

## Ch 12 [Solutions](#) (also covers dispersions, which are not solutions)

### Solution

- A solution is a uniform, homogeneous, single-phase mixture of ions and/or molecules.
- [Solutions](#) can be solid, liquid, or gas. [Examples](#) include not just aqueous solutions, but also solids dissolved together (like alloys), as well as gases dissolved together (like air).
- All substances are completely dissolved (mixed) to the ionic/molecular level.
- The substances are not aggregating or clumping together as particles or multiple molecules.

### Dispersion

- If there are particles in the mixture, and the particles are small and distributed uniformly so that the mixture appears homogeneous, the mixture is called a colloid, or colloidal dispersion.

### Solute (dissolved species)

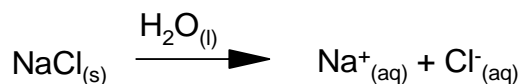
- If a solid or gas is dissolved within a liquid, that solid or gas is considered the solute.
- The solute is no longer a solid or gas when part of a liquid solution, it is dissolved within it.
- In all other cases, such as with two different liquids that are dissolved together, the solute is the component that is present in smaller quantity.

### Solvent (the substance which the solute is dissolved in)

- If a solid or gas is dissolved within a liquid, the liquid is considered the solvent.
- In all other cases, the solvent is the component present in greater quantity.

### Dissolving Processes

- Dissolving an ionic compound in water requires [dissociation](#) into ions, which is a chemical process.



- Sodium and chloride ions are the solutes in this aqueous solution, and water is the solvent.
- Sodium chloride is not the actual solute, because it has dissociated and is not truly present.
- Dissolving a completely covalent compound does not involve dissociation, so it is a physical process only. Covalent molecules, such as ammonia, are dissolved intact.  $\text{NH}_{3(g)} \rightarrow \text{NH}_{3(aq)}$ .

### Solubility

- Solubility is the maximum amount of a solute that will dissolve in a solvent.
- Solubility can be measured in either mol/L (M) or g/L (grams of solute per liter of solvent).

### Miscibility and Miscible Fluids

- The term miscibility describes the ability of liquids to dissolve in each other. The term is sometimes, though rarely, applied to solids and gases as well.
- [Miscible liquids](#) are those that always dissolve together completely in all proportions. They always form a single solution, no how much of each component is present.
- Ethyl alcohol ( $\text{CH}_3\text{CH}_2\text{OH}$ ) and water are miscible.
- They always dissolve together completely to form a single liquid-phase solution.

## Immiscible Liquids

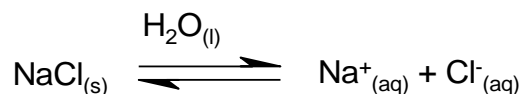
- [Immiscible liquids](#) do not dissolve together in any proportion (like oil and H<sub>2</sub>O).
- n-Hexyl alcohol only dissolves up to 0.6 g per 100 g of water. They are *nearly* immiscible together.
- Immiscible liquids of different densities typically form layers, with the lighter fluid (lower density) on top (oil floats on water).

## Partially Miscible

- [Partially-miscible liquids](#) dissolve together as a single solution until the solubility is reached.
- Any amount of a component in excess of the solubility will form a second liquid layer.
- n-Butyl alcohol will dissolve completely (not partially) in water up to 7.9 g per 100 g H<sub>2</sub>O. But any amount of the alcohol in excess of the solubility will not dissolve at all. The undissolved excess alcohol is less dense than the solution, and it will float on top.

## Dynamic Equilibrium

- Equilibrium, for a solution, is when concentrations within a solution are no longer changing because the quantity of solute dissolving (entering the solution) is the same as the quantity crystallizing (leaving the solution).



- An arrow is shown for each direction because both reactions occur simultaneously.
- At equilibrium, both reactions have the same rate, so the concentrations are not changing.

## LeChatelier's Principle

- When a system in equilibrium is disturbed by a change in T, P, or concentration, the system shifts its equilibrium to counteract this change.
- The presence of equilibrium implies a reversible reaction, which is a reaction that can proceed in both the forward and reverse directions simultaneously.

