

IMPORTANT TERMS

- **Solution**—a homogeneous mixture of two or more substances in a single phase. Does *not* have to involve liquids. Air is a solution of nitrogen, oxygen, carbon dioxide etc.; solder is a solution of lead, tin etc.

- **solute**—component in lesser concentration; dissolvee
- **solvent**—component in greater concentration; dissolver

- **solubility**—maximum amount of material that will dissolve in a given amount of solvent at a given temperature to produce a stable solution. In other words, the solution is saturated.

Study the solubility rules!!

- **molar solubility**—the number of moles of solute that dissolves in exactly 1.0 L of solvent, expressed in units of molarity, M or the use of square brackets.

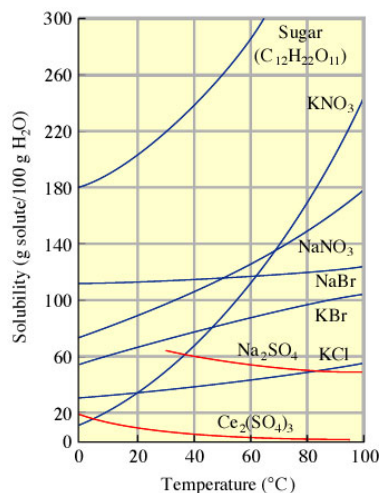


Table 11.1 Various Types of Solutions

Example	State of Solution	State of Solute	State of Solvent
Air, natural gas	Gas	Gas	Gas
Vodka in water, antifreeze	Liquid	Liquid	Liquid
Brass	Solid	Solid	Solid
Carbonated water (soda)	Liquid	Gas	Liquid
Seawater, sugar solution	Liquid	Solid	Liquid
Hydrogen in platinum	Solid	Gas	Solid

- **Saturated solution**—a solution containing the maximum amount of solute that will dissolve under a given set of pressure and temperature conditions. Saturated solutions are at dynamic equilibrium with any excess undissolved solute present. Solute particles dissolve and recrystallize at equal rates.

- **Unsaturated solution**—a solution containing less than the maximum amount of solute that will dissolve under a given set of conditions. (more solute can dissolve)

- **Supersaturated solution**—oxymoron—a solution that has been prepared at an elevated temperature and then slowly cooled. It contains more than the usual maximum amount of solution dissolved. A supersaturated solution is very unstable and agitation (stirring, pouring, etc.) or the addition of a “seed crystal” will cause all excess solute to crystallize out of solution leaving the remaining solvent saturated. (rock candy is made this way as are those liquid hand warmer packets of sodium acetate solution, pictured at left, folks use hunting, at football games, skiing, etc.)

- **miscible**—When two or more liquids mix (ex. Water and food coloring)
- **immiscible**—When two or more liquids DON'T mix.—they usually layer if allowed to set for a while. (ex. Water and oil)

Concentration Units

- **Molarity (M)** = # of moles of solute per liter of solution; IS temperature dependent. The liquid solvent can expand and contract with changes in temperature. Thus, not a constant ratio of solute:solvent particles. Most molar solutions are made at 25° C so this point is subtle and picky, but important none the less!!

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- **Mass percent (weight percent)** = percent by mass of the solute in the solution

$$\text{Mass percent} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$$

- **Mole fraction (χ)** = ratio of the number of moles of a given component to the total number of moles of present.

$$\text{Mole fraction}_a = \chi_a \quad \text{Mole Fraction}_a = \chi_a = \frac{n_a}{n_a + n_b + \dots}$$

- **Molality (m)** = # of moles of solute per kilogram of solvent; NOT temperature dependent. Represents a ratio of solute:solvent molecules at all times.

$$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Exercise 1 Various Methods for Describing Solution Composition

A solution is prepared by mixing 1.00g ethanol (C_2H_5OH) with 100.0g water to give a final volume of 101 mL. Calculate the molarity, mass percent, mole fraction, and molality of ethanol in this solution.

molarity = 0.215 *M*
mass percent = 0.990% C_2H_5OH
mole fraction = 0.00389
molality = 0.217 *m*

**Exercise 2 Calculating Various Methods of Solution Composition
 from the Molarity**

The electrolyte in automobile lead storage batteries is a 3.75 *M* sulfuric acid solution that has a density of 1.230 g/mL. Calculate the mass percent, molality, and normality of the sulfuric acid.

mass percent = 29.9% H_2SO_4
molality = 4.35 *m*
normality is 7.50 *N*

THE SOLUTION PROCESS**➤ Energies involved in solution formation**

When a solute is dissolved in a solvent, the attractive forces between solute and solvent particles are great enough to overcome the attractive forces within the pure solvent and within the pure solute. The solute becomes *solvated* (usually by dipole—dipole or ion—dipole forces). When the solvent is water the solute is *hydrated*.

