

Chapter 12 SOLUTIONS

These Notes are to SUPPLEMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests! To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test!

The author is providing these notes as an addition to the students reading the text book and listening to the lecture. Although the author tries to keep errors to a minimum, the student is responsible for correcting any errors in these notes.

12.1 Solution: a homogeneous mixture of two or more substances consisting of ions or molecules

Colloid appears to be homogeneous, but consists of comparatively large particles of one substance dispersed throughout another substance or solution.

Solutions may be Gases, Liquids or Solids

Solute: for a gas or solid dissolved in a liquid, the gas or solid is the solute

Else it is the component in the smaller amount

Solvent: for a gas or solid dissolved in a liquid, it is the liquid, else the component in the greater amount.

12.1a Gaseous Solutions: Nonreactive gases can mix in all proportions to give a gaseous solution. Air is a mixture of Nitrogen, Oxygen and trace amounts of other gases.

Miscible Fluid: fluids that mix with or dissolve in each other in all proportions

Immiscible: do not mix, but form two layers,

12.1b Liquid Solutions: are a liquid solution obtained by dissolving a gas, liquid or solid to form a liquid.

Gas / Liquid: dissolve carbon dioxide in water to form carbonic acid

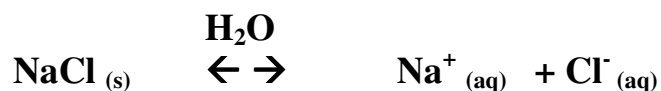
Liq / Liq: dissolve acetone in water, or alcohol in water

Solid / liq: dissolve sodium chloride in water to form the ocean

Solid / Solid: mix potassium and sodium metals to form a liquid mixture

12.1c Solid Solutions (alloys) Dental Fillings = mercury / silver alloy, Brass = copper and zinc, bronze is copper and tin, pewter is zinc and tin.

12.2 Solubility: 36.0 g of NaCl will dissolve in 100 ml of water at 20° C. It's solubility is 36.0 g / 100 ml. If you put 40.0 g of NaCl into 100 ml of water, 36.0 g will dissolve and the rest will remain on the bottom.

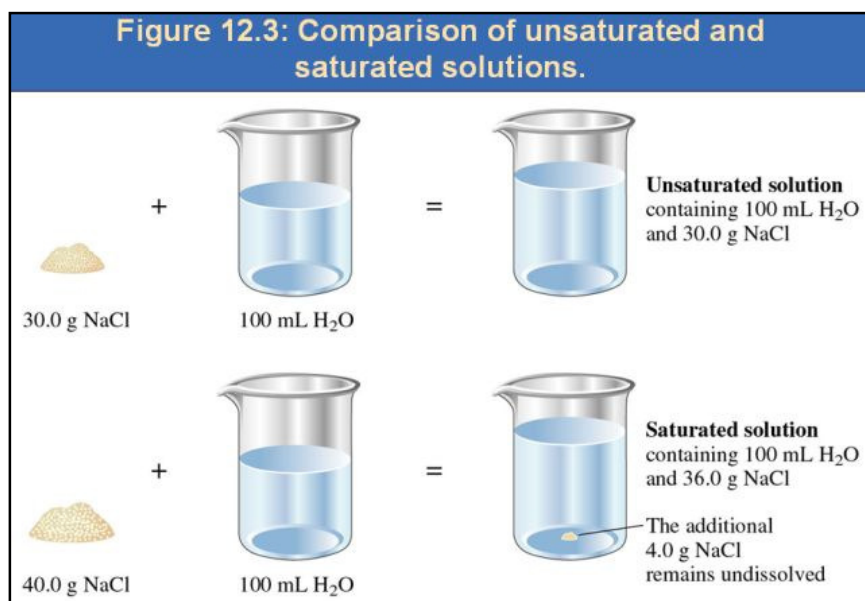


Sodium and Chlorine ions will come off of the salt on the bottom into solution until the maximum concentration is reached. At the same time the ions are going from the solid crystal to the ion form, the reverse is happening – ions are recrystallizing into the solid form.

1. Saturated Solution: one that is in equilibrium with respect to a given dissolved substance.

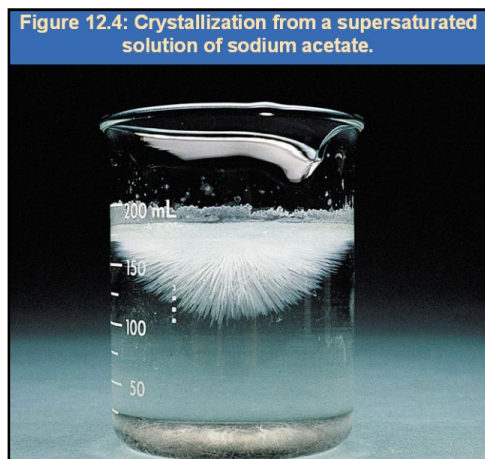
Solubility is the amount that dissolves in a given quantity of water at a given temperature to give a saturated solution.

2. Unsaturated Solution is one that is not in equilibrium with respect to a given dissolved substance and in which more of the substance can dissolve. If you mix 30.0 g of NaCl in 100 ml of water you will get an Unsaturated Solution.



