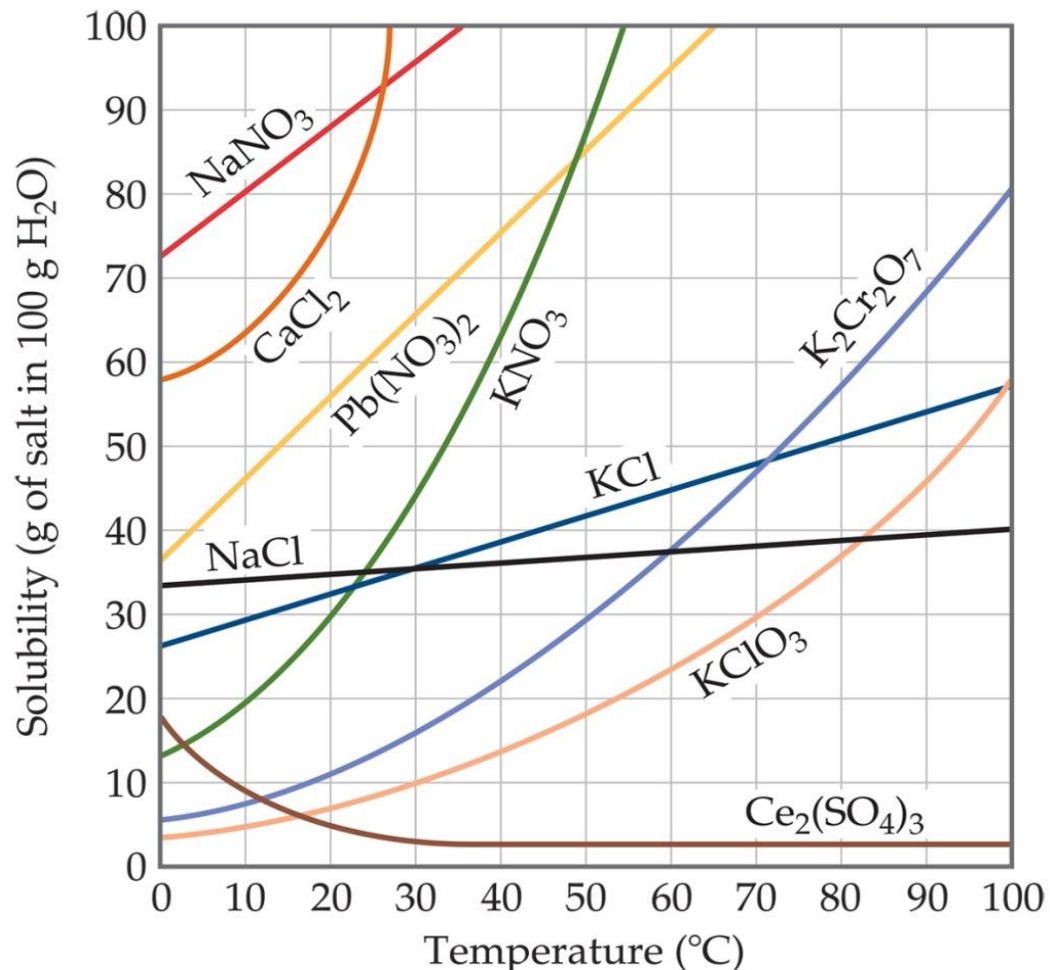
Calculate the concentration of CO₂ in a soft drink that is bottled with a partial pressure of CO₂ of 4.0 atm over the liquid at 25 °C. The Henry's law constant for CO₂ in water at this temperature is 3.1×10^{-2} mol/L-atm.

Solution

Solve: $S_{\text{CO}_2} = kP_{\text{CO}_2} = (3.1 \times 10^{-2} \text{ mol/L-atm})(4.0 \text{ atm})$
 $= 0.12 \text{ mol/L} = 0.12 \text{ M}$

Temperature Effects

Generally, the solubility of most solid solutes in water increases as the temperature of the solution increases.

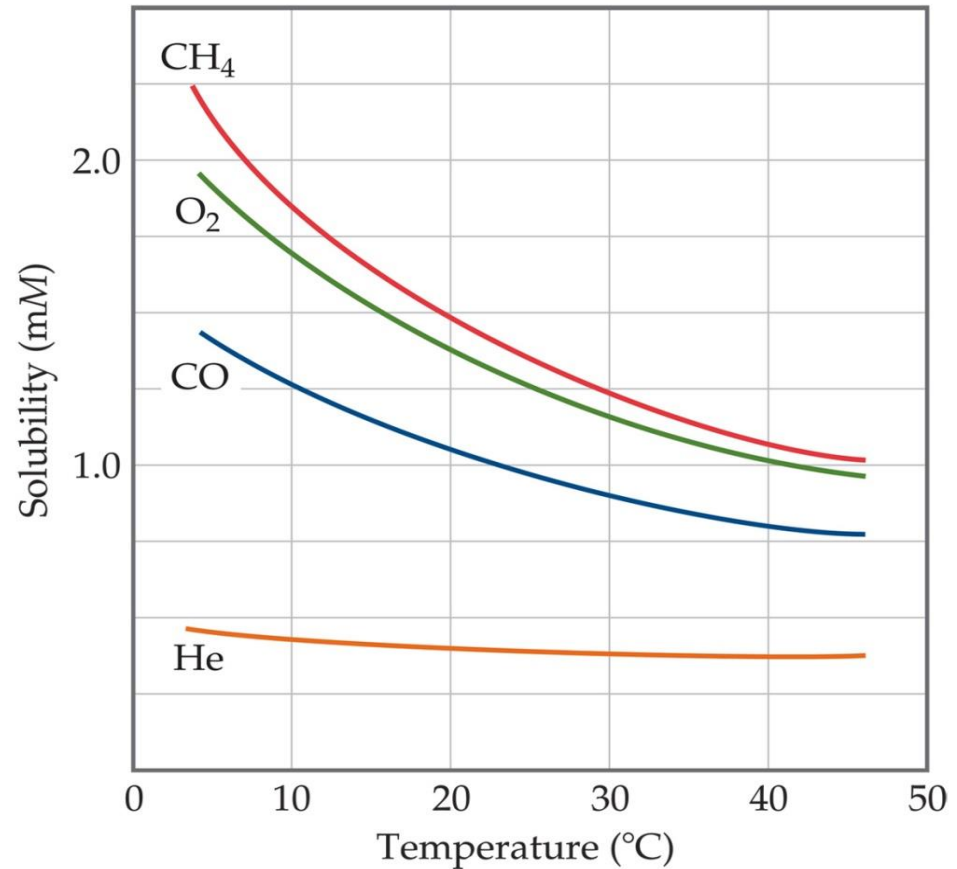


The opposite is true of gases.

The solubility of gases in water decreases with increasing temperature.

Examples:

- As the temperature of the solution increases, the solubility of CO_2 decreases, and $\text{CO}_2(\text{g})$ escapes from the solution.
- Carbonated soft drinks are more “bubbly” if stored in the refrigerator.
- When a glass of cold tap water is warmed, you can see bubbles of air on the inside of the glass.
- Warm lakes and streams have less O_2 dissolved in them than cool lakes (thermal pollution), fish may suffocate and die under these conditions.



13.5

Colligative Properties

Colligative Properties

Some physical properties of solutions differ in important ways from those of pure solvent. For example, pure water freezes at 0 °C, but aqueous solutions freeze at lower temperatures.