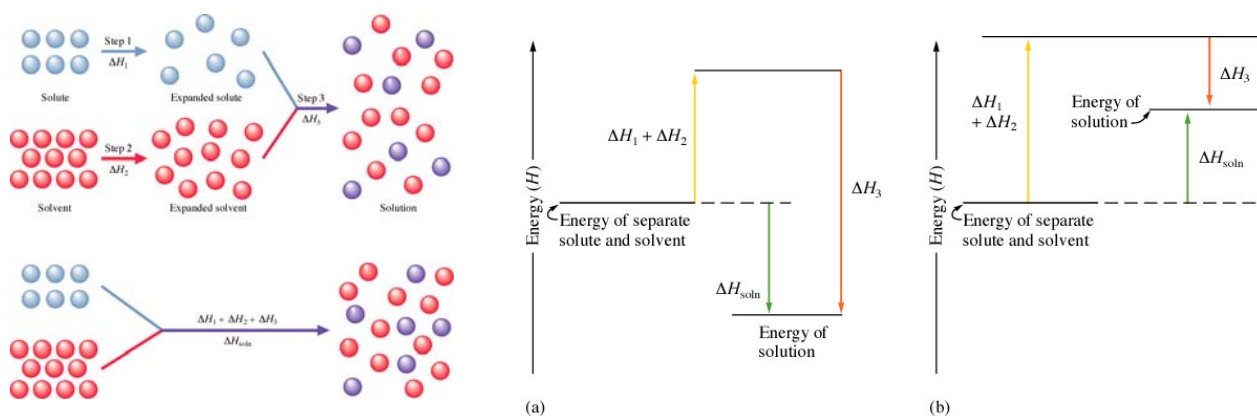
- Substances with similar types of intermolecular forces dissolve in each other. “Like dissolves like.”
 - Polar solvents dissolve polar or ionic solutes.
 - Nonpolar solvents dissolve nonpolar solutes.

- Water dissolves many salts because the stronger ion—dipole attractions water forms with the ions of the salt are very similar to the strong attractions between the ions themselves. The same salts are *insoluble* in hexane (C₆H₁₄) because the weaker London dispersion forces their ions could form with this nonpolar solvent are much weaker than the attraction between the ions of the salt.
- Oil does not dissolve in water. Oil is immiscible in water due to the fact that any weak dipole-induced dipole attractions that form between oil and water cannot overcome the stronger dipole-dipole hydrogen bonding that water molecules have for each other.
- Solubilities of alcohols in water:** As the hydrocarbon portion of the alcohol increases in length, the alcohol becomes less soluble. (More of the molecule is nonpolar; the dipole moment is diminished.)
- Solubilities of alcohol in nonpolar solvents:** As the hydrocarbon portion of the alcohol increases in length, the alcohol becomes *more* soluble in a nonpolar solvent such as hexane.

➤ **Enthalpy of solution (ΔH_{soln})** = the enthalpy change associated with the formation of a solution (just the sum of all of the steps involved!)

$$3 \text{ steps: } \Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

- ΔH_{soln} can be positive (endothermic) or negative (exothermic).



Step 1 (ΔH_1)

- Separating the solute into individual components of the solute (expanding the solute). This requires *E* be *added* to the system, therefore **endothermic**. The magnitude of the value is high in ionic and polar solutes, low in nonpolar solutes.

$$\Delta H_{\text{solute}} = -\Delta H_{\text{lattice energy}}$$

Step 2 (ΔH_2)

- Overcoming IMFs in solvent to make room for the solute (expanding the solvent). Requires that *E* be added to the system, therefore **endothermic**. The magnitude of the value is high in polar solvents, low in nonpolar solvents.

Step 3 (ΔH_3)

- Interaction of solute and solvent to form the solution. Energy must be released here, else the solution would never form since nature always tends toward a lower energy state, therefore **exothermic**. The magnitude of this value is high in polar solute—polar solvent interactions, low in other types of interactions.

$$\Delta H_2 + \Delta H_3 = \text{enthalpy of hydration } (\Delta H_{\text{hyd}})$$

Enthalpy of hydration is more negative for small ions and highly charged ions.

Table 11.3 The Energy Terms for Various Types of Solutes and Solvents

	ΔH_1	ΔH_2	ΔH_3	ΔH_{soln}	Outcome
Polar solvent, polar solute	Large	Large	Large, negative	Small	Solution forms
Polar solvent, nonpolar solute	Small	Large	Small	Large, positive	No solution forms
Nonpolar solvent, nonpolar solute	Small	Small	Small	Small	Solution forms
Nonpolar solvent, polar solute	Large	Small	Small	Large, positive	No solution forms

- Some heats of solution are positive (endothermic). The reason that the solute dissolves is that the solution process greatly increases the entropy (disorder) which overrides the cost of the small positive heat of solution. This makes the process spontaneous. The solution process involves two factors; the change in heat and the change in entropy, and the relative magnitude of these two factors determine whether a solute dissolves in a solvent.
- Hot and cold packs:
These often consist of a heavy outer pouch containing water and a thin inner pouch containing a salt. A squeeze on the outer pouch breaks the inner pouch and the salt dissolves. Some hot packs use anhydrous CaCl_2 ($\Delta H_{\text{soln}} = -82.8 \text{ kJ/mol}$) whereas many cold packs use NH_4NO_3 ($\Delta H_{\text{soln}} = +25.7 \text{ kJ/mol}$). We discussed other hot packs earlier that function on the principle of a supersaturated solution crystallizing, releasing the heat of crystallization. Other hot packs contain iron filings and the process of rusting is sped up thus, producing heat energy.