3. Supersaturated Solution: is a solution that contains more dissolved substances than a saturated solution does.

Sodium Thiosulfate dissolves in water at 50 g / 100 ml at room temperature. At 100 °C, 231 g will dissolved in 100 ml of water. If you heat some water to 100 °C and add 230 g, it will all dissolved. If you let the solution cool back to room temperature without being disturbed, you will have a supersaturated solution. If you then add one crystal of sodium thiosulfate, or just scratch the container to create a “Seed Crystal”, crystals of sodium thiosulfate will rapidly form and the solution concentration will be reduce to the 50 g / 100 ml. Supersaturated Solutions ARE UNSTABLE.

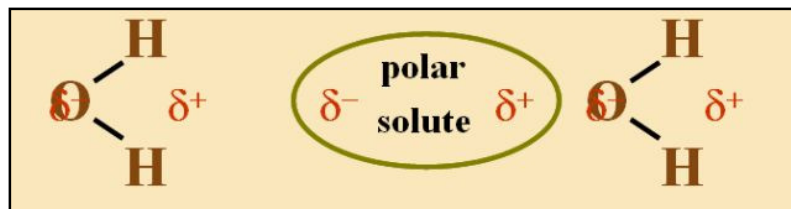


Factors in Solubility: “Like dissolves Like”

Polar solvents dissolve polar compounds, non-polar solvents dissolve non-polar compounds.

Polar: NaCl dissolves in water.

Non-Polar: Oil dissolves in gasoline.



Polar dissolves polar due to dipole-dipole interactions

Non-Polar dissolves non-polar due to the weak London forces.

Molecular Solutions: one gas dissolves in another gas. Gases mix, therefore gases are miscible.

Oil dissolves in Gasoline just as the organic heptane C_7H_{16} will dissolve in octane C_8H_{10} . These Organics are straight chains and the intermolecular attraction is due to London Forces.

Octane [an oil] does not dissolve in water because the hydrogen bonding of the water needs to break and the London Forces are not strong enough to overcome breaking the H-bonds.

As an example, as the **chain length** of the following organic alcohols gets larger, the **solubility in water decreases**. The compound has more organic [London Forces] structure rather than Hydroxyl [-OH] and the London Forces overcome the Hydrogen Bonding and the water solubility goes down:

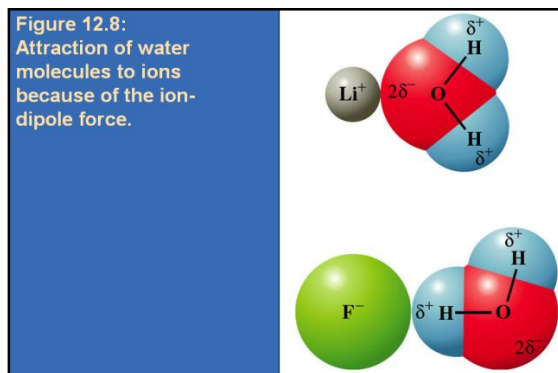
Solubilities of Alcohols in Water		
Name	Formula	Solubility in H ₂ O (g/100 g H ₂ O at 20°C)
Methanol	CH ₃ OH	Miscible
Ethanol	CH ₃ CH ₂ OH	Miscible
1-Propanol	CH ₃ CH ₂ CH ₂ OH	Miscible
1-Butanol	CH ₃ CH ₂ CH ₂ CH ₂ OH	7.9
1-Pentanol	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ OH	2.7
1-Hexanol	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ OH	0.6

Ionic Solutions: Ionic compounds differ markedly in their solubility in water.

NaCl dissolves to 36 g / 100 ml of water, $Ca_3(PO_4)_2$ only to 0.002 g / 100 ml of water.

The solubility difference is usually due to the difference between the attraction of the ions in the crystal and the attractions between the ion and water.

1. Ion / Water attraction is due to ion – dipole forces. Water is polar and will orient as such, the negative Oxygen will orient towards the Na^+ of NaCl , Li^+ with the Oxygen and F^- with the Hydrogen.



Hydration is the attraction for water molecules. See Fig 12.10, p 485

The energy of hydration depends on the ions charge. The smaller the ion, the larger is the relative charge and the more it is attracted towards water and is more soluble. The energy of hydration also increases with charge so Mg^{+2} has greater hydration energy than Na^{+1} .

2. Lattice Energy is the energy holding the ions of a crystal together.

Lattice Energy depends on the charges of the ions and the distance between them. The greater the charge, the greater the lattice energy and the less the solubility. So X^{+2} will have a greater lattice energy than X^{+1} and be less soluble. Na^+ is more soluble than Mg^{+2} .

Lattice Energy is proportional to the distance between the ions. The greater the distance between the ions, the greater the solubility. In going down the periodic table: $\text{Mg}(\text{OH})_2$, $\text{Ca}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$, $\text{Ba}(\text{OH})_2$, the radius increases so the lattice energy decreases so the solubility increases in going down the column of hydroxides.

To be soluble, the Hydration energy must be stronger than the crystals lattice energy.

12.3 Effect of Temperature on Solubility: Solubility depends on temperature.

Gases: most gases become less soluble in water at higher temperature.

Most ionic compounds increase solubility with increasing temperature and some decrease solubility.

Heat of Solution: Heat released or absorbed when an ionic substance is dissolved in water.

Exothermic: heat is released. Dissolve NaOH or CaCl_2 or MgSO_4 in water. [Hot Packs]

Endothermic: heat is absorbed. Dissolve ammonium nitrate [NH_4NO_3] in water. [Cold Packs]

Effect of Pressure on Solubility: The solubility of a gas is affected by pressure.

La Chatelier's Principle states when a system in equilibrium is disturbed by a change of temperature, pressure or concentration, the system shifts in equilibrium composition in a way that tends to counteract the change.