## Saturated

- Saturated means that the maximum amount of a solute is already dissolved in the solvent.
- The solute's saturated concentration is the same as its solubility.
- Even if the solute and solution are in contact with each other, no more solute dissolves, and the solute concentration still does not change.
- The system remains in equilibrium because forward and reverse rates are equal.

## Unsaturated

- Unsaturated means that less than the maximum amount is dissolved.
- More solute can dissolve into the solution.
- If the solute and solution are in contact, then the solute concentration can increase.
- The system is not yet at equilibrium if the solute and solution are in contact with each other.

## Supersaturated

- Supersaturated means that more than the maximum equilibrium amount is dissolved. The situation seems illogical, but it is not impossible.
- Solubility [changes with temperature](#). Usually it increases at higher T. A saturated solution can be heated so that more solute can be dissolved.
- The hot saturated solution can be cooled slowly without losing the additional solute. The now supersaturated solution is somewhat stable with the excess amount still dissolved.
- The solute needs solid particles present to crystallize onto before it can leave the solution.
- If the supersaturated solution comes in contact with the solid particles, then the solute will be able to leave the solution. The solute will continue to separate and become solid until equilibrium is reached and the solution is no longer [supersaturated](#).

## Solubility of Gases

- All gases are miscible together.
- Solubility of gases does not depend on intermolecular forces because the molecules are much farther apart in gas phase.
- As a result of the distance between the molecules, their attractive forces are not strong enough for any separation from the solution.

## Solubility of Molecular Liquids (Like Dissolves Like)

- Polar liquids have permanent  $\delta^+$  and  $\delta^-$  dipoles on each molecule.
- Dipoles are attracted to their opposite charge. This holds these molecules together strongly.
- Hydrogen bonding is an especially strong type of polar intermolecular force. It exists between a  $\delta^+$  H and a  $\delta^-$  O, N or F atom, for example between water and methanol.
- Nonpolar liquids, like the hydrocarbons hexane and octane, do not have permanent dipoles. They are held together by weaker instantaneous (random/temporary) dipoles only. These molecules dissolve together though, because their intermolecular forces are similar.
- Unlike substances do not dissolve together. If polar and nonpolar liquids are placed together, the stronger polar forces will hold the polar molecules together strongly, and exclude the nonpolar molecules. So, the two substances will remain separated.

## Entropy

- Entropy is a measure of the natural tendency towards disorder.
- The amount of overall entropy always increases as a result of spontaneous reactions.
- Substances will mix and become disordered if no forces prevent them from doing so.
- If only London forces are present, then there are no large differences in attractions. The molecules will move freely together as long as T is much less than their boiling points.
- The same is true of polar molecules if all intermolecular forces are similar. For example,  $\text{NH}_3$  and  $\text{H}_2\text{O}$  are both very polar and both have hydrogen bonding. As a result, they are completely miscible.

## Aqueous Ionic Solutions

- Hydration is the attraction of ions for dipoles on H<sub>2</sub>O. It is the result of ion-dipole forces. The δ<sup>+</sup> H atom is attracted to anions (-), while the δ<sup>-</sup> O atom is attracted to cations (+).
- Lattice energy is the attraction of oppositely charged ions to each other to form the ionic crystal. It is the result of ionic bonds.
- Ionic solids dissolve when hydration energy released is greater than lattice energy absorbed.
- In general, compounds with singly-charged ions, like Na<sup>+</sup> and K<sup>+</sup>, are very soluble in water because they have relatively small lattice energies.
- As more solute is dissolved, fewer water molecules are available for bonding to the ions, so the overall hydration energy that is available decreases.
- Hydration decreases as more solute is dissolved until it is the same as the lattice energy. At that point, no more solute can dissolve. The system is then saturated and at equilibrium.

## Effect of Temperature on Aqueous Solubility

- Gases are less soluble in water at higher T because vapor pressures increase as T increases. Higher vapor pressure results in more molecules leaving the liquid solution for the gas phase.
- Lower T makes gases less likely to vaporize and leave the solution, so they remain dissolved.
- Ionic salts generally, though not necessarily, become more soluble as T increases. This is because higher T generally increases the hydration more than the lattice energy.
- There are some exceptions, though. CaSO<sub>4</sub>, for instance, becomes less soluble at higher T, because higher T increases its lattice energy more so than its hydration.