Exercise 3

Differentiating Solvent Properties

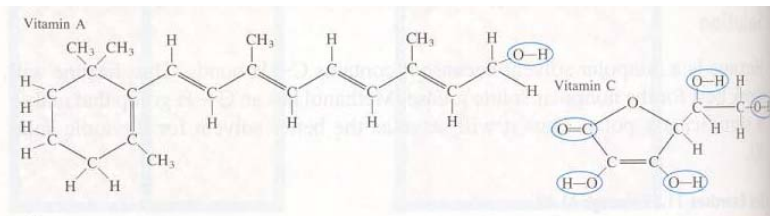
Decide whether liquid hexane (C_6H_{14}) or liquid methanol (CH_3OH) is the more appropriate solvent for the substances grease ($\text{C}_{20}\text{H}_{42}$) and potassium iodide (KI).

hexane → grease
methanol → KI

FACTORS AFFECTING SOLUBILITY

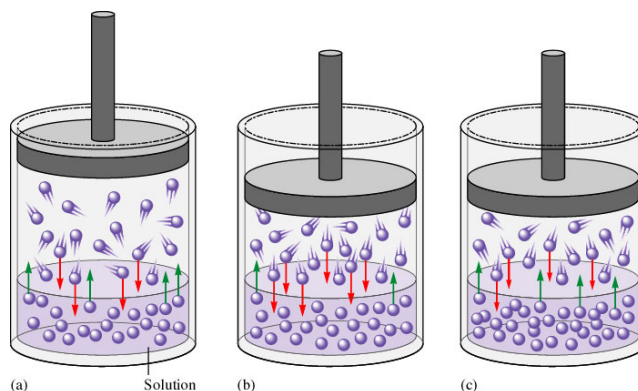
➤ Molecular Structure:

- Fat soluble vitamins, (A,D,E,K) –nonpolar (can be stored in fatty body tissues)
- Water soluble vitamins, (B&C) –polar (are not stored, wash away and must be consumed regularly)
- Hydrophobic— water fearing (nonpolar)
- Hydrophilic – water loving (polar)



➤ The Effect of Increasing Pressure

- The solubility of a gas increases with increasing pressure. Increasing pressure has very little effect on the solubility of liquids and solids. (carbonated beverages must be bottled at high pressures to ensure a high concentration of carbon dioxide in the liquid)
- Henry's Law— the amount of a gas dissolved in a solution is directly proportional to the pressure of the gas above the solution. Henry's Law is obeyed best for dilute solutions of gases that don't dissociate or react with the solvent.



$$\text{Henry's Law: } C = kP$$

P = partial pressure of the gaseous solute above the solution

k = constant (depends on the solution)

C = concentration of dissolved gas

Exercise 4

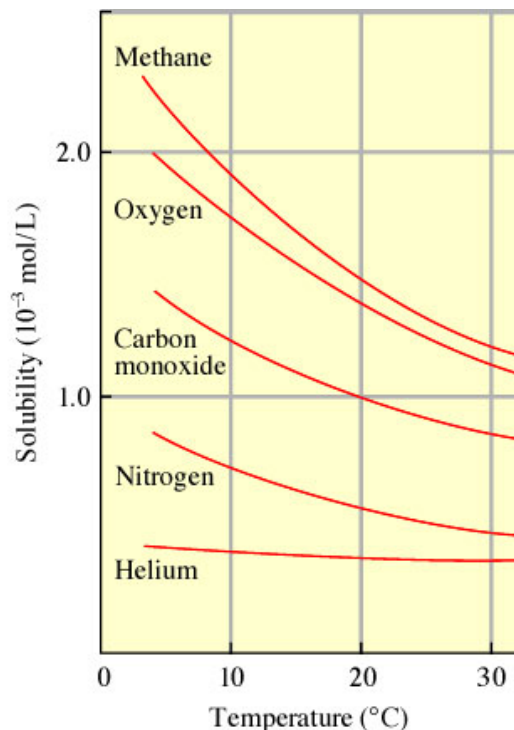
Calculations Using Henry's Law

A certain soft drink is bottled so that a bottle at 25°C contains CO₂ gas at a pressure of 5.0 atm over the liquid. Assuming that the partial pressure of CO₂ in the atmosphere is 4.0×10^{-4} atm, calculate the equilibrium concentrations of CO₂ in the soda both before and after the bottle is opened. The Henry's law constant for CO₂ in aqueous solution is 0.031 mol/L • atm at 25°C.