Example: Shake and release the cap of a soda bottle. The CO_2 will foam out. Throw an aerosol can into a fire – it'll explode from the increase in pressure.

Henry's Law: the solubility of a gas is directly proportional to the partial pressure of the gas above the solution:

$$S = k_H P \quad - \text{ or } - \quad \frac{S_1}{S_2} = \frac{P_1}{P_2}$$

S = Solubility, k_H = Henry's constant, P = partial pressure of the gas.

Example 12.1 Acetylene [HCCH] at 1.0 atm partial pressure is soluble in acetone at 27 g / liter. What is the solubility at 12 atm pp?

$$\frac{S_2}{27 \text{ g / L}} = \frac{12 \text{ atm}}{1.0 \text{ atm}} \quad S_2 = 324 \text{ g / L} = 3.2 \times 10^2 \text{ g / L}$$

Example 12.1b One Liter of water at 25 °C dissolves 0.0404 g of Oxygen with the Oxygen partial pressure of 1.00 atm. How much Oxygen is dissolved in 1 L of water at 159 mm Hg?

$$\frac{4.04 \times 10^{-2} \text{ g/L}}{S_2} = \frac{1.00 \text{ atm}}{159 \text{ mm Hg} / 760 \text{ mm Hg}}$$

$$S_2 = 8.45 \times 10^{-3} \text{ g/L}$$

Colligative property depends on the concentration of the solute molecules or ions, but not the chemical identity of the solute.

12.4 Concentration: of a solute is the amount of solute dissolved in a given quantity of solvent.

<u>Symbol</u>	<u>Name</u>	<u>Top</u>	<u>Bottom</u>	<u>Use</u>
M	Molarity	Moles of Solute	Liters of Solution	Titration, Lab Use
	Mass % of Solute	Mass of Solute * 100 %	Mass of Solution	
m	Molality	Moles of Solute	Kg of solvent	BP Raise, FP Lower
X_A	Mole Fraction	Moles of A	Total Moles of Solution	Partial Pressure

M = Molarity = Moles of Solute / Liters of Solution

Molarity relates to exactly how much of each compound you're dealing with – how many atoms of each.

Mass % of Solute = (Mass of Solute / Mass of Solution) * 100%

m = Molality = Moles of Solute / Kg of solvent

X_A = Mole Fraction = Moles of A / Total moles of solution

1. MOLARITY: is the moles of solute per Liter of solution.

What is the Molarity using 0.20 mol of ethylene glycol added to enough water to give 2.0 L.

$$M = \text{Molarity} = \frac{\text{Moles of Solute}}{\text{Liters of Solution}} = \frac{0.20 \text{ mol ethylene glycol}}{2.0 \text{ L Solution}} = 0.1 \text{ M ethylene glycol.}$$

2. MASS PERCENT: is the percentage by mass of solute contained in a solution.

2a. What is the Mass % of solute when 3.5 g of NaCl is dissolved in 96.5 g of water?

$$\text{Mass \% of Solute} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} * 100\% = \frac{3.5 \text{ g NaCl}}{[3.5 \text{ g NaCl} + 96.5 \text{ g H}_2\text{O}]} = 3.5 \%$$

CLASS: Determine all of the above concentrations for a solution made from 10.0 g of NaOH in 100 g H₂O.

Example 12.2 How would you make 425 g of a water solution containing 2.40 Mass % of Sodium Acetate [NaC₂H₃O₂]?

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

Rearrange to:

$$\begin{aligned} \text{Mass of Solute} &= \text{Mass of Solution} * \text{Mass \% of Solute} / 100\% \\ &= 425 \text{ g} * 2.40 \text{ Mass \%} / 100 \% = 10.2 \text{ g of Sodium Acetate} \end{aligned}$$

$$\text{The quantity of water} = 425 \text{ g [total weight]} - 10.2 \text{ g NaC}_2\text{H}_3\text{O}_2 = 415 \text{ g}$$

You'd prepare this solution by **adding 10.2 g NaC₂H₃O₂ to 415 g of water.**

Exercise 12.5 How would you prepare 35.0 g of 20.2% HCl by mass?

$$\begin{aligned} \text{Mass of Solute} &= \text{Mass of Solution} * \text{Mass \% of Solute} / 100\% \\ &= 35.0 \text{ g} * 0.202 = 7.07 \text{ g HCl} \end{aligned}$$

$$\text{The quantity of water} = 35.0 \text{ g total} - 7.07 \text{ g HCl} = 27.93 = 27.9 \text{ g water}$$

3. MOLALITY: is the moles of solute per kilogram of solvent

What is the molality of a solution of 0.20 mol of ethylene glycol dissolved in 2.0 kg of water?

$$\text{m} = \text{Molality} = \frac{\text{Moles of solute}}{\text{Kg of solvent}} = \frac{0.20 \text{ mol ethylene glycol}}{2.0 \text{ kg water}} = 0.10 \text{ m}$$

Example 12.3 What is molality of a solution of 5.67 g of glucose [C₆H₁₂O₆, mw = 180.2 g/mole] dissolved in 25.2 g of water.

$$\text{Moles of glucose} = \text{g} / \text{mw} = 5.67 \text{ g} / 180.2 \text{ g} / \text{mole} = 0.0315 \text{ mole} = 3.15 \times 10^{-2} \text{ mole}$$

$$\text{m} = \text{Molality} = \frac{\text{Moles of Solute}}{\text{Kg of solvent}} = \frac{3.15 \times 10^{-2} \text{ mole}}{2.52 \times 10^{-2} \text{ kg}} = 1.25 \text{ m}$$