Ethylene glycol is added to the water in radiators of cars as an antifreeze to lower the freezing point of the solution. It also raises the boiling point of the solution above that of pure water.

Colligative properties depend only on the number or concentration of solute particles present, not on the identity or kind of the solute particles.

Among colligative properties are:

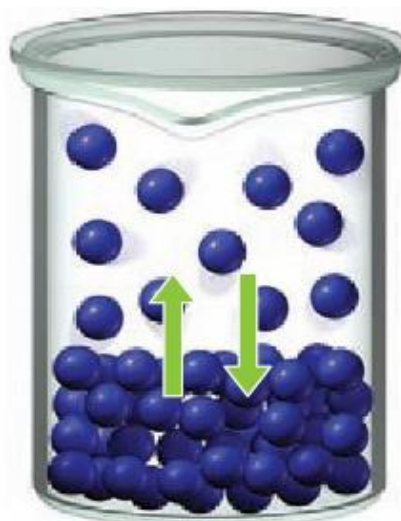
- Vapor pressure lowering.
- Boiling point elevation.
- Melting point depression.
- Osmotic pressure.

Lowering the Vapor Pressure

A liquid in a closed container will establish equilibrium with its vapor. When that equilibrium is reached, the pressure exerted by the vapor is called the **vapor pressure**.

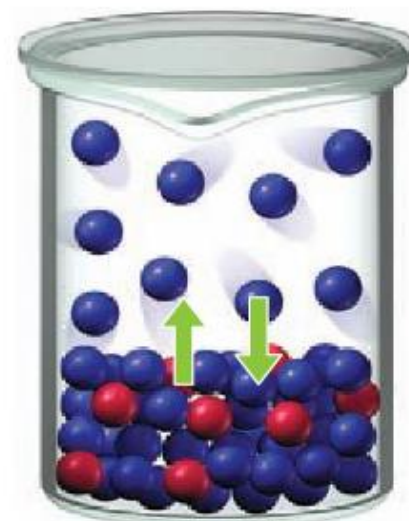
As solute molecules are added to a solution, the solvent become less volatile (decreased vapor pressure).

Solute-solvent interactions contribute to this effect, because of solute-solvent intermolecular attraction, higher concentrations of nonvolatile solutes make it harder for solvent to escape to the vapor phase.



Solvent alone

(a)



Solvent + solute

(b)

Raoult's Law

The extent to which a nonvolatile solute lowers the vapor pressure is proportional to its concentration. This relationship is expressed by Raoult's law.

Raoult's law states that the partial pressure exerted by solvent vapor above a solution, P_A , equals the product of the mole fraction of the solvent in the solution, X_A , times the vapor pressure of the pure solvent, P_A° .

$$P_A = X_A P_A^\circ$$

where:

X_A : is the mole fraction of solvent A in solution.

P_A° : is the normal vapor pressure of pure solvent A at that temperature.

P_A : is the vapor pressure above a solution.

NOTE: This is one of those times when you want to make sure you have the vapor pressure of the **solvent**.

For example

The vapor pressure of water is 17.5 torr at 20 °C. Imagine holding the temperature constant while adding glucose ($C_6H_{12}O_6$) to the water so that the resulting solution has $X_{H_2O} = 0.8$ and $X_{C_6H_{12}O_6} = 0.2$.

The vapor pressure of water over the solution will be 80% of that of pure water:

$$P_{H_2O} = (0.8) \times (17.5 \text{ torr}) = 14.0 \text{ torr}$$

In other words, the presence of the nonvolatile solute lowers the vapor pressure of the volatile solvent by $17.5 \text{ torr} - 14.0 \text{ torr} = 3.5 \text{ torr}$.

Sample Exercise Calculation of Vapor-Pressure Lowering

Glycerin ($\text{C}_3\text{H}_8\text{O}_3$) is a nonvolatile nonelectrolyte with a density of 1.26 g/mL at 25 °C. Calculate the vapor pressure at 25 °C of a solution made by adding 50.0 mL of glycerin to 500.0 mL of water. The vapor pressure of pure water at 25 °C is 23.8 torr, and its density is 1.00 g/mL.

Solution

To calculate the mole fraction of water in the solution, we must determine the number of moles of $\text{C}_3\text{H}_8\text{O}_3$ and H_2O :

$$\text{Moles } \text{C}_3\text{H}_8\text{O}_3 = (50.0 \text{ mL } \text{C}_3\text{H}_8\text{O}_3) \left(\frac{1.26 \text{ g } \text{C}_3\text{H}_8\text{O}_3}{1 \text{ mL } \text{C}_3\text{H}_8\text{O}_3} \right) \left(\frac{1 \text{ mol } \text{C}_3\text{H}_8\text{O}_3}{92.1 \text{ g } \text{C}_3\text{H}_8\text{O}_3} \right) = 0.684 \text{ mol}$$

$$\text{Moles } \text{H}_2\text{O} = (500.0 \text{ mL } \text{H}_2\text{O}) \left(\frac{1.00 \text{ g } \text{H}_2\text{O}}{1 \text{ mL } \text{H}_2\text{O}} \right) \left(\frac{1 \text{ mol } \text{H}_2\text{O}}{18.0 \text{ g } \text{H}_2\text{O}} \right) = 27.8 \text{ mol}$$

$$X_{\text{H}_2\text{O}} = \frac{\text{mol } \text{H}_2\text{O}}{\text{mol } \text{H}_2\text{O} + \text{mol } \text{C}_3\text{H}_8\text{O}_3} = \frac{27.8}{27.8 + 0.684} = 0.976$$

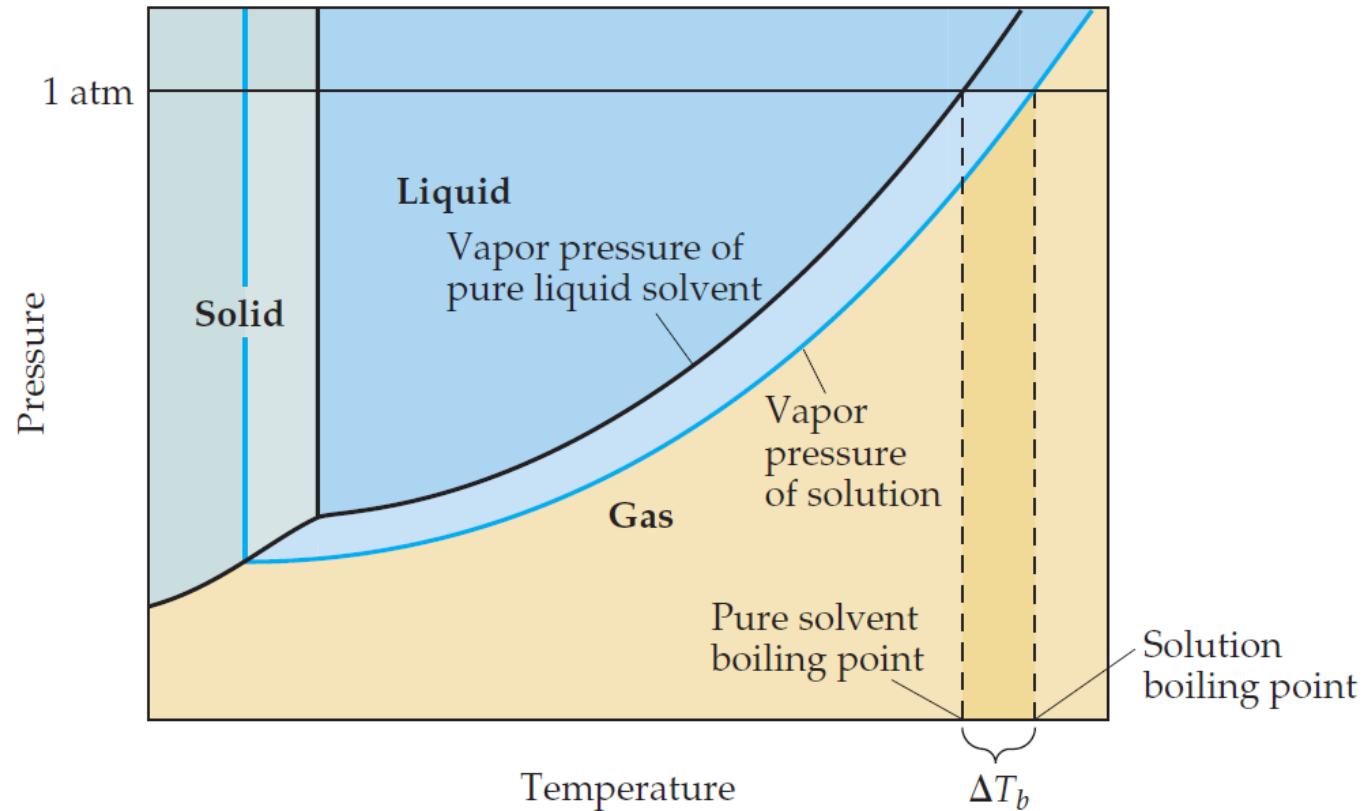
$$P_{\text{H}_2\text{O}} = X_{\text{H}_2\text{O}} P_{\text{H}_2\text{O}}^\circ = (0.976)(23.8 \text{ torr}) = 23.2 \text{ torr}$$