before = 0.16 mol/L
after = 1.2×10^{-5} mol/L

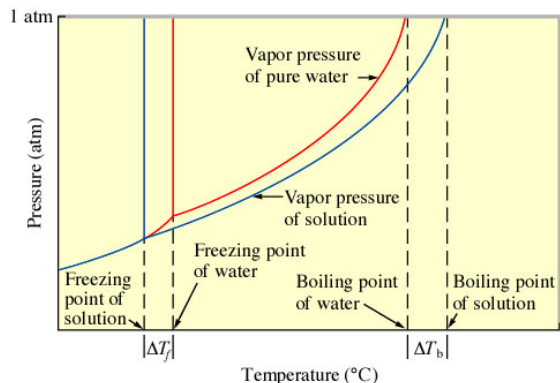
➤ The Effect of Increasing Temperature

- The amount of solute that will dissolve *usually* increases with increasing temperature since most solution formation is endothermic. Solubility generally increases with temperature if the solution process is endothermic ($\Delta H_{\text{soln}} > 0$). Solubility generally decreases with temperature if the solution process is exothermic ($\Delta H_{\text{soln}} < 0$). Potassium hydroxide, sodium hydroxide and sodium sulfate are three compounds that become less soluble as the temperature rises. This can be explained by LeChatelier's Principle.
- Remember, the *dissolving* of a solid occurs more rapidly with an increase in temperature, but the *amount of solid* may increase or decrease with an increase in temperature. It is very difficult to predict what this solubility may be—experimental evidence is the only sure way.
- The solubility of a gas in water **always decreases** with increasing temperature.
- There are all types of environmental issues involved with the solubility of a gas at higher temperatures. Thermal pollution – water being returned to its natural source at a higher ambient temperature has killed much wildlife as less oxygen is dissolved in the water. Boiler scale is another problem. This is where a coating builds up on the walls of containers such as industrial boilers and pipes causing inefficient heat transfer and blockage.



COLLIGATIVE PROPERTIES

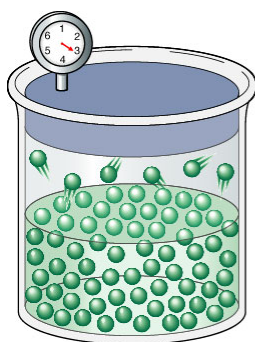
Colligative Properties— properties that depend on the number of dissolved particles; not on the identity of the particle. Intermolecular forces of the solvent are interrupted when the solute is added. This changes the properties of the solvent. These properties include: vapor pressure lowering, boiling-point elevation, freezing-point depression, and osmotic pressure.



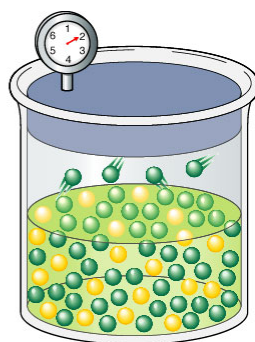
- **Vapor Pressure Lowering**— The presence of a nonvolatile solute lowers the vapor pressure of a solvent. This is because the dissolved nonvolatile solute decreases the number of solvent molecules per unit volume. (Nonvolatile solute dilutes the solution and interferes with the IMFs of the solvent). There are fewer solvent molecules on the surface to escape. This can be mathematically expressed by

$$\text{Raoult's Law: } P_{\text{solution}} = (\chi_{\text{solvent}}) (P^{\circ}_{\text{solvent}})$$

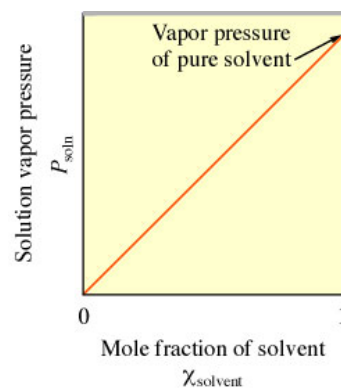
- P_{solution} = observed vapor pressure of the solvent in the solution
- χ_{solvent} = mole fraction of solvent
- $P^{\circ}_{\text{solvent}}$ = vapor pressure of the pure solvent
- i = van't Hoff factor (moles of electrolyte must be multiplied by this)
number of moles particles in solution/number of moles particles dissolved



Pure solvent



Solution with a nonvolatile solute



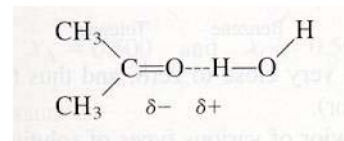
- The vapor pressure of a solution is directly proportional to the mole fraction of solvent present. If the solute ionizes the number of ions further affects (lowers) the vapor pressure. The moles of solute must be multiplied by the number of ions the given solute breaks into. For instance, if we had 1 mole of NaCl as the solute, we would use 2 moles of particles for our mole fraction calculations, AlCl₃ would yield a van't Hoff factor of 4 and so on.
- For nonelectrolytes, $i = 1$. For electrolytes, i = the number of particles formed when one formula unit of the solute dissolves in the solvent.
- The experimental value of i is often less than the expected value of i because of a phenomenon called “**ion pairing**”. Especially in concentrated solutions, oppositely charged ions can pair up and thus, we have fewer particles than expected. Conversely, nonelectrolytes may also pair up forming dimers (2 units bound together).