## Heat of Solution

- Heat is released or absorbed when ionic substances are dissolved.
- Dissolving ionic substances is a chemical change.
- As a result of this change, the lattice energy is absorbed to separate the ions, and the energy from hydration is released when the ions bond to water molecules.
- The overall chemical change can be exothermic (increases T) or endothermic (decreases T).
- NaOH and H<sub>2</sub>SO<sub>4</sub> each create heat when dissolved in water. Is this exothermic or endothermic? What happens to T?
- Slowly add acid to water, so that water can absorb heat without boiling (or spattering).
- NH<sub>4</sub>NO<sub>3</sub> removes heat when dissolved and is used in cold packs.

## Effect of Partial Pressure on Solubility of a Gas in a Liquid

- Solubilities of solids and nonvolatile liquids are generally not affected by pressure.
- Solubility of a gas, though, is determined by the pressure it exerts. Henry's Law states that the solubility (S) of a gas in a liquid is proportional to the partial pressure (P) of the gas.
- $S = k_H P_{\text{partial}}$  and  $k_H = \frac{S_1}{P_1} = \frac{S_2}{P_2}$  where  $k_H$  is the Henry's law constant for the system.
- We can rearrange the equation so that  $S_2 = \frac{S_1 P_2}{P_1}$ , and then it is not necessary to know  $k_H$ .

Example 12.01  $S_2 = \frac{S_1 P_2}{P_1} = \frac{(0.759 \frac{\text{g}}{\text{mL}} \text{ of CO}_2 \text{ in H}_2\text{O})(2 \text{ atm of CO}_2)}{(1 \text{ atm of CO}_2)} = 1.52 \text{ g of CO}_2 \text{ per mL of H}_2\text{O}$

Concentration – the amount of solute dissolved in a given amount of solvent or solution

Molarity (M) is moles of solute per liter of solution

- [Molarity](#) depends on solution volume. It cannot be directly related to the quantity of solvent unless the molar volumes of both solvent and solute are known (L/mol of each component).
- $0.30 \text{ mol solute} / 2.0 \text{ L solution} = 0.15 \text{ mol/L} = 0.15 \text{ M}$  (See [pages 5-6](#) for more examples.)

Mass Percentage       $\text{Mass \%} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100 \%$

- Denominator is not solvent mass alone, but the mass of solvent can be calculated directly.  
(Mass of Solution) = (Mass of Solute) + (Mass of Solvent)

Example 12.02 Determine Masses Needed for 254.0 g of Solution with 3.50 % Solute by Mass

- $3.50 \div 100 = \frac{0.0350 \text{ g Solute}}{1.0000 \text{ g Solution}} \quad \frac{0.0350 \text{ g Solute}}{1.0000 \text{ g Solution}} \times 254.0 \text{ g Solution} = 8.89 \text{ g Solute}$
- $1.0000 - 0.0350 = \frac{0.9650 \text{ g Solvent}}{1.0000 \text{ g Solution}} \quad \frac{0.9650 \text{ g Solvent}}{1.0000 \text{ g Solution}} \times 254.0 \text{ g Solution} = 245.1 \text{ g Solvent}$

molality       $c_m = \frac{\text{Moles of Solute}}{\text{Mass of Solvent in kg}}$

- [molality](#) ( $c_m$ ) does not depend on volume at all
- molality relates the quantity of solute (moles) directly to the quantity of solvent (kg)
- molality is used to determine the colligative properties on the pages which follow

Example 12.03 Determine molality for 8.67 g of fructose dissolved in 25.4 g of water

- moles fructose =  $\frac{8.67 \text{ g}}{180.16 \frac{\text{g}}{\text{mol}}} = 0.0481 \text{ mol}$       kg water =  $\frac{25.4 \text{ g}}{1000 \frac{\text{g}}{\text{kg}}} = 0.0254 \text{ kg}$
- Molality is moles of solute divided by kg of solvent:       $c_m = \frac{0.0481 \text{ mol fru}}{0.0254 \text{ kg water}} = 1.89 \text{ molal}$

- Mole Fraction       $X_A = \frac{\text{Moles of A}}{\text{Total Moles of Solution}}$

- [Mole fraction](#) is the moles of solute (A) divided by the total moles of solution.
- Mole fraction does not depend on either mass or volume of the components.
- The sum of all the mole fractions in a solution is always one.