Exercise 12.6 What is molality of a solution containing 35.6 g of toluene [C₆H₅CH₃] and 125 g of benzene?

$$\text{Moles of toluene} = \text{g} / \text{mw} = 35.6 \text{ g} / 92.14 \text{ g} / \text{mole} = 0.3863 = 0.386 \text{ moles toluene}$$

$$\text{m} = \text{Molality} = \frac{\text{Moles of Solute}}{\text{Kg of solvent}} = \frac{0.386 \text{ g toluene}}{0.125 \text{ kg benzene}} = 3.088 = 3.09 \text{ m toluene}$$

4. Mole Fraction: is the moles of a component divided by the total moles of solution [solute + solvent].

Example 12.4 What is the Mole Fraction of glucose in a solution of 5.67 g of glucose [$C_6H_{12}O_6$, mw = 180.2 g/mole] in 25.2 g of water?

$$\text{Moles of glucose} = g / \text{mw} = 5.67 \text{ g} / 180.2 \text{ g / mole} = 0.03146 = 0.0315 \text{ mole glucose}$$

$$\text{Moles of water} = g / \text{mw} = 25.2 \text{ g} / 18.02 \text{ g / mole} = 1.398 = 1.40 \text{ mole } H_2O$$

$$\text{Total moles} = 0.0315 \text{ mole} + 1.40 \text{ mole} = 1.43 \text{ mole}$$

$$\text{Mole Fraction Glucose} = \text{Moles of A} / \text{Total moles of solution} = 0.0315 \text{ m} / 1.43 \text{ m} = 0.0220 = \mathbf{2.20 \times 10^{-2}}$$

$$\text{Mole Fraction Water} = \text{Moles of B} / \text{Total moles of solution} = 1.40 \text{ m} / 1.43 \text{ m} = 0.979 \text{ m} = \mathbf{9.79 \times 10^{-1}}$$

NOTE: The sum of the mole fractions = 1.

Conversion of Concentration Units!

Example 12.5 Converting Molality to Mole Fractions: 0.120 m glucose in water, what are the mole fractions of each component.

$$\text{Molality} = \text{Moles Solute} / \text{kg solvent} \quad \text{Mole Fraction} = \text{Moles A} / \text{Total Moles}$$

$$\text{Need to convert kg } H_2O \text{ to Moles. } 1.00 \times 10^3 \text{ g} / 18.0 \text{ g/mole} = 55.6 \text{ moles } H_2O$$

$$\text{Mole Fraction of Glucose} = 0.120 \text{ mol glucose} / [0.120 \text{ mol} + 55.6 \text{ mol}] = 0.00215 = \mathbf{2.15 \times 10^{-3}}$$

$$\text{Mole Fraction of } H_2O = 55.6 \text{ mol } H_2O / [0.120 \text{ mol} + 55.6 \text{ mol}] = 0.998 = \mathbf{9.98 \times 10^{-1}}$$

Exercise 12.8 A solution is 0.120 m of methanol [CH_3OH] dissolved in ethanol [H_3CCH_2OH]. Calculate the mole fractions of each?

$$\text{Molality} = \text{Moles Solute} / \text{kg solvent} \quad \text{Mole Fraction} = \text{Moles A} / \text{Total Moles}$$

$$\text{Need to convert kg ethanol to Moles. } 1.00 \times 10^3 \text{ g} / 46.07 \text{ g/mole} = 21.7 \text{ moles ethanol}$$

$$\text{Total Moles} = 21.7 \text{ moles ethanol} + 0.120 \text{ m methanol} = 21.8 \text{ moles}$$

$$\text{Mole Fraction of Ethanol} = 21.7 \text{ moles ethanol} / 21.8 \text{ moles total} = \mathbf{0.995}$$

$$\text{Mole Fraction of Methanol} = 0.120 \text{ moles methanol} / 21.8 \text{ moles total} = 0.00550 = \mathbf{5.50 \times 10^{-3}}$$

Example 12.6 Converting Mole Fractions to Molality: 0.150 mole fraction glucose [$C_6H_{12}O_6$] and 0.850 mole fraction of water. What is the molality of the glucose in the solutions?

$$\text{Mole Fraction} = \text{Moles A} / \text{Total Moles} \quad \text{Molality} = \text{Moles Solute} / \text{kg solvent}$$

Need to convert Total Moles to kg of solvent

One mole of solution contains 0.150 m of glucose and 0.850 mole of H_2O .

Convert the 0.850 mole of H_2O to kg.

$$0.850 \text{ mole} * 18.0 \text{ g / mole} = 15.3 \text{ g } H_2O = 1.53 \times 10^{-2} \text{ kg } H_2O$$

$$\text{Molality} = \text{Moles Solute} / \text{kg solvent} = 0.150 \text{ mole glucose} / 1.53 \times 10^{-2} \text{ kg } H_2O = \mathbf{9.80 \text{ m glucose}}$$

Exercise 12.9 0.250 mole fraction of methanol and 0.750 mole fraction in ethanol. What is the molality of methanol?

$$\text{Mole Fraction} = \text{Moles A} / \text{Total Moles} \quad \text{Molality} = \text{Moles Solute} / \text{kg solvent}$$

Need to convert Total Moles to kg of solvent

One mole of solution contains 0.250 m of methanol and 0.750 mole of ethanol.