Table 11.6 Expected and Observed Values of the van't Hoff Factor for 0.05 m Solutions of Several Electrolytes

Electrolyte	<i>i</i> (expected)	<i>i</i> (observed)
NaCl	2.0	1.9
MgCl ₂	3.0	2.7
MgSO ₄	2.0	1.3
FeCl ₃	4.0	3.4
HCl	2.0	1.9
Glucose*	1.0	1.0

*A nonelectrolyte shown for comparison.

- An **ideal solution** is a solution that obeys Raoult's Law. There is no such thing. In very dilute solutions, Raoult's Law works fairly well. Solutions are most ideal when the solute and the solvent are very similar. If hydrogen bonding occurs between solute and solvent, vapor pressure is less than expected. We call this a *negative deviation from Raoult's law*. This can often be predicted when the enthalpy of solution formation is large and negative (highly exothermic)
 - A great example of this negative deviation is acetone and water.



EXAMPLE: Calculate the vapor pressure caused by the addition of 100.g of sucrose, C₁₂H₂₂O₁₁, to 1000.g of water if the vapor pressure of the pure water at 25°C is 23.8 torr.

$$100 \text{ g sucrose} \times \frac{1 \text{ mol sucrose}}{342.0 \text{ g sucrose}} = 0.292 \text{ mol}$$

$$1000 \text{ g water} \times \frac{1 \text{ mole water}}{18.0 \text{ g water}} = 55.6 \text{ mol water}$$

$$\chi_{\text{water}} = \frac{55.6}{0.292 + 55.6} = 0.995$$

$$P_{\text{soln}} = 0.995 \times 23.8 = \mathbf{23.7 \text{ torr}}$$

Exercise 5 Calculating the Vapor Pressure of a Solution

Calculate the expected vapor pressure at 25°C for a solution prepared by dissolving 158.0 g of common table sugar (sucrose, molar mass = 342.3 g/mol) in 643.5 cm³ of water. At 25°C, the density of water is 0.9971 g/cm³ and the vapor pressure is 23.76 torr.

$$= \mathbf{23.46 \text{ torr}}$$

Exercise 6 Calculating the Vapor Pressure of a Solution Containing Ionic Solute

Predict the vapor pressure of a solution prepared by mixing 35.0 g solid Na_2SO_4 (molar mass = 142 g/mol) with 175 g water at 25°C . The vapor pressure of pure water at 25°C is 23.76 torr.

= 22.1 torr

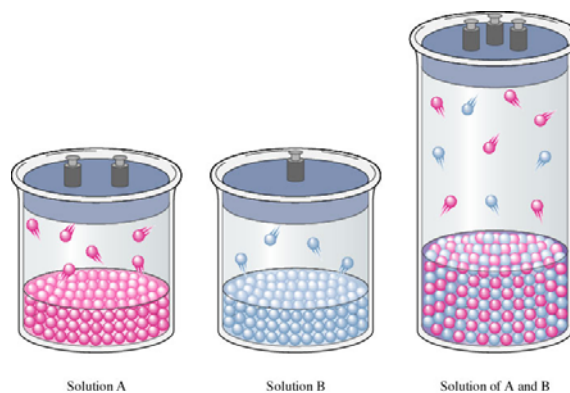
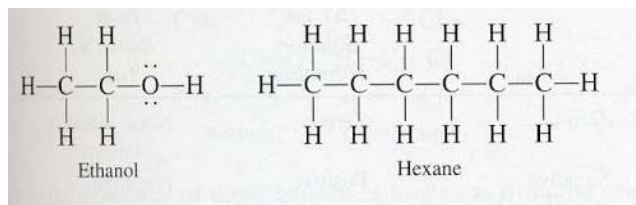
- We can determine the **molecular mass** of a solute by using the vapor pressure of a solution if the mass of the solute is known.