Example 12.04 Determine mole fraction for 8.67 g of fructose dissolved in 25.4 g of water

- moles of fructose =  $\frac{8.67 \text{ g}}{180.2 \frac{\text{g}}{\text{mol}}} = 0.0481 \text{ mol}$       moles of water =  $\frac{25.4 \text{ g}}{18.0 \frac{\text{g}}{\text{mol}}} = 1.41 \text{ mol}$
- $X_{\text{Fru}} = \frac{0.0481 \text{ moles of fructose}}{(0.0481 + 1.41 \text{ total moles of solution})} = \frac{0.0481 \text{ mol fructose}}{1.4581 \text{ mol solution}} = 0.0330 \frac{\text{mol fructose}}{\text{mol solution}}$
- $X_{\text{Water}} = \frac{1.41 \text{ moles of water}}{1.4581 \text{ total moles of solution}} = 0.967 \frac{\text{mol water}}{\text{mol solution}}$
- $X_{\text{Water}} + X_{\text{Fru}} = 0.967 + 0.0330 = 1.000$  (See [page 4](#) for more examples.)

Example 12.08 Convert Molarity (M = mol/L) of 1.832 M NaCl into molality (m = mol/kg)

- Use M and molar mass to find solute g/L  $\left(\frac{1.832 \text{ mol}}{\text{L}}\right)\left(\frac{58.44 \text{ g}}{\text{mol}}\right) = 107.1 \frac{\text{g}}{\text{L}}$  of solute
- Find solution g/L from density  $\left(\frac{1.0707 \text{ g}}{\text{ml}}\right)\left(\frac{1000 \text{ ml}}{1 \text{ L}}\right) = 1070.7 \frac{\text{g}}{\text{L}}$  of solution
- Solvent g/L is solution g/L minus solute g/L  $1070.7 \frac{\text{g}}{\text{L}} - 107.1 \frac{\text{g}}{\text{L}} = 963.6 \frac{\text{g}}{\text{L}}$  of water
- Convert solvent g/L to kg/L  $\left(\frac{963.6 \text{ g}}{\text{L}}\right)\left(\frac{1 \text{ kg}}{1000 \text{ g}}\right) = 0.9636 \frac{\text{kg}}{\text{L}}$  of water
- Divide solute mol/L by solvent kg/L  $\left(\frac{1.832 \text{ mol}}{\text{L}}\right) \div \left(\frac{0.9636 \text{ kg}}{\text{L}}\right) = 1.901 \frac{\text{mol solute}}{\text{kg of water}}$

Vapor Pressure is caused by molecules leaving the liquid solution for the gas phase.

- [Raoult's Law](#) states that the partial pressure ( $P_A$ ) of a substance's vapor above its solution equals its pure liquid's vapor pressure ( $P_A^0$ ) times its mole fraction ( $X_A$ ).  $P_A = (P_A^0)(X_A)$
- Dalton's Law states that total pressure is the sum of partial pressures.  $P_{\text{Total}} = P_A + P_B$
- When adding a volatile solute to a volatile solvent, we can combine the equations.

$$P_{\text{Total}} = (P_A^0)(X_A) + (P_B^0)(X_B)$$

$$P_{\text{Total}} = (P_A^0)(X_A) + (P_B^0)(1 - X_A)$$

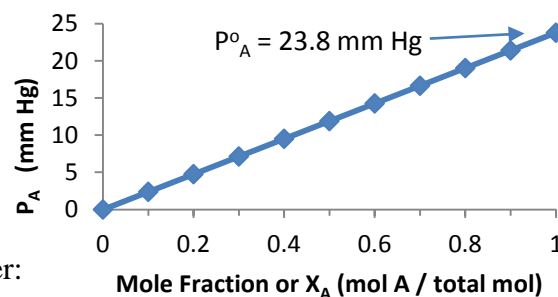
- If we add a nonvolatile solute instead, we still have  $X_A + X_B = 1$ , but  $P_B^0 = 0$ .