The vapor pressure of the solution has been lowered by 0.6 torr relative to that of pure water

Boiling Point Elevation

Nonvolatile solute-solvent interactions also cause solutions to have higher boiling points than the pure solvent.

The vapor pressure curve of the solution shifted downward relative to the vapor pressure of the pure liquid at any given temperature.



Phase diagram for a pure solvent and for a solution of a nonvolatile solute.

At the normal boiling point of the pure liquid, the vapor pressure of the solution will be less than 1 atm. Therefore, a higher temperature is required to attain a vapor pressure of 1 atm. Thus, the boiling point of the solution is higher than that of the pure liquid.

The change in boiling point is proportional to the **molality** (m) of the solution:

$$\Delta T_b = K_b m$$

where:

K_b is the molal boiling point elevation constant, a property of the solvent.

ΔT_b is added to the normal boiling point of the solvent.

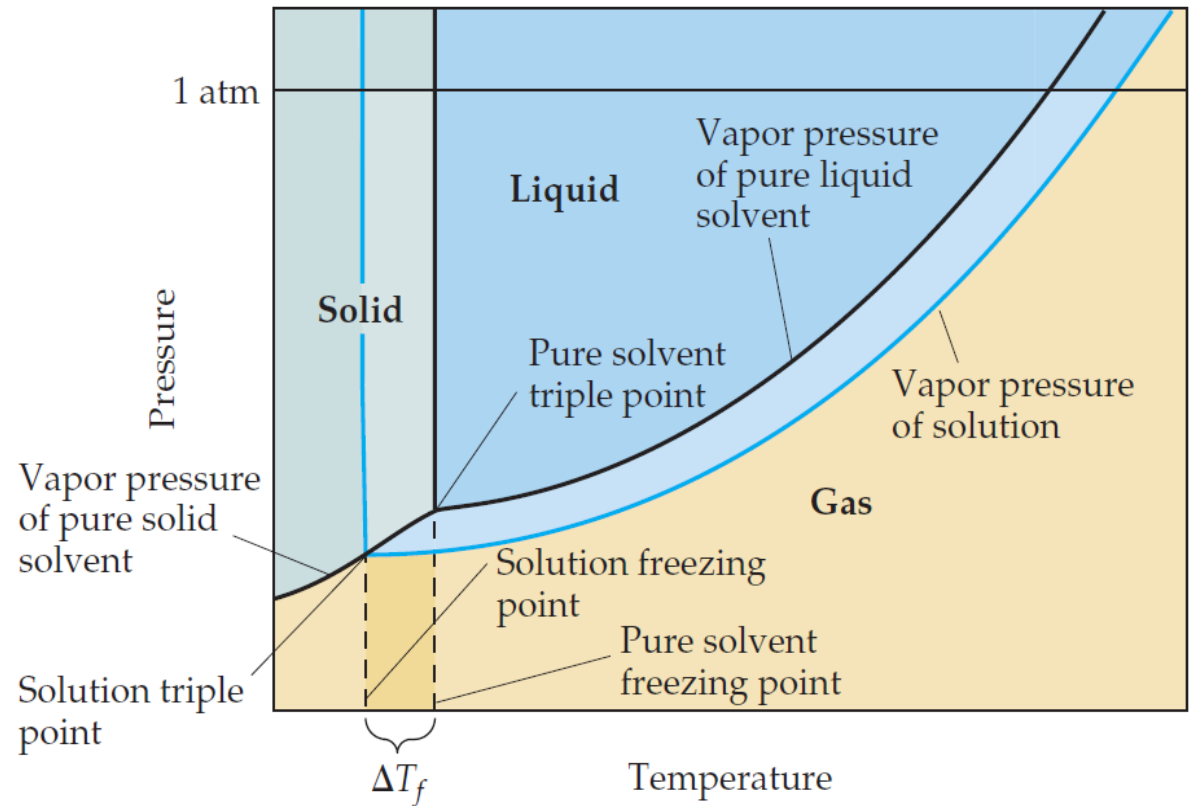
The increase in boiling point relative to that of the pure solvent, ΔT_b is a +ve quantity obtained by subtracting the boiling point of the pure solvent from the boiling point of the solution.

$$\Delta T_b = T_{soln} - T_{pure\ solvent}$$

Freezing Point Depression

Nonvolatile solute-solvent interactions also cause solutions to have lower freezing points than the pure solvent.

e.g., Antifreeze being added to an automobile radiator. Antifreeze consists of an aqueous solution of ethylene glycol, $\text{CH}_2(\text{OH})\text{CH}_2(\text{OH})$.



Phase diagram for a pure solvent and for a solution of a nonvolatile solute.

The change in freezing point can be found similarly:

$$\Delta T_f = K_f m$$

where:

K_f is the molal freezing point depression constant of the solvent.

ΔT_f is subtracted from the normal freezing point of the solvent.

The decrease in freezing point, ΔT_f is a +ve quantity obtained by subtracting the freezing point of the solution from the freezing point of the pure solvent.