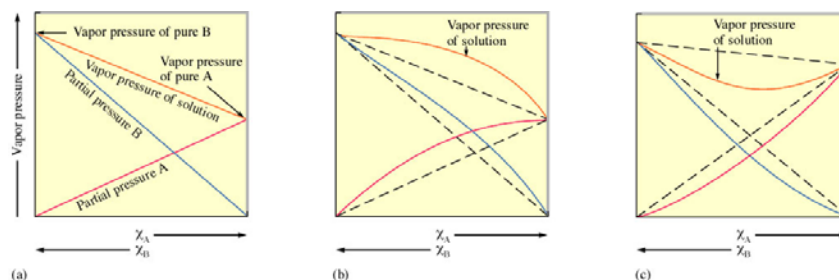
$$\chi_{\text{solvent}} = \frac{P_{\text{soln}}}{P_{\text{solvent}}^{\circ}}$$

$$\chi_{\text{solvent}} = \frac{n_{\text{solvent}}}{n_{\text{solvent}} + n_x}$$

solve for n_x and remember, $\frac{\text{g}}{\text{mol}} = MM$

- Solutions in which both solute and solvent are liquid and the liquids are volatile *do not* behave ideally. Both solute and solvent contribute to the vapor pressure. If the solute is more volatile than the solvent, the vapor pressure of the solution is higher than the vapor pressure of the solvent. In this case, the molecules have a higher tendency to escape than expected. We call this a *positive deviation from Raoult's law*. The enthalpy of solution for this type of deviation is positive. (endothermic)





$$P_{total} = P_A + P_B = \chi_A P_A^0 + \chi_B P_B^0 \quad (\text{same as Dalton's Law})$$

Exercise 7 Calculating the VP of a Solution Containing Two Liquids

A solution is prepared by mixing 5.81 g acetone (C_3H_6O , molar mass = 58.1 g/mol) and 11.9 g chloroform ($CHCl_3$, molar mass = 119.4 g/mol). At $35^\circ C$, this solution has a total vapor pressure of 260. torr. Is this an ideal solution? The vapor pressures of pure acetone and pure chloroform at $35^\circ C$ are 345 and 293 torr, respectively.

Not an ideal solution

- **Boiling-point Elevation**— Because vapor pressure is lowered by the addition of a nonvolatile solute, more E must be added to the system to achieve a vapor pressure equal to the atmospheric pressure (definition of normal BP), so the boiling point is increased.

$$\text{Boiling Point Elevation: } \Delta T_b = K_b \times m \times i$$

- K_b = molal boiling-point elevation constant (for water = $0.51^\circ C \text{ kg/mol}$)
 - i = van't Hoff factor
 - m = concentration in molality
 - ΔT = change in temperature
- **Freezing-point Depression**— Because vapor pressure is lowered by the addition of a nonvolatile solute, less E must be added to the system to achieve a temperature at which the vapor pressure of the solid and the liquid are equal (definition of FP or MP). If the vapor pressure of the liquid is lowered, the freezing point decreases. This is why $NaCl$ and $CaCl_2$ are used on icy roads and sidewalks to keep them from freezing in really cold weather.

WHY IS A SOLUTION'S FREEZING POINT DEPRESSED? Molecules cluster in order to freeze. They must be attracted to one another and have a spot in which to cluster. Solute molecules get in the way! The more ions in solution, the greater the effect on the freezing point and the boiling point. A *solution* does not have a sharply defined freezing point, a *solvent* does. Useful for separation purposes in fractional crystallization.

$$\text{Freezing Point Depression: } \Delta T = K_f \times m \times i$$

- K_f = molal freezing-point depression constant (for water = 1.86 °C kg/mol)
- add 6 qts. of antifreeze to 12 qts. cooling system in order to lower the FP to -34°F and raise the BP to $+226^\circ\text{F}$

➤ **Determining the molar mass (*MM*) of a solute using freezing-point depression or boiling-point elevation**

- solute concentration must be low ($0.10m$)
- disadvantage—compound must be nonvolatile and stable at the boiling point.
- still used widely
- remember that the units you seek are grams/mole!

EXAMPLE: Calculate the freezing point and boiling point of a solution of 100. g ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$) in 900. g of water.

Table 11.5 Molal Boiling-Point Elevation Constants (K_b) and Freezing-Point Depression Constants (K_f) for Several Solvents

<i>Solvent</i>	<i>Boiling Point</i> (°C)	K_b (°C · kg/mol)	<i>Freezing Point</i> (°C)	K_f (°C · kg/mol)
Water (H_2O)	100.0	0.51	0	1.86
Carbon tetrachloride (CCl_4)	76.5	5.03	-22.99	30.
Chloroform (CHCl_3)	61.2	3.63	-63.5	4.70
Benzene (C_6H_6)	80.1	2.53	5.5	5.12
Carbon disulfide (CS_2)	46.2	2.34	-111.5	3.83
Ethyl ether ($\text{C}_4\text{H}_{10}\text{O}$)	34.5	2.02	-116.2	1.79
Camphor ($\text{C}_{10}\text{H}_{16}\text{O}$)	208.0	5.95	179.8	40.