$$P_{\text{Total}} = P_A = (P_A^0)(X_A)$$

$$-\Delta P = P_A^0 - P_A = (P_A^0)(1 - X_A) = (P_A^0)(X_B)$$

- If a nonvolatile component (like fructose) is added to water:

$$P_{\text{Fructose}}^0 = 0 \quad \text{and} \quad P_{\text{Total}} = P_{\text{Water}} = (P_{\text{Water}}^0)(X_{\text{Water}})$$



Example 12.09 Calculate the Vapor Pressure of a Fructose Solution with Raoult's Law

- The aqueous solution from example 12.04 is at 25 °C, where  $P_{\text{Water}}^0 = 23.8 \text{ mmHg}$ .
- Fructose is nonvolatile ( $P_{\text{Fructose}}^0 = 0$ ) and  $X_{\text{Water}} = 0.967$ .
- $P_{\text{Total}} = P_{\text{Fructose}} + P_{\text{Water}} = P_{\text{Water}} = (P_{\text{Water}}^0)(X_{\text{Water}}) = (23.8 \text{ mmHg})(0.967) = 23.0 \text{ mmHg}$

Boiling Point Elevation ( $\Delta T_b$ )

- A liquid boils at  $T_b$ , where its vapor pressure equals one atmosphere (ambient pressure).
- If solute B is added to solvent A at the  $T_b$  of A, and  $P_B^0$  is small, then  $P_{\text{Total}}$  decreases. Now,  $P_{\text{Total}} < P_A^0$  and  $P_A^0 = 1 \text{ atm}$ , so the solution does not boil.
- Since  $P_{\text{Total}}$  increases as T increases, a higher T is needed to boil the solution. This means  $T_b$  of solvent A is less than the new boiling point ( $T_b$  of the solution).
- So, the boiling point ( $T_b$ ) increases as solute B is added to solvent A. The solution  $T_b$  equals the solvent  $T_b$  plus  $\Delta T_b$ , that is solution  $T_b = \text{solvent } T_b + \Delta T_b$ .
- This result is called Boiling Point Elevation, which is a Colligative Property. A colligative property depends on concentration of solute only, not on solute's identity.
- For a nonvolatile molecular solute:  $\Delta T_b = k_b c_m$ . (The lab manual uses m rather than  $c_m$ .) The increase in  $T_b$  equals the BP elevation constant ( $k_b$ ) times solute molality ( $c_m$ ).

$$\Delta T_b = k_b c_m$$

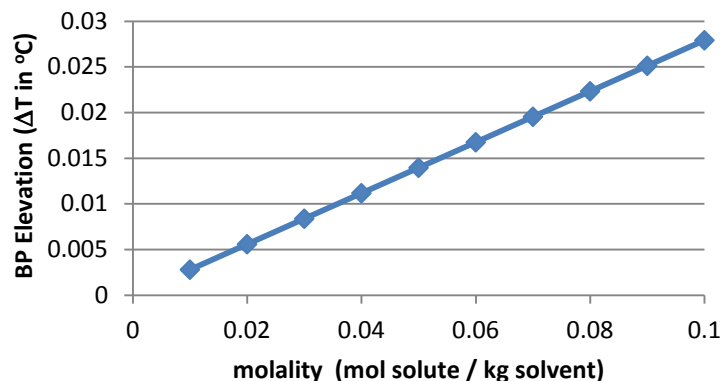
$\Delta T_b$  is the BP elevation

$c_m$  is molality in  $\frac{\text{mol}}{\text{kg}}$

$k_b$  is the BP constant

Units for  $k_b$  are  $\frac{^\circ\text{C}\cdot\text{kg}}{\text{mol}}$

$k_b$  is the slope of the line



### Freezing Point Depression ( $\Delta T_f$ )

- Freezing occurs when kinetic energy of molecules has been decreased by lowering T, so that the molecules can no longer overcome the attractive forces between them.
- Addition of solute increases the average distance between solvent molecules.
- This effectively decreases the solvent's intermolecular forces because these forces are a function of the distance between molecules.
- So,  $T_f$  of the solution is lower (with even less kinetic energy) than  $T_f$  of the solvent.
- The result is a colligative property called Freezing Point Depression.

Its equation is  $\Delta T_f = k_f c_m$ , where  $k_f$  is the FP depression constant.