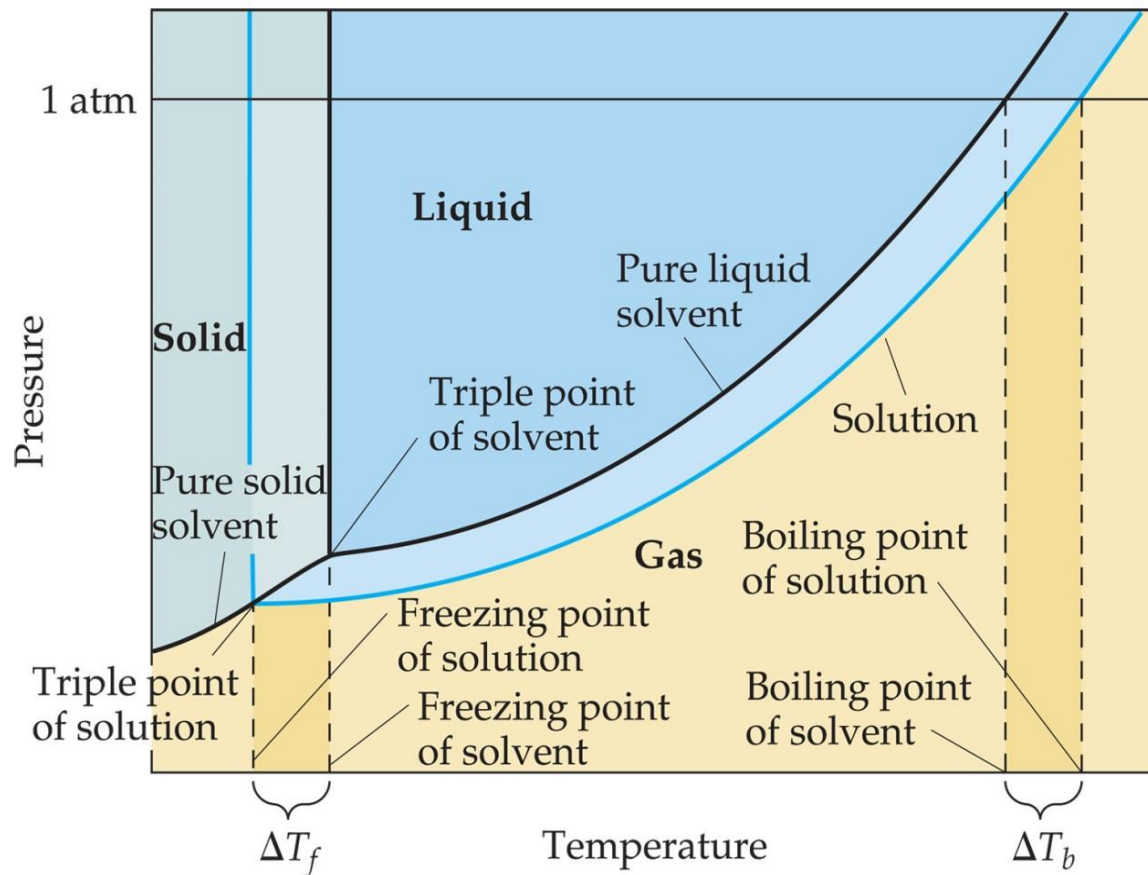
$$\Delta T_f = T_{\text{pure solvent}} - T_{\text{soln}}$$

Boiling Point Elevation and Freezing Point Depression

$$\Delta T_b = K_b \cdot m$$

$$\Delta T_f = K_f \cdot m$$

Note that in both equations, ΔT does not depend on what the solute is, but only on how many particles are dissolved.



Phase diagram for a pure solvent and for a solution of a nonvolatile solute.

TABLE 13.4 ■ Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/ <i>m</i>)	Normal Freezing Point (°C)	K_f (°C/ <i>m</i>)
Water, H ₂ O	100.0	0.51	0.0	1.86
Benzene, C ₆ H ₆	80.1	2.53	5.5	5.12
Ethanol, C ₂ H ₅ OH	78.4	1.22	-114.6	1.99
Carbon tetrachloride, CCl ₄	76.8	5.02	-22.3	29.8
Chloroform, CHCl ₃	61.2	3.63	-63.5	4.68

For water, K_b is 0.51 °C/*m*; therefore, a 1 *m* aqueous solution of sucrose or any other aqueous solution that is 1 *m* in nonvolatile solute particles will boil 0.51 °C higher than pure water.

The boiling-point elevation is proportional to the concentration of solute particles, regardless of whether the particles are molecules or ions.

TABLE 13.4 ■ Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/ <i>m</i>)	Normal Freezing Point (°C)	K_f (°C/ <i>m</i>)
Water, H ₂ O	100.0	0.51	0.0	1.86
Benzene, C ₆ H ₆	80.1	2.53	5.5	5.12
Ethanol, C ₂ H ₅ OH	78.4	1.22	-114.6	1.99
Carbon tetrachloride, CCl ₄	76.8	5.02	-22.3	29.8
Chloroform, CHCl ₃	61.2	3.63	-63.5	4.68

For water, K_f is 1.86 °C/*m*; therefore, a 1 *m* aqueous solution of sucrose or any other aqueous solution that is 1 *m* in nonvolatile solute particles (such as 0.5 *m* **NaCl**) will freeze 1.86 °C lower than pure water.

The freezing-point lowering caused by solutes explains the use of antifreeze in cars and the use of calcium chloride (**CaCl₂**) to melt ice on roads during winter.

Sample Exercise Calculation of Boiling-Point Elevation and Freezing-Point Lowering

Automotive antifreeze consists of ethylene glycol, $\text{CH}_2(\text{OH})\text{CH}_2(\text{OH})$, a nonvolatile nonelectrolyte. Calculate the boiling point and freezing point of a 25.0 mass % solution of ethylene glycol in water.

Solution

The molality of the solution is calculated as follows:

$$\begin{aligned}\text{Molality} &= \frac{\text{moles C}_2\text{H}_6\text{O}_2}{\text{kilograms H}_2\text{O}} = \left(\frac{250 \text{ g C}_2\text{H}_6\text{O}_2}{750 \text{ g H}_2\text{O}}\right) \left(\frac{1 \text{ mol C}_2\text{H}_6\text{O}_2}{62.1 \text{ g C}_2\text{H}_6\text{O}_2}\right) \left(\frac{1000 \text{ g H}_2\text{O}}{1 \text{ kg H}_2\text{O}}\right) \\ &= 5.37 \text{ m}\end{aligned}$$

We can now use Equations to calculate the changes in the boiling and freezing points:

$$\begin{aligned}\Delta T_b &= K_b m = (0.51 \text{ }^\circ\text{C}/m)(5.37 \text{ m}) = 2.7 \text{ }^\circ\text{C} \\ \Delta T_f &= K_f m = (1.86 \text{ }^\circ\text{C}/m)(5.37 \text{ m}) = 10.0 \text{ }^\circ\text{C}\end{aligned}$$

Hence, the boiling and freezing points of the solution are

$$\begin{aligned}\text{Boiling point} &= (\text{normal bp of solvent}) + \Delta T_b \\ &= 100.0 \text{ }^\circ\text{C} + 2.7 \text{ }^\circ\text{C} = 102.7 \text{ }^\circ\text{C} \\ \text{Freezing point} &= (\text{normal fp of solvent}) - \Delta T_f \\ &= 0.0 \text{ }^\circ\text{C} - 10.0 \text{ }^\circ\text{C} = -10.0 \text{ }^\circ\text{C}\end{aligned}$$

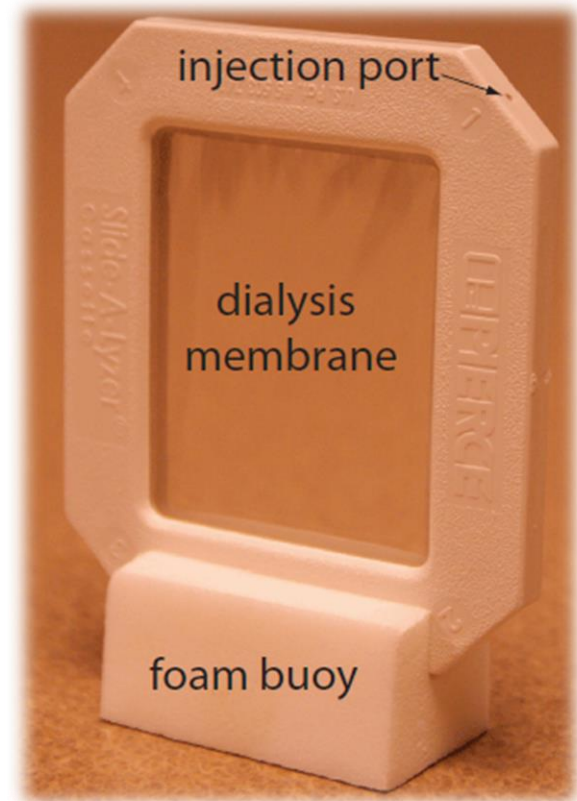
Comment: Notice that the solution is a liquid over a larger temperature range than the pure solvent.