Convert 0.750 mole of ethanol to kg

$$0.750 \text{ mole ethanol} * 46.07 \text{ g / mole} = 34.55 = 34.6 \text{ g ethanol}$$

$$\text{Molality} = \text{Moles Solute} / \text{kg solvent} = 0.250 \text{ mole methanol} / 0.0346 \text{ kg ethanol} = 7.225 = \mathbf{7.23}$$

Example 12.7 Converting Molality to Molarity: A KCl solution is 0.273 m. What is the molar concentration of KCl? The density of the solution is $1.011 \times 10^3 \text{ g/L}$.

$$\text{Molality} = \text{Moles Solute} / \text{kg solvent} \quad \text{Molarity} = \text{Moles of Solute} / \text{Liters of Solution}$$

Need to convert kg of solvent to liters.

$$0.273 \text{ m KCl} = 0.273 \text{ moles KCl} / 1 \text{ kg H}_2\text{O}$$

$$\text{g KCl} = 0.273 \text{ moles KCl} * 74.6 \text{ g / mole} = \mathbf{20.4 \text{ g KCl}}$$

$$\text{Total mass of the solution is } 20.4 \text{ g KCl} + 1.00 \times 10^3 \text{ g H}_2\text{O} = 1.02 \times 10^3 \text{ g}$$

The volume is determined from the density:

$$\text{Vol} = 1.02 \times 10^3 \text{ g} / 1.011 \times 10^3 \text{ g/L} = 1.01 \text{ L}$$

$$\text{Molarity} = \text{Moles of Solute} / \text{Liters of Solution} = 0.273 \text{ moles KCl} / 1.01 \text{ L} = 0.270 \text{ M KCl} =$$

$$\mathbf{2.70 \times 10^{-1} \text{ M KCl}}$$

Note: Molarity vs Molality concentrations usually do not differ greatly.

Exercise 12.10 A solution is 3.42 m [**m = molality = moles / kg**] in urea [$\text{H}_2\text{N-CO-NH}_2$]. It's density is 1.045 g/mL. What is the molar [**M = moles / Liter**] concentration of urea?

$$3.42 \text{ m urea} * 60.05 \text{ g / mole} = 205.4 \text{ g urea}$$

$$\text{Total mass of solution is } 205.4 \text{ g urea} + 1.00 \times 10^3 \text{ g H}_2\text{O} = 1.21 \times 10^3 \text{ g}$$

The volume is determined from the density:

$$\text{Vol} = 1.21 \times 10^3 \text{ g} / 1.045 \text{ g/mL} = 1158 \text{ ml} = 1.16 \text{ L}$$

$$\text{Molarity} = \text{Moles of Solute} / \text{Liters of Solution} = 3.42 \text{ moles urea} / 1.16 \text{ L} = \mathbf{2.95 \text{ M urea}}$$

Example 12.8 Converting Molarity to Molality: A 0.907 M [moles / Liter] lead nitrate [$\text{Pb}(\text{NO}_3)_2$] solution has the density of 1.252 g/mL. What is its Molality [moles / Kg] ?

$$\text{Molarity} = \text{Moles of Solute} / \text{Liters of Solution} \quad \text{Molality} = \text{Moles Solute} / \text{kg solvent}$$

Convert Liters to kg

0.907 Moles of $\text{Pb}(\text{NO}_3)_2$ in 1 L of H_2O

$$\text{Mass of Solution} = 1 \text{ L} * 1000 \text{ mL} / \text{L} * 1.252 \text{ g} / \text{mL} = 1.252 \times 10^3 \text{ g}$$

$$\text{Mass of Lead Nitrate} = 0.907 \text{ moles} * 331.2 \text{ g} / \text{mole} = 3.00 \times 10^2 \text{ g Pb}(\text{NO}_3)_2$$

$$\text{The mass of water} = \text{mass of solution} - \text{mass of Pb}(\text{NO}_3)_2 = 1.252 \times 10^3 \text{ g} - 3.00 \times 10^2 \text{ g} =$$

$$1252 - 300 = 952 = 9.52 \times 10^2 \text{ g water}$$

$$\text{Molality} = \text{Moles Solute} / \text{kg solvent} = 0.907 \text{ moles Pb}(\text{NO}_3)_2 / 9.52 \times 10^2 \text{ g H}_2\text{O} = \mathbf{0.953 \text{ m Pb}(\text{NO}_3)_2}$$

Exercise 12.11 An 2.00 M urea aqueous solution has a density of 1.029 g/ml. What is the molality of urea?

Vapor Pressure

Vapor Pressure lowering of a solvent is a colligative property equal to the vapor pressure of the pure solvent minus the vapor pressure of the solution.

Raoult's Law: The partial pressure of solvent P_A , over a solution equals the vapor pressure of the pure solvent P_A^0 times the **mole fraction** of solvent X_A in the solution:

$$P_A = P_A^0 X_A$$

The vapor pressure of the solvent, with a nonvolatile solute, is always less than the pure solvent.

$$\text{For a change in vapor pressure:} \quad \Delta P_A = P_A^0 X_A$$

The vapor pressure depends on the concentration, but not the identity of the solute.

e.g. Antifreeze is ethylene glycol. If the mole fraction of ethylene glycol is doubled from 0.010 to 0.020, the vapor pressure lowering is doubled from 0.18 to 0.36 mm Hg, and thusly the BP is increased!