Exercise 8 Calculating the Molar Mass by Boiling-Point Elevation

A solution was prepared by dissolving 18.00 g glucose in 150.0 g water. The resulting solution was found to have a boiling point of 100.34°C. Calculate the molar mass of glucose. Glucose is a molecular solid that is present as individual molecules in solution.

= 180 g/mol

Exercise 9 Freezing-point Depression

What mass of ethylene glycol ($C_2H_6O_2$, molar mass = 62.1 g/mol), the main component of antifreeze, must be added to 10.0 L water to produce a solution for use in a car's radiator that freezes at $-10.0^\circ F$ ($-23.3^\circ C$)? Assume the density of water is exactly 1.00 g/mL.

= 7.76×10^3 g (or 7.76 kg)

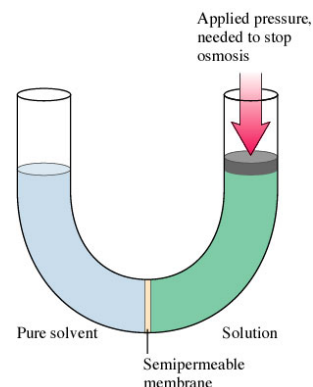
Exercise 10 Determining Molar Mass by Freezing-Point Depression

A chemist is trying to identify a human hormone, which controls metabolism, by determining its molar mass. A sample weighing 0.546 g was dissolved in 15.0 g benzene, and the freezing-point depression was determined to be 0.240°C. Calculate the molar mass of the hormone.

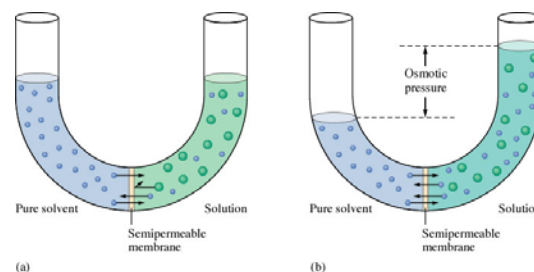
= 776 g/mol

➤ Osmotic Pressure: a special use of $PV=nRT$

- Semipermeable or selectively permeable membrane – membrane which allows solvent but not solute molecules to pass through (small molecules can pass but large ones cannot)
- Osmosis—a special case of diffusion; water travels along a concentration gradient.
 - Osmosis occurs when water molecules move through a semipermeable membrane from a region of lower solute concentration to a region of higher solute concentration. Is driven by the need nature has to establish an equilibrium.



- Osmotic pressure—(π) – the pressure that must be applied to a solution to prevent the net movement of water from solvent to solution (osmosis). The osmotic pressure of a solution is proportional to the number of solute particles in a given volume of solution, that is, to the molarity. The equation is similar to the ideal gas law since both relate the pressure of a system to its concentration and temperature.



$$\text{Osmotic Pressure: } \pi = MRTi \quad \text{or} \quad \pi = \left(\frac{n}{V}\right)RTi$$

π = osmotic pressure in atm
 M = molarity of the solution
 R = 0.08206 L • atm/mol • K
 T = temperature in Kelvin
 i = van't Hoff factor

- The use of osmotic pressure calculations for determining the molecular mass of an unknown substance is more accurate than the use of freezing-point depression or boiling point elevation data because a small concentration of solute produces a relatively large osmotic pressure.
- Ideal for measuring molar masses of large molecules of biological importance.

EXAMPLE: The concentration of hemoglobin in blood is roughly 15.0g/100.mL of solution. Assume that a solution contains 15.0g of hemoglobin dissolved in water to make 100.mL of solution and that the osmotic pressure of this solution is found to be 0.0500 atm at 25°C. What is the molecular mass of hemoglobin? (The osmotic pressure of a 1 m solution at 25°C is 24.45 atm)

$$\pi = \frac{nRT}{V} \text{ or } MM = \frac{dRT}{\pi}; \text{ so, } MM = \frac{\left(\frac{15}{0.1}\right)(0.08206)(298)}{0.050} = 73,400 \frac{\text{g}}{\text{mol}}$$

Exercise 11 Determining Molar Mass from Osmotic Pressure

To determine the molar mass of a certain protein, 1.00×10^{-3} g of it was dissolved in enough water to make 1.00 mL of solution. The osmotic pressure of this solution was found to be 1.12 torr at 25.0°C. Calculate the molar mass of the protein.

$$= 1.66 \times 10^4 \text{ g/mol}$$

Exercise 12 Isotonic Solutions

What concentration of sodium chloride in water is needed to produce an aqueous solution isotonic with blood ($\pi = 7.70$ atm at 25°C)?

$$= 0.158 \text{ M}$$

Exercise 13

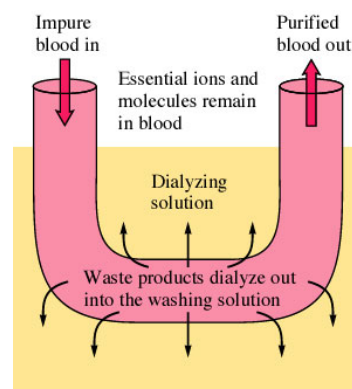
Osmotic Pressure

The observed osmotic pressure for a 0.10 M solution of $\text{Fe}(\text{NH}_4)_2(\text{SO}_4)_2$ at 25°C is 10.8 atm. Compare the expected and experimental values for i .