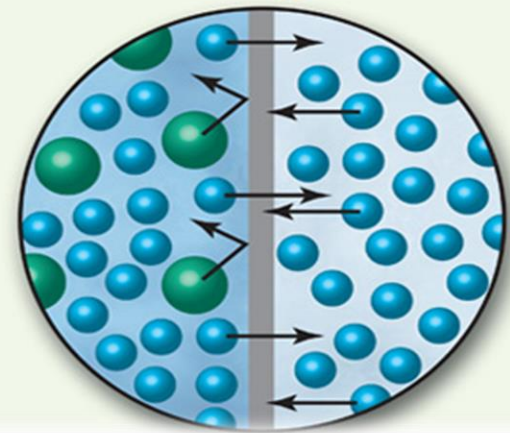
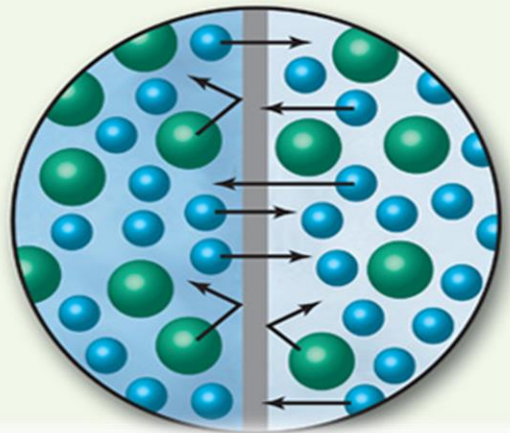
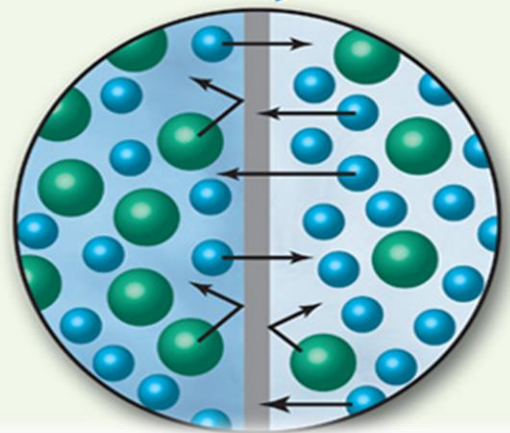
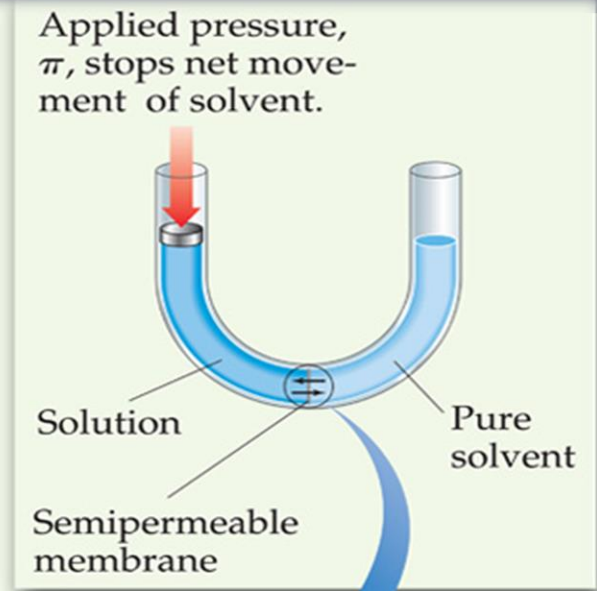
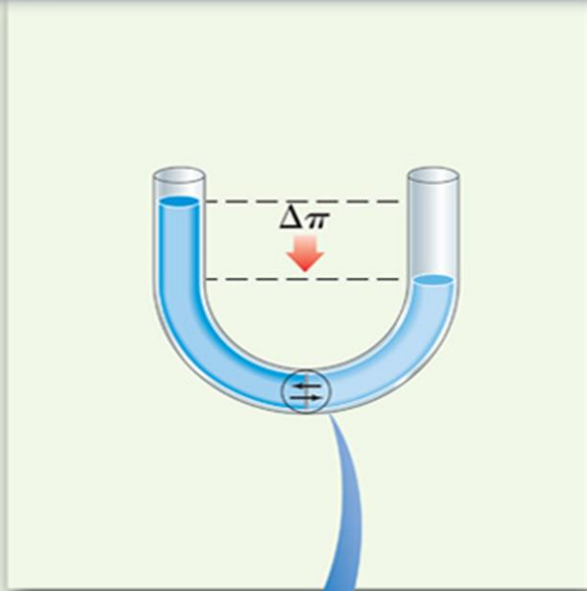
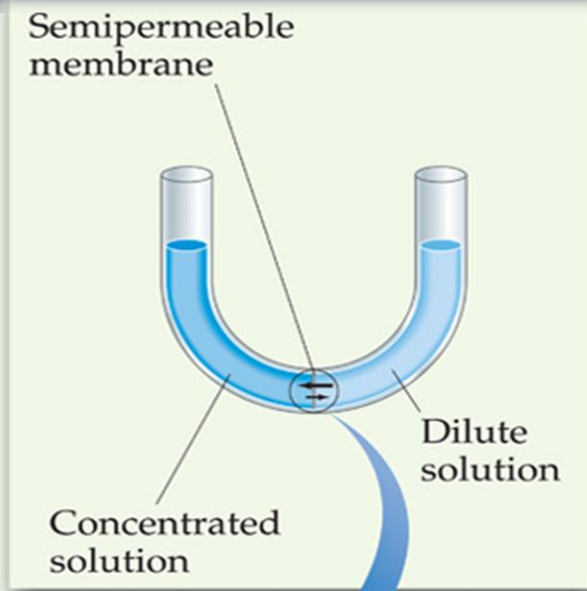
Osmosis

Consider a situation in which only solvent molecules are able to pass through a membrane (**semipermeable membranes**, allowing some smaller particles to pass through, but blocking other larger particles).

If such a membrane is placed between two solutions of different concentration, solvent molecules move in both directions through the membrane.

A net movement of solvent molecules **from the less concentrated solution into the more concentrated one**. In this process, called **osmosis**, the net movement of solvent is always toward the solution with the higher solute concentration.





Net movement of solvent from the pure solvent or a solution with low solute concentration to a solution with high solute concentration.

Osmosis stops when the column of a soln on the left becomes high enough to exert sufficient pressure at the membrane to counter the net movement of solvent. At this point the soln on the left has become more dilute, but there still exists a difference in concentrations between the two solns.