- The solution  $T_f$  equals the solvent  $T_f$  minus  $\Delta T_f$ , that is solution  $T_f = \text{solvent } T_f - \Delta T_f$ .
- BP elevation and FP depression are both proportional to molality (m), not molarity (M). This works because molality uses solvent mass, not solution volume like molarity.
- The constants ( $k_b$  and  $k_f$ ) for water and other common solvents are listed in [Table 11.2](#) as well as [this table](#).

Example 12.10 Find boiling and freezing points for a 0.305 m solution of fructose in water

- $\Delta T_b = k_b c_m = (0.512 \frac{^\circ\text{C}\cdot\text{kg}}{\text{mol}})(0.305 \frac{\text{mol}}{\text{kg}}) = 0.156 \text{ }^\circ\text{C}$        $T_{b \text{ sol'n}} = T_{b \text{ water}} + \Delta T_b = 100.156 \text{ }^\circ\text{C}$
- $\Delta T_f = k_f c_m = (1.858 \frac{^\circ\text{C}\cdot\text{kg}}{\text{mol}})(0.305 \frac{\text{mol}}{\text{kg}}) = 0.567 \text{ }^\circ\text{C}$        $T_{f \text{ sol'n}} = T_{f \text{ water}} - \Delta T_f = -0.567 \text{ }^\circ\text{C}$

Example 12.12 Finding Molecular Mass from  $\Delta T_f$  (similar to colligative properties experiment)

- Suppose 59.5 mg of a solute is dissolved in 1.000 g of acetic acid and  $T_{f \text{ sol'n}} = 15.3 \text{ }^\circ\text{C}$
- molality =  $c_m = \frac{\Delta T_f}{k_f} = (16.6 \text{ }^\circ\text{C} - 15.3 \text{ }^\circ\text{C}) \left( \frac{1 \text{ mol/kg}}{3.9 \text{ }^\circ\text{C}} \right) = (1.3 \text{ }^\circ\text{C}) \left( \frac{1 \text{ mol/kg}}{3.9 \text{ }^\circ\text{C}} \right) = 0.33 \frac{\text{mol solute}}{\text{kg solvent}}$
- mass of solvent =  $(1.000 \text{ g solvent}) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 1.000 \times 10^{-3} \text{ kg solvent}$
- moles of solute =  $(\text{molality})(\text{solvent mass}) = \left( \frac{\text{mol solute}}{\text{kg solvent}} \right) (\text{kg solvent})$   
 $= (0.33 \frac{\text{mol solute}}{\text{kg solvent}})(1.000 \times 10^{-3} \text{ kg solvent}) = 3.3 \times 10^{-4} \text{ mol solute}$
- molar mass of solute =  $\frac{(59.5 \text{ mg}) \left( \frac{1 \text{ g}}{10^3 \text{ mg}} \right)}{3.3 \times 10^{-4} \text{ mol}} = \frac{5.95 \times 10^{-2} \text{ g}}{3.3 \times 10^{-4} \text{ mol}} = 180 \text{ g/mol} = 180 \text{ amu/molecule}$

## Boiling and Freezing Points of Ionic Solutions

- BP and FP are affected by all ions present, because ionic compounds dissociate.
- Each ion is a separate solute with its own molality, and each ion affects both BP and FP.
- The [Van't Hoff factor](#) ( $i$ ) is the total number of ions per formula unit for compound.  
$$\Delta T_b = ik_b c_m \quad \text{and} \quad \Delta T_f = ik_f c_m$$
- For  $\text{CaCl}_2$ , we have one  $\text{Ca}^{+2}$  ion and two  $\text{Cl}^{-1}$  ions, which gives us  $i = 1 + 2 = 3 \frac{\text{moles of ions}}{\text{mole CaCl}_2}$ .  
$$\Delta T_b = 3k_b c_m \quad \text{and} \quad \Delta T_f = 3k_f c_m$$
 where  $c_m$  is the molality of  $\text{CaCl}_2$  or  $\frac{\text{mol CaCl}_2}{\text{kg of solvent}}$
- We can assume that  $i$  is an exact integer in these equations, although more accurate equations for  $\Delta T$  also include an ionic activity coefficient ( $\alpha$ ) that is slightly less than 1.