Example 12.9 Calculate the vapor pressure lowering of water when 5.67 g of glucose [$\text{C}_6\text{H}_{12}\text{O}_6$] is dissolved in 25.2 g of water at 25 °C. The vapor pressure of water at 25 °C is 23.8 mm Hg.

This solution is 0.0220 mole fraction in glucose [see above]

$$\Delta P_A = P_A^0 X_A = 23.8 \text{ mm Hg} * 0.0220 = \mathbf{0.524 \text{ mm Hg lower vapor pressure}}$$

$$\text{The vapor pressure of the solution is } 23.8 \text{ mm Hg} - 0.524 \text{ mm Hg} = \mathbf{23.3 \text{ mm Hg}}$$

The boiling point of the water is raised!

Exercise 12.12 0.525 g of naphthalene [C_{10}H_8] is dissolved in 60.8 g of chloroform [CHCl_3]. Calculate the vapor pressure lowering at 20 °C. The vapor pressure of chloroform at 20 °C is 156 mm Hg. What is the vapor pressure of the solution?

$$\Delta P = \mathbf{1.22 \text{ mm Hg}} \quad P = \mathbf{155 \text{ mm Hg}} \quad [\text{ Prove it }]$$

For ideal solutions, solutions that are chemically similar, the substances follow Raoult's law:

$$P = P_A^0 X_A + P_b^0 X_b$$

The total vapor pressure equals the sum of the partial pressures [pressure * mole fraction]

Example: a solution is 0.70 mf in benzene and 0.30 mf in toluene. The vapor pressure of benzene is 75 mm Hg and of toluene is 22 mm Hg.

The total vapor pressure is: $(75 \text{ mm Hg} * 0.70) + (22 \text{ mm Hg} * 0.30) = 59 \text{ mm Hg}$

The vapor pressure of benzene is: $(75 \text{ mm Hg} * 0.70) = 52 \text{ mm Hg}$

The mole fraction of benzene in the vapor is $52 \text{ mm Hg} / 59 \text{ mm Hg} = \mathbf{0.9}$

So in the liquid form, benzene is 0.70 mf, in the vapor it's 0.9 or much richer.

If you distill this mixture, the product is the richer in benzene distilling first.

Fractional Distillation: a fractionating column is used and the vapors condense on the cooler packing material, this liquid which is richer in benzene and then vaporizes to condense further up – it becomes richer in benzene.

12.6 Boiling Point Elevation & Freezing Point Depression

Boiling Point of a liquid is the temperature at which it's vapor pressure equals 1 atm.

The addition of a nonvolatile solute to a liquid lowers the liquids vapor pressure. The BP is then increased as the temp must be increased to a greater value then normal to achieve a vapor pressure of 1 atm.

BP elevation is a colligative property of a solution and is equal to the BP of the solution minus the BP of the pure solvent.

The BP elevation is proportional to the **molal concentration** [c_m] of the solution: $\Delta T_b = K_b c_m$
 K_b is the boiling point elevation constant and is solvent dependant.

K_b for benzene is 2.61 °C/m, so 0.100 m solution of a nonvolatile in benzene will raise its BP 0.261 °C, or from 80.2 °C to 80.5 °C.

FP depression is a colligative property of a solution equal to the FP of the pure solvent minus the FP of he solution:

The FP depression is proportional to the **molal concentration** [c_m] of the solution: $\Delta T_f = K_f c_m$
 K_f is the freezing point depression constant.

The K_f for benzene is 5.07 °C/m, so a 0.100 m solution freezes at 0.507 °C below benzene's normal FP of 5.46 °C or 4.95 °C.

Example 12.10 You have an aqueous solution of 0.0222 m glucose. What are the BP elevation and FP lowering?

K_b for water us 0.512 °C/m and the K_f for water is 1.86 °C/m.

$$\Delta T_b = K_b c_m = 0.512 \text{ °C/m} * 0.0222 \text{ m} = 0.0114 \text{ °C} \quad \text{BP} = 100.000 \text{ °C} + 0.0114 \text{ °C} = 100.011 \text{ °C}$$

$$\Delta T_f = K_f c_m = 1.86 \text{ °C/m} * 0.222 \text{ m} = 0.0413 \text{ °C} \quad \text{FP} = 0.000 \text{ °C} - 0.0413 \text{ °C} = -0.041 \text{ °C}$$

FP depression is used in cars where ethylene glycol is added to the radiator to lower the freezing point and it raises it's BP.

Sodium Chloride is added to icy roads to lower the MP of the ice below that of the air, so the ice will liquefy.

Freezing Point Depression is used to determine molecular weights.

You can determine [**BSOTC**] the freezing point depression based on the Molality of the solution. From this, you can determine the molecular weight of the compound.

Example 12.11 0.131g of an unknown compound is dissolved in 25.4 g of water. The molality of the solution is determined by FP depression to be 0.056 m. What is the mw of the compound?

$$\text{Molality} = \frac{\text{Moles Solute}}{\text{kg solvent}} = 0.056 \text{ m} = \frac{\text{moles of compound}}{2.54 \times 10^{-3} \text{ kg H}_2\text{O}}$$

$$\text{moles of compound} = 0.056 \text{ m} * 2.54 \times 10^{-3} \text{ kg H}_2\text{O} = 1.42 \times 10^{-3} \text{ moles} = \text{g} / \text{mw} = 0.131 \text{ g} / \text{mole}$$

$$\text{mw} = 0.131 \text{ g} / 1.42 \times 10^{-3} \text{ moles} = \mathbf{92.3 \text{ g} / \text{mole}}$$

Exercise 12.14 0.930 g of ascorbic acid [Vit C] dissolved in 95.0 g water. The concentration was determined by FP depression to be 0.0555 m. What is the mw of ascorbic acid [mw = 176 g/mole] ?