Applied pressure on the left arm of the apparatus stops net movement of solvent from the right side of the semipermeable membrane. This applied pressure is the osmotic pressure of the soln.

The pressure required to stop or to prevent osmosis by pure solvent, known as **osmotic pressure**, π , of the solution.

The osmotic pressure obeys a law similar in form to the ideal gas law:

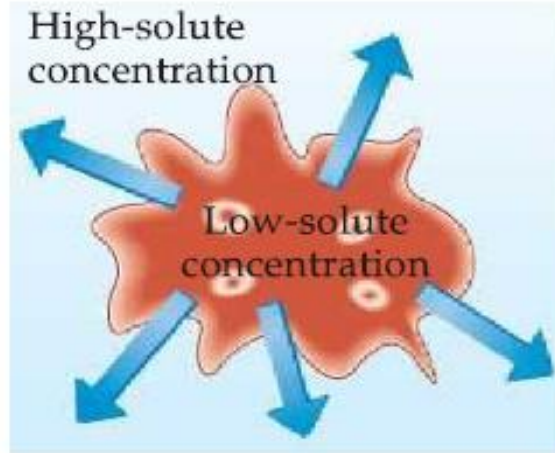
$$\Pi = \left(\frac{n}{V} \right) RT = MRT$$

Where: M is the molarity of the solution, V is the volume of the solution, n is the number of moles of solute, R is the ideal gas constant, and T is the temperature on the Kelvin scale

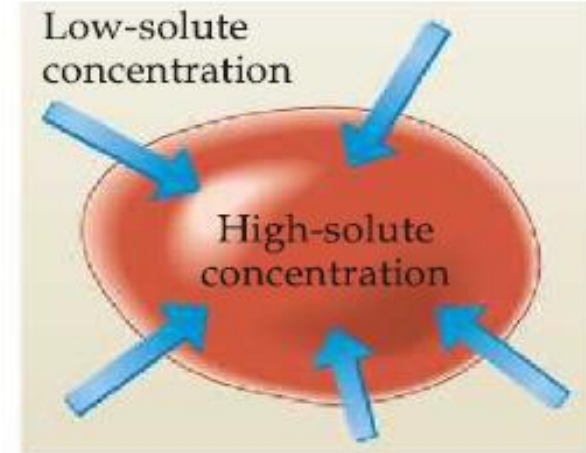
If two solutions of identical osmotic pressure are separated by a semipermeable membrane, no osmosis will occur. The two solutions are **isotonic**. If one solution is of lower osmotic pressure, it is **hypotonic** with respect to the more concentrated solution. The more concentrated solution is **hypertonic** with respect to the dilute solution.

Osmosis plays a very important role in living systems.

The membranes of red blood cells, for example, are semipermeable.



(a) Crenation



(b) Hemolysis

The blue arrows represent the net movement of water molecules.

Water will flow out of the cell, and **crenation** results (causes the cell to **shrivel**).

Water will flow into the cell, and **hemolysis** results (causes the cell to **rupture**).

Sample Exercise Calculations Involving Osmotic Pressure

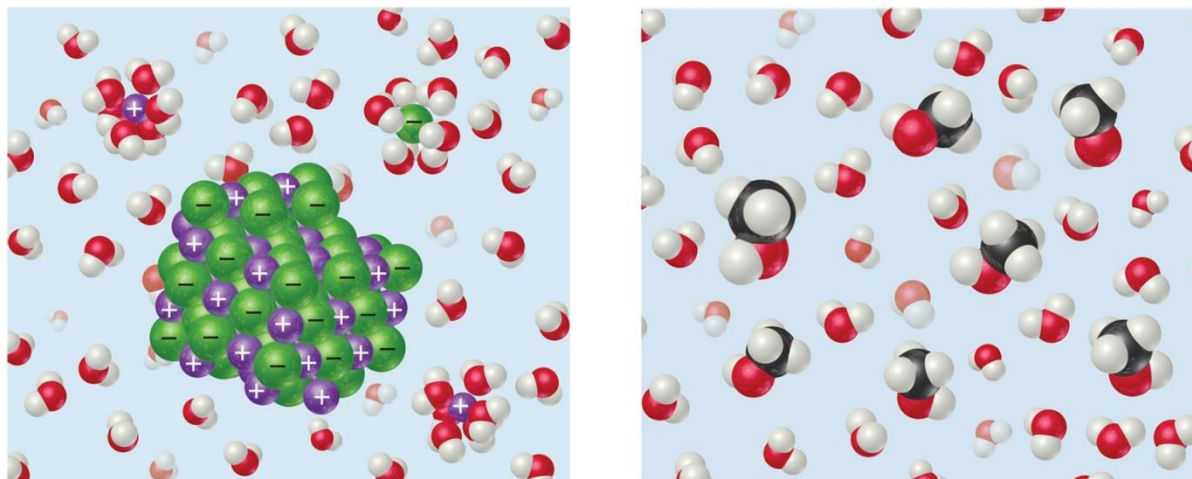
The average osmotic pressure of blood is 7.7 atm at 25 °C. What molarity of glucose ($C_6H_{12}O_6$) will be isotonic with blood?

Solution

$$\begin{aligned} \Pi &= MRT \\ M &= \frac{\Pi}{RT} = \frac{7.7 \text{ atm}}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 0.31 \text{ M} \end{aligned}$$

Colligative Properties of Electrolytes

Since these properties depend on the number of particles dissolved, solutions of electrolytes (which dissociate in solution) should show greater changes than those of nonelectrolytes. The ideal value of i can be determined for a salt from the number of ions per formula unit.

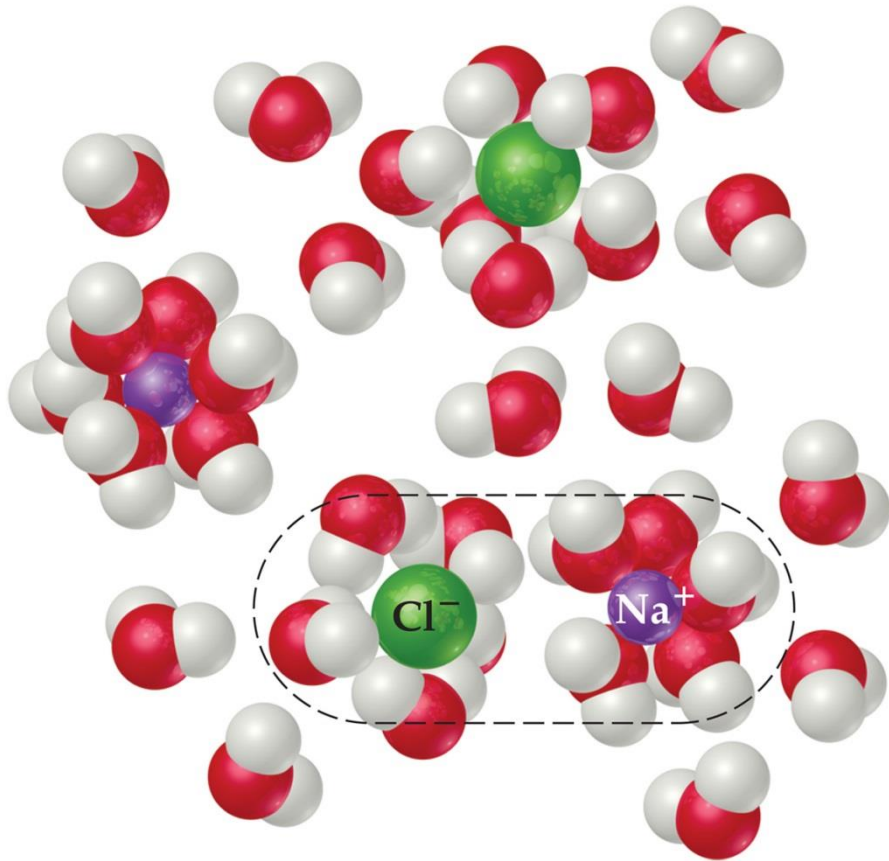


e.g. **NaCl** dissociates to form 2 ion particles; consists of one Na^+ and one Cl^- per formula unit, its limiting van't Hoff factor is 2.