## Osmosis

- Suppose two solutions are separated by a membrane which is permeable only to the solvent molecules, and not to the solute molecules.  
Osmosis is the solvent flow which acts to equalize solute concentrations.
- The direction of solvent flow will be from the side with lower solute concentration to the side with higher solute concentration.  
This will eventually cause the two concentrations to become equal.
- Osmosis occurs because the solute molecules block the membrane, and this prevents reverse flow of solvent away from the side with higher solute concentration.

## Osmotic Pressure

- [Osmotic pressure](#) is a colligative property of two solutions separated by a membrane.
- It is the pressure needed on the more concentrated side to stop the solvent flow into it from the fluid with lower solute concentration.  
$$\Pi = MRT$$
, where  $\Pi$  is osmotic pressure and  $M$  is molarity
- The equation uses molarity because it is derived from the ideal gas law as  $\Pi V = nRT$ .  
This rearranges to  $\Pi = (n/V)RT$ , and  $n/V$  is equivalent to  $M$  (mol/liter).
- For an ionic solution,  $\Pi = iMRT$ , where  $i$  is the Van't Hoff factor discussed previously.

### Example 12.13 Determining Osmotic Pressure of 1.000 g fructose in 0.2250 L Aqueous Solution

- $(1.000 \text{ g}) \left( \frac{1 \text{ mol}}{180.2 \text{ g}} \right) = 5.549 \times 10^{-3} \text{ mol}$
- $M = (5.549 \times 10^{-3} \text{ mol}) \div (0.2250 \text{ L}) = 2.466 \times 10^{-2} \frac{\text{mol}}{\text{L}}$
- $\Pi = MRT = (2.466 \times 10^{-2} \frac{\text{mol}}{\text{L}}) (0.082057 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}) (298 \text{ K}) = 0.6030 \text{ atm}$

## Reverse Osmosis

- If the concentrated side exerts a higher pressure ( $P > \Pi$ ), then the flow [reverses](#).
- Solvent then flows from higher solute concentration to lower solute concentration.
- This is called reverse osmosis.
- Reverse osmosis is a common method of purifying water, and is often used for bottled drinking waters and water fill stations.



## Colloids

- [Colloids](#) are dispersions of particles in a bulk medium.
- Colloids are not solutions because particles contain multiple molecules, and are therefore not dissolved.
- The particles are called the dispersed phase, the bulk medium is the continuous phase.
- Most [types of colloids](#) have familiar names.
- An aerosol is composed of liquid droplets or solid particles dispersed in a gas, like hair spray and smoke.
- An emulsion is composed of liquid droplets dispersed in another liquid, like oil and vinegar salad dressing.
- A sol is composed of solid particles dispersed in a liquid, like silt in a river.
- A gel is composed of liquid particles dispersed in a solid, like toothpaste.
- A foam is composed of gas particles dispersed in either a solid or a liquid, like styrofoam and beer foam.

Colloids can be either hydrophilic or hydrophobic when water is the continuous phase.

- Hydrophilic is a strong attraction between dispersed and continuous phases. One example would be a protein is dispersed in water, where both components have H-bonding. The colloid is stable because the protein molecules have greater attractive forces towards each other than they do towards the water molecules.
- Hydrophobic is a lack of attraction between phases. One example would small particles of an insoluble ionic salt formed by a reaction in water. The product is a type of sol containing the dispersed solid particles of a precipitate.

## Coagulation

- Coagulation is when dispersed particles aggregate, then separate from the continuous phase.
- This causes the colloid to break down, and the particles are no longer dispersed.

## Micelles and Association Colloids

- A [micelle](#) is a dispersed particle in an association colloid.
- A micelle's outside layer is hydrophilic. That is, the outer layer likes water because its atoms have ionic charges.
- Inside of this charged outer layer, the remainder of the micelle is hydrophobic. That is, the inner portion does not like water because its atomic bonds are nonpolar.
- A micelle is composed of molecules, like soaps (which are fatty acid salts), that have both hydrophilic (ionic) and hydrophobic (nonpolar) ends.
- Nonpolar dirt becomes trapped inside of the micelle, where it is nonpolar and hydrophobic.
- The hydrophilic outside layer allows causes the micelle particles to be dispersed in water, even though the dirt itself is actually insoluble in water.