Example 12.12 Camphor melts at 179.5 °C. The freezing point depression [K_f] is 40 °C/m. A 1.07 mg sample of a compound was dissolved in 78.1 mg of camphor and melted at 176.0 °C. What is the mw of the compound? The empirical formula for camphor is CH, what is the molecular formulae?

$$\Delta T = [179.5 \text{ }^\circ\text{C} - 176.0 \text{ }^\circ\text{C}] = 3.5 \text{ }^\circ\text{C}$$

$$\Delta T_f = K_f c_m$$

$$c_m = \text{molal concentration} = \Delta T_f / K_f = 3.5 \text{ }^\circ\text{C} / 40 \text{ }^\circ\text{C/m.} = 0.088 \text{ m} = \text{Moles} / \text{Kg of Solvent}$$

$$\text{Moles} = c_m * \text{Kg of Solvent} = 0.088 \text{ m} * 78.1 \times 10^{-6} \text{ kg} = 6.87 = 6.9 \times 10^{-6} \text{ moles}$$

$$\text{Moles} = \text{wt} / \text{mw} = 1.07 \times 10^{-3} \text{ g} / 6.9 \times 10^{-6} \text{ moles} = 1.55 = 1.6 \times 10^{+2} \text{ g/mole} = @ 160$$

$$n = \text{mw} / \text{emp form} = 160 / 13 = @12$$

$$\text{Mole Form} = \text{C}_{12}\text{H}_{12}$$

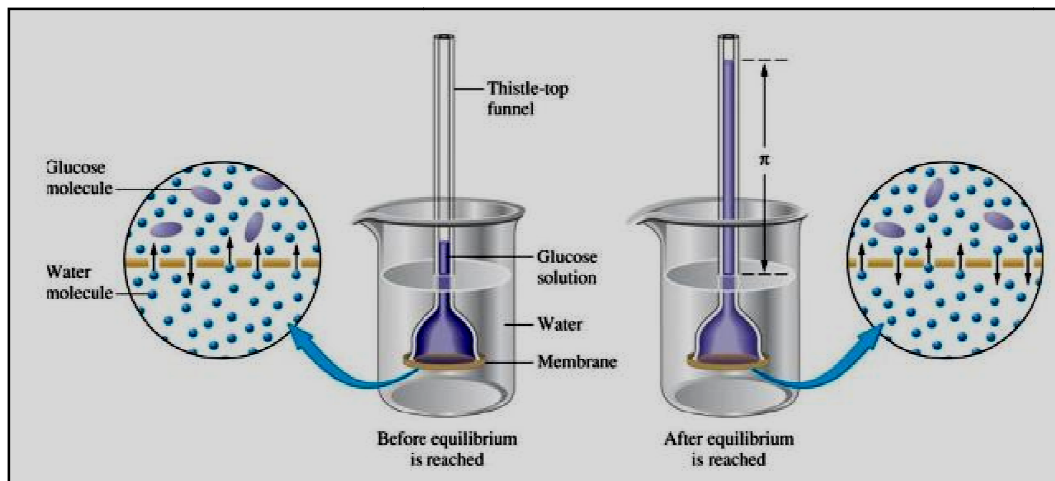
Osmosis: is the phenomenon of solvent flow through a semipermeable membrane to equalize the solute concentration on both sides of the membrane.

Semipermeable membrane – allows solvents to pass through, but not solutes

Osmotic Pressure [Π - pi] is a colligative property of a solution equal to the pressure that when applied to the solution, just stops osmosis:

$$\mathbf{\Pi = M R T}$$

M = osmotic pressure, R is the gas constant, T is the absolute temp.



Example: A solution with a molarity of 0.02 m has a FP depression of 0.04 °C. At dilute solutions, molarity and molality are approximately the same. Hence the Osmotic Pressure at 25 °C [298 °K] is:

$$\Pi = M R T = 0.02 \text{ mol / L} * 0.082 \text{ L atm / (}^\circ\text{K mol)} * 298 \text{ }^\circ\text{K} = \mathbf{0.5 \text{ atm.}}$$

0.5 atm = a water column 4.5 meters high

Example 12.13 A starch with a formula of $(\text{C}_6\text{H}_{10}\text{O}_5)_n$ where n is around 200. 0.798 g is dissolved in 100.0 ml of water. What is the osmotic pressure at 25°C?

Mw of $(\text{C}_6\text{H}_{10}\text{O}_5)_{200}$ is around 32,400 g / mole

$$\# \text{ moles} = 0.798 \text{ g} / 32,400 \text{ g/mole} = 2.46 \times 10^{-5} \text{ mole}$$

$$\mathbf{M = Molarity} = \frac{\text{Moles of Solute}}{\text{Liters of Solution}} = \frac{2.46 \times 10^{-5} \text{ mole}}{0.10000 \text{ L}} = 2.46 \times 10^{-4} \text{ mol / L}$$

$$\Pi = M R T = 2.46 \times 10^{-4} \text{ mol / L} * 0.0821 \text{ L atm / (}^\circ\text{K mol)} * 298 \text{ }^\circ\text{K} = x$$

$$x * 760 \text{ mm Hg / 1 atm}$$

$$\text{Osmotic Pressure} = \Pi = \mathbf{4.58 \text{ mm Hg.}}$$

As a comparison, the FP depression is $4.58 \times 10^{-4} \text{ }^\circ\text{C}$ – very difficult to measure this small difference in the lab.