For **K_2SO_4** it is 3 because it consists of two K^+ and one SO_4^{-2} .

van't Hoff Factor

One mole of NaCl in water does not really give rise to two moles of ions.



Ion pairing

Some **Na⁺** and **Cl⁻** reassociate for a short time (as a result of electrostatic attractions between ions),

Therefore, the true concentration of particles is somewhat less than two times the concentration of **NaCl**.

One measure of the extent to which electrolytes dissociate is the **van't Hoff factor, i** .

This factor is the ratio of the actual value of a colligative property to the value calculated when the substance is assumed to be a nonelectrolyte.

Using the freezing point depression, for example, we have

$$i = \frac{\Delta T_f(\text{measured})}{\Delta T_f(\text{calculated for nonelectrolyte})}$$

Reassociation is more likely at higher concentration.

Therefore, the number of particles present is concentration-dependent.

TABLE 13.5 ■ van't Hoff Factors for Several Substances at 25 °C

Compound	Concentration			Limiting Value
	0.100 <i>m</i>	0.0100 <i>m</i>	0.00100 <i>m</i>	
Sucrose	1.00	1.00	1.00	1.00
NaCl	1.87	1.94	1.97	2.00
K ₂ SO ₄	2.32	2.70	2.84	3.00
MgSO ₄	1.21	1.53	1.82	2.00

We modify the previous equations by multiplying by the van't Hoff factor, *i*.

i = 1 for non-electrolytes

$$\Delta T_b = K_b \times m \times i$$

$$\Delta T_f = K_f \times m \times i$$

Sample Exercise Freezing-Point Depression in Aqueous Solutions

List the following aqueous solutions in order of their expected freezing point:
 $0.050\ m\ \text{CaCl}_2$, $0.15\ m\ \text{NaCl}$, $0.10\ m\ \text{HCl}$, $0.050\ m\ \text{CH}_3\text{COOH}$, $0.10\ m\ \text{C}_{12}\text{H}_{22}\text{O}_{11}$.

Solution

CaCl_2 , NaCl , and HCl are strong electrolytes, CH_3COOH (acetic acid) is a weak electrolyte, and $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ is a nonelectrolyte. The molality of each solution in total particles is as follows:

$0.050\ m\ \text{CaCl}_2 \Rightarrow 0.050\ m\ \text{in}\ \text{Ca}^{2+}\ \text{and}\ 0.10\ m\ \text{in}\ \text{Cl}^- \Rightarrow 0.15\ m\ \text{in}\ \text{particles}$

$0.15\ m\ \text{NaCl} \Rightarrow 0.15\ m\ \text{Na}^+\ \text{and}\ 0.15\ m\ \text{in}\ \text{Cl}^- \Rightarrow 0.30\ m\ \text{in}\ \text{particles}$

$0.10\ m\ \text{HCl} \Rightarrow 0.10\ m\ \text{H}^+\ \text{and}\ 0.10\ m\ \text{in}\ \text{Cl}^- \Rightarrow 0.20\ m\ \text{in}\ \text{particles}$

$0.050\ m\ \text{CH}_3\text{COOH} \Rightarrow \text{weak electrolyte} \Rightarrow \text{between}\ 0.050\ m\ \text{and}\ 0.10\ m\ \text{in}\ \text{particles}$

$0.10\ m\ \text{C}_{12}\text{H}_{22}\text{O}_{11} \Rightarrow \text{nonelectrolyte} \Rightarrow 0.10\ m\ \text{in}\ \text{particles}$

Because the freezing points depend on the total molality of particles in solution, the expected ordering is $0.15\ m\ \text{NaCl}$ (lowest freezing point), $0.10\ m\ \text{HCl}$, $0.050\ m\ \text{CaCl}_2$, $0.10\ m\ \text{C}_{12}\text{H}_{22}\text{O}_{11}$, and $0.050\ m\ \text{CH}_3\text{COOH}$ (highest freezing point).

Determination of Molar Mass from Colligative Properties

Sample Exercise

A solution of an unknown nonvolatile nonelectrolyte was prepared by dissolving 0.250 g of the substance in 40.0 g of CCl_4 . The boiling point of the resultant solution was $0.357\text{ }^\circ\text{C}$ higher than that of the pure solvent. Calculate the molar mass of the solute.

Solution

$$\text{Molality} = \frac{\Delta T_b}{K_b} = \frac{0.357\text{ }^\circ\text{C}}{5.02\text{ }^\circ\text{C}/m} = 0.0711\text{ } m$$

$$(0.0400\text{ kg CCl}_4) \left(0.0711 \frac{\text{mol solute}}{\text{kg CCl}_4} \right) = 2.84 \times 10^{-3}\text{ mol solute}$$

$$\text{Molar mass} = \frac{0.250\text{ g}}{2.84 \times 10^{-3}\text{ mol}} = 88.0\text{ g/mol}$$

Sample Exercise

The osmotic pressure of an aqueous solution of a certain protein was measured to determine the protein's molar mass. The solution contained 3.50 mg of protein dissolved in sufficient water to form 5.00 mL of solution. The osmotic pressure of the solution at 25 °C was found to be 1.54 torr. Treating the protein as a nonelectrolyte, calculate its molar mass.

Solution

$$\text{Molarity} = \frac{\Pi}{RT} = \frac{(1.54 \text{ torr})\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 8.28 \times 10^{-5} \frac{\text{mol}}{\text{L}}$$

Because the volume of the solution is 5.00 ml = 5.00×10^{-3} L, the number of moles of protein must be

$$\text{Moles} = (8.28 \times 10^{-5} \text{ mol/L})(5.00 \times 10^{-3} \text{ L}) = 4.14 \times 10^{-7} \text{ mol}$$

The molar mass is the number of grams per mole of the substance. The sample has a mass of 3.50 mg = 3.50×10^{-3} g. The molar mass is the number of grams divided by the number of moles:

$$\text{Molar mass} = \frac{\text{grams}}{\text{moles}} = \frac{3.50 \times 10^{-3} \text{ g}}{4.14 \times 10^{-7} \text{ mol}} = 8.45 \times 10^3 \text{ g/mol}$$



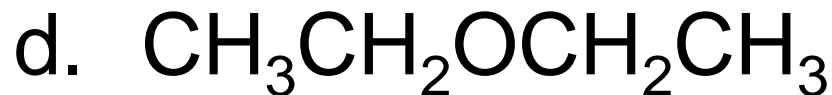
Q & A



The “like dissolves like” rule states that substances with similar _____ will dissolve in each other.

- a. molecular weights
- b. molecular shapes
- c. intermolecular attractive forces
- d. densities

Which substance below is miscible with water?



_____ Law states that the solubility of a gas in a liquid increases as the pressure of the gas above the liquid increases.