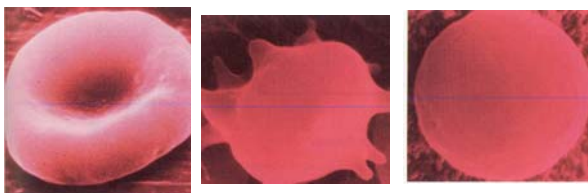
expected = 5
experimental = 4.4

➤ Applications of Osmosis:

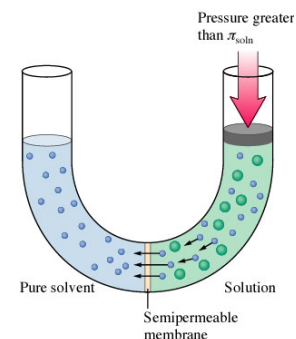
- **Dialysis**— a phenomenon in which a semipermeable membrane allows transfer of both solvent molecules and small solute molecules and ions.
- occurs in walls of most plant and animal cells
- kidney dialysis is one of most important applications – waste molecules move into the “wash” solution and filter the blood.



- **isotonic solution**—solutions that have the same osmotic pressure (Ex. IV fluids)
- **hypertonic**— solution has higher osmotic pressure (cells bathed in a hypertonic solution would shrivel—crenation) Treating the surface of food with salt causes this to happen to bacteria, thus protecting or preserving the food.
- **hypotonic** – solution has lower osmotic pressure (cells bathed in a hypotonic solution would burst—hemolysis)

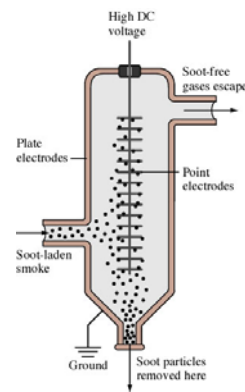


- **reverse osmosis**— the process occurring when the high external pressure on a solution causes a net flow of solvent through a semipermeable membrane from the solution to the solvent
- used in desalination (the membrane here acts as a “molecular filter” to remove solute particles) The need for this process will probably increase as the need for drinkable water increases.



➤ **Colloids**— (also called colloidal dispersions)

- Thomas Graham, 1860—albumin, starch, gelatin and glue diffuse only very slowly and could not be crystallized. He called these substances colloids.
- a suspension of tiny particles in some medium
- The dispersed colloidal particles are larger than a simple molecule but small enough to remain distributed and not settle out.
- A colloidal particle has a diameter between 1 and 1000 nm and may contain many atoms, ions, or molecules.
- Because of their small particle size, colloids have an enormous total surface area.
- The particles stay suspended because of electrostatic repulsion.
- hydrophobic/hydrophilic ends
- Coagulation, destruction of a colloid, occurs by heating (particles collide so hard that they stick together) or by the addition of an electrolyte (neutralizes ion layers). This process is important in removal of soot from smoke so that air quality has improved somewhat in industrialized cities.
- **Tyndall effect**— the scattering of light by particles
 - Used to distinguish between a suspension and a true solution. A true solution has particles that are too small to scatter light.
- **Brownian motion**— a characteristic movement in which the particles change speed and direction erratically (solvent molecules collide with the colloidal particles).
 - **Suspensions** are temporary solutions. They will settle eventually—colloids will not do this.
 - **Solutions** are permanent. Particles are really small. Colloids lie in between solutions and suspensions!



▪ **Examples of Some Common Colloids:**

- **Foam**— colloidal dispersion of a gas dispersed in a liquid or solid (ex. Whipped cream and marshmallows)
- **Aerosol**— colloidal dispersion of a liquid or solid dispersed in a gas (ex. Fog and smoke)
- **Emulsion**— colloidal dispersion of a liquid dispersed in a solid or liquid (ex. Butter and milk)
- **Sol**— colloidal dispersion of a solid dispersed in a liquid or solid (ex. Paint or ruby)

Examples	Dispersing Medium	Dispersed Substance	Colloid Type
Fog, aerosol sprays	Gas	Liquid	Aerosol
Smoke, airborne bacteria	Gas	Solid	Aerosol
Whipped cream, soap suds	Liquid	Gas	Foam
Milk, mayonnaise	Liquid	Liquid	Emulsion
Paint, clays, gelatin	Liquid	Solid	Sol
Marshmallow, polystyrene foam	Solid	Gas	Solid foam
Butter, cheese	Solid	Liquid	Solid emulsion
Ruby glass	Solid	Solid	Solid sol