Therefore it's easier to measure the Osmotic Pressure and then determine the molecular weight of a compound than to determine it by the freezing point depression.

Osmotic pressure is very important in biological processes.

Colloids: A colloid is a dispersion of particles of one substance throughout another substance or solution. Fog is a colloid.

Tyndall effect is the scattering of light by colloidal size particles.

Types and Examples of Colloids are

<u>Continuous Phase</u>	<u>Dispersed Phase</u>		
	SOLID	LIQUID	GAS
SOLID	Ruby Glass	Jelly	Plastic Foam
LIQUID	AgCl in HOH	Mayonnaise	Whipped Cream
GAS	Smoke	Fog	N/A

Hydrophilic colloid is a colloid in which there is a **STRONG ATTRACTION** between the dispersed phase and the continuous phase [water].

Hydrophobic colloid is a colloid in which there is a **LACK OF ATTRACTION** between the dispersed phase and the continuous phase [water]. These are unstable and will change into larger particles over time.

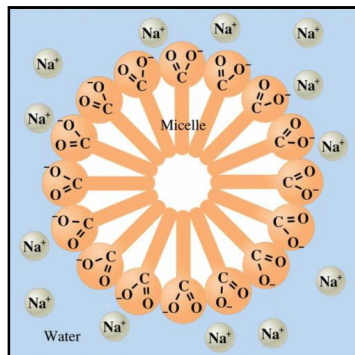
Example Iron III Hydroxide

Coagulation is the process by which the dispersed phase of a colloid is made to aggregate and thereby separate from the continuous phase.

Example: curdling of milk.

Micelle is a colloidal size particle formed in water by the association of molecules or ions that each have a hydrophobic end and a hydrophilic end.

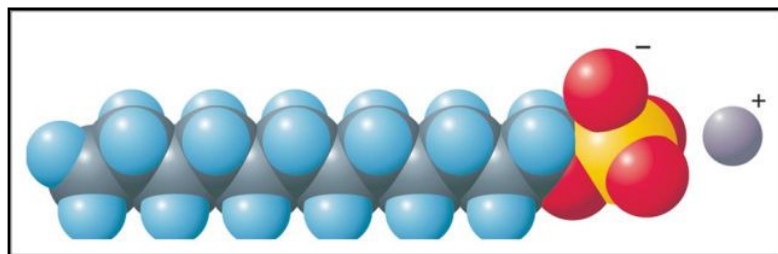
Example soap



Association Colloid is a colloid in which the dispersed phase consists of micelles.

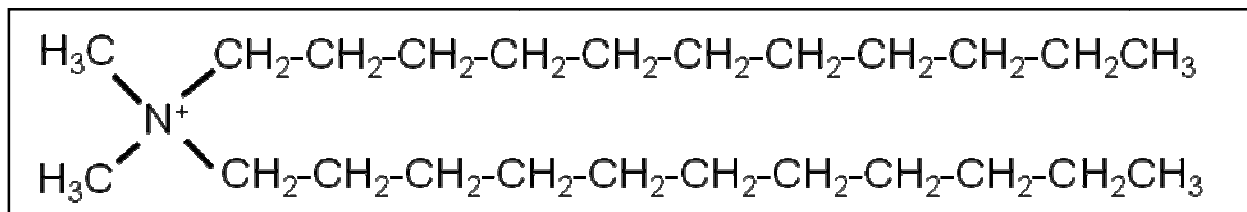
Some detergent molecules have a **negative end** at the hydrophilic end, these are **cationics**

Example Acid Salts



Some have a **positive end** – these are called **Anionics**

Example Ammonium Salts



Alternate Example 12.2 Calculating Mass Percentage of Solute

An experiment calls for 36.0 g of a 5.00% aqueous solution of potassium bromide. Describe how you would make up such a solution.

Answer: Dissolve 1.8 g KBr in 34.2 g H₂O.

Alternate Example 12.3 Calculating the Molality of Solute

Iodine dissolves in various organic solvents, such as methylene chloride, in which it forms an orange solution. What is the molality of I₂ in a solution of 5.00 g of iodine, I₂, in 30.0 g of methylene chloride, CH₂Cl₂?

Answer: 0.657 m

First Alternate Example 12.4 Calculating the Mole Fractions of Components

A solution of iodine in methylene chloride, CH₂Cl₂, contains 1.50 g I₂ and 56.00 g CH₂Cl₂. What are the mole fractions of each component in the solution?

Answer: 8.89 · 10⁻³ mole fraction I₂; 0.9911 mole fraction CH₂Cl₂

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Alternate Examples for Lecture 91

Second Alternate Example 12.4 Calculating the Mole Fractions of Components

A bottle of bourbon is labeled 94 proof, or 47% by volume alcohol in water. What is the mole fraction of ethyl alcohol, C₂H₅OH, in the bourbon? The density of ethyl alcohol is 0.80 g/mL.

Answer: 0.22 mole fraction C₂H₅OH

Practice Questions:

Review Questions:

Concept Questions:

Practice Problems: 12.33, 12.34, 12.41
12.43, 12.44, 12.45, 12.46
12.47, 12.48, 12.49, 12.50
12.51, 12.52
12.56, 12.61
12.77, 12.78
12.93