- a. Boyle's
- b. Charles'
- c. Henry's
- d. Raoult's

_____ Law states that the vapor pressure of a solution is directly proportional to the mole fraction of solvent present.

- a. Boyle's
- b. Charles'
- c. Henry's
- d. Raoult's

Which aqueous solution should have the lowest freezing point?

- a. 0.100 *m* NaCl
- b. 0.100 *m* CaCl₂
- c. 0.100 *m* KI
- d. 0.100 *m* LiNO₃

Isotonic solutions have the same:

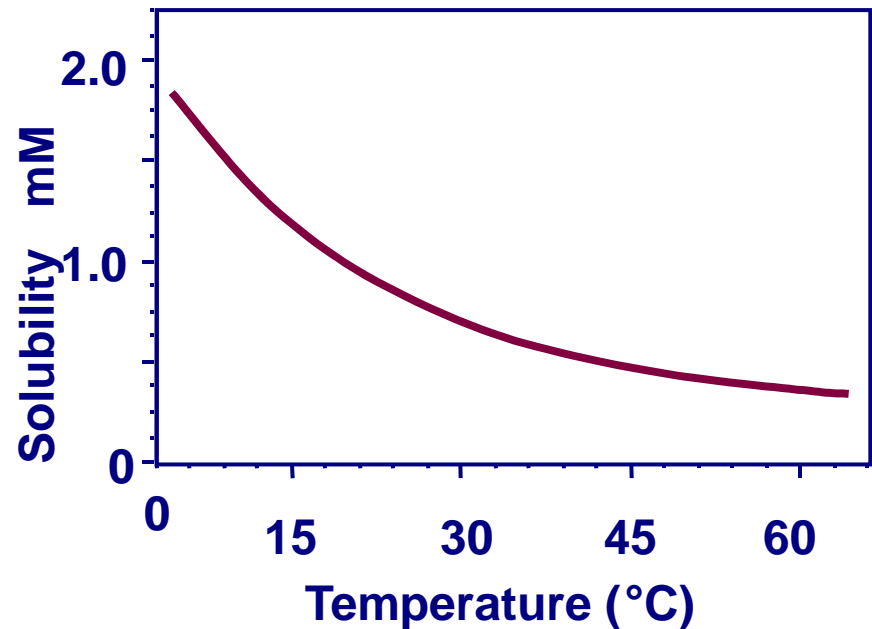
- a. density
- b. volume
- c. vapor pressure
- d. osmotic pressure

A solution contains 0.02755 g of a protein in 25.0 mL of solution. The osmotic pressure is 3.22 torr at 298 K. What is the molecular weight of the protein?

- a. 254 g/mole
- b. 6360 g/mole**
- c. 159,000 g/mole
- d. 2,340,000 g/mole

The curve describes the solubility properties of

- a liquid
- an ionic solid
- a covalent solid
- **a gas**
- a gas or liquid



Considering the substances given below, predict which will be soluble in water.

- 1



- 2

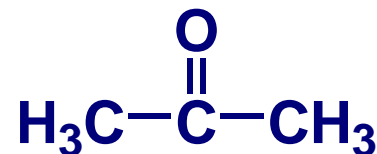
1

- 3

- 2, 3



2



3

- 1, 2, 3

Predict which aqueous solution will have the lowest freezing point.

- 0.25 *m* C₂H₅OH
- 0.15 *m* CaCl₂
- 0.20 *m* NaCl
- 0.15 *m* NH₄NO₃
- 0.15 *m* Na₃PO₄

Arrange the aqueous solutions according to increasing boiling point.

- $\text{AlCl}_3 < \text{KNO}_3 < \text{Na}_2\text{SO}_4$
- $\text{Na}_2\text{SO}_4 < \text{AlCl}_3 < \text{KNO}_3$
- $\text{Na}_2\text{SO}_4 < \text{KNO}_3 < \text{AlCl}_3$
- $\text{KNO}_3 < \text{AlCl}_3 < \text{Na}_2\text{SO}_4$
- $\text{KNO}_3 < \text{Na}_2\text{SO}_4 < \text{AlCl}_3$

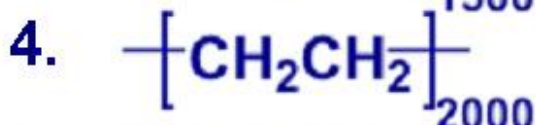
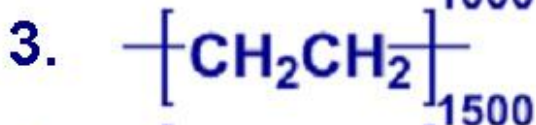
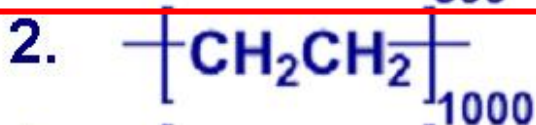
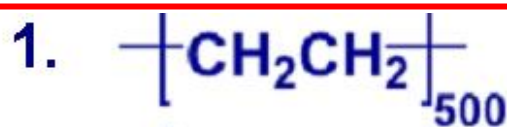
0.10 m Na_2SO_4

0.15 m AlCl_3

0.20 m KNO_3

Which polyethylene solution will produce the largest osmotic pressure?

Assume that all solutions use the same solvent and have the same concentration



5. All will have the same osmotic pressure.

Calculate the concentration of CO_2 in a soft drink after the bottle is opened and equilibrates at $25\text{ }^\circ\text{C}$ under a CO_2 partial pressure of $3.0 \times 10^{-4}\text{ atm}$.

Answer: $9.3 \times 10^{-6}\text{ M}$

The vapor pressure of pure water at 110 °C is 1070 torr. A solution of ethylene glycol and water has a vapor pressure of 1.00 atm at 110 °C. Assuming that Raoult's law is obeyed, what is the mole fraction of ethylene glycol in the solution?

Answer: 0.290

Calculate the freezing point of a solution containing 0.600 kg of CHCl_3 and 42.0 g of eucalyptol ($\text{C}_{10}\text{H}_{18}\text{O}$), a fragrant substance found in the leaves of eucalyptus trees. (See Table 13.4.)

Answer: $-65.6\text{ }^\circ\text{C}$

Which of the following solutes will produce the largest increase in boiling point upon addition to 1 kg of water: 1 mol of $\text{Co}(\text{NO}_3)_2$, 2 mol of KCl , 3 mol of ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$)?

Answer: 2 mol of KCl because it contains the highest concentration of particles, 2 *m* K^+ and 2 *m* Cl^- , giving 4 *m* in all

What is the osmotic pressure at 20 °C of a 0.0020 M sucrose ($C_{12}H_{22}O_{11}$) solution?

Answer: 0.048 atm, or 37 torr

Camphor ($\text{C}_{10}\text{H}_{16}\text{O}$) melts at $179.8\text{ }^\circ\text{C}$, and it has a particularly large freezing-point-depression constant, $K_f = 40.0\text{ }^\circ\text{C}/\text{m}$. When 0.186 g of an organic substance of unknown molar mass is dissolved in 22.01 g of liquid camphor, the freezing point of the mixture is found to be $176.7\text{ }^\circ\text{C}$. What is the molar mass of the solute?

***Answer:* 110 g/mol**

A sample of 2.05 g of polystyrene of uniform polymer chain length was dissolved in enough toluene to form 0.100 L of solution. The osmotic pressure of this solution was found to be 1.21 kPa at 25 °C. Calculate the molar mass of the polystyrene.

Answer: 4.20×10^4 g/mol

