

Solutions

Chapter 12

Why does a raw egg swell or shrink when placed in different solutions?

Solutions

Chapter 12

**Dissolving the
shell in vinegar**

**Egg in
Pure
water**

**Egg in
Corn
syrup**

**Why does a raw egg swell or shrink
when placed in different solutions?**

Chapter 12

Solutions

1. Important Definitions
2. Expressing Concentrations of Solutions
3. The Thermodynamics of the Dissolving Process
4. Dynamic Equilibrium
5. The Solubility of Solids, Liquids, and Gases
6. Colligative Properties
7. Colloids (*NOT Solutions!*)

Important Definitions

Solution - A homogeneous mixture consisting of one or more substances uniformly dispersed as separate atoms, molecules, or ions in another substance.

Solvent - The component of a solution that is the “dissolving” medium. The solvent determines the physical state of the solution (*solid, liquid, or gas*).

Solute - The components of a solution that are “dissolved” by the medium.

Aqueous Solution - A solution wherein water is the solvent.

When the solute and solvent are both in the same physical state, the one in the largest quantity is the solvent.

Solubility

- When one substance (solute) dissolves in another (solvent) it is said to be **soluble**
 - ◆ salt is soluble in water
 - ◆ bromine is soluble in methylene chloride
- When one substance does not dissolve in another it is said to be **insoluble**
 - ◆ oil is insoluble in water
- The solubility of one substance in another depends on two factors – nature's tendency towards mixing, and the types of intermolecular attractive forces

Spontaneous Mixing

Solubility

- There is usually a limit to the solubility of one substance in another
 - ◆ gases are *always* soluble in each other
 - ◆ two liquids that are mutually soluble are said to be **miscible**
 - ☞ alcohol and water are miscible
 - ☞ oil and water are immiscible
- The maximum amount of solute that can be dissolved in a given amount of solvent is called the **solubility**
- The solubility of one substance in another varies with temperature and pressure

Expressing Concentration

Concentration - The amount of **solute** present in a given quantity of **solvent** or **solution**.

1. $\frac{\text{mass solute}}{\text{volume solvent}}$ $\frac{35.7 \text{ g NaCl}}{100 \text{ mL H}_2\text{O}}$ $\frac{203 \text{ g Sugar}}{100 \text{ mL H}_2\text{O}}$

(Typical way to express solubilities)

2. Percent by Mass

$$\left[\frac{\text{mass solute}}{\text{mass solution}} \right] (100 \%)$$

What is the percent by mass of sugar in a solution containing 50.0 g of sugar and 65.0 g H₂O?

(mass of solute/mass of solution)(100%) =

$$\left[\frac{50.0 \text{ g sugar}}{50.0 \text{ g sugar} + 65.0 \text{ g H}_2} \right] (100.0\%) =$$

43.5 % by mass

3. Molarity (M) (*Temperature Dependent*)

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liter solution}}$$

Practice Problem: A 650-mL solution contains 1.22 mol of KNO_3 . Calculate the molarity of the solution.

$$\left[\frac{1.22 \text{ mol KNO}_3}{650 \text{ mL soln}} \right] \left[\frac{1000 \text{ mL soln}}{\text{L soln}} \right] = 1.88 \text{ M KNO}_3$$

Practice Problem: Calculate the molarity of a solution that contains 5.25 g AgNO₃ in 175 mL of solution.

A. Calculate the formula mass of AgNO₃.

$$\begin{array}{r} 1 \text{ Ag@ } 107.868 \text{ u} = 107.868 \text{ u} \\ 1 \text{ N@ } 14.0067 \text{ u} = 14.0067 \text{ u} \\ 3 \text{ O@ } 15.9994 \text{ u} = 47.9982 \text{ u} \\ \hline 1 \text{ AgNO}_3 = 169.873 \text{ u} \end{array}$$

B. Determine the moles of AgNO₃ in 5.25 g AgNO₃.

$$(5.25 \text{ g AgNO}_3)(1 \text{ mol AgNO}_3/169.873 \text{ g AgNO}_3) =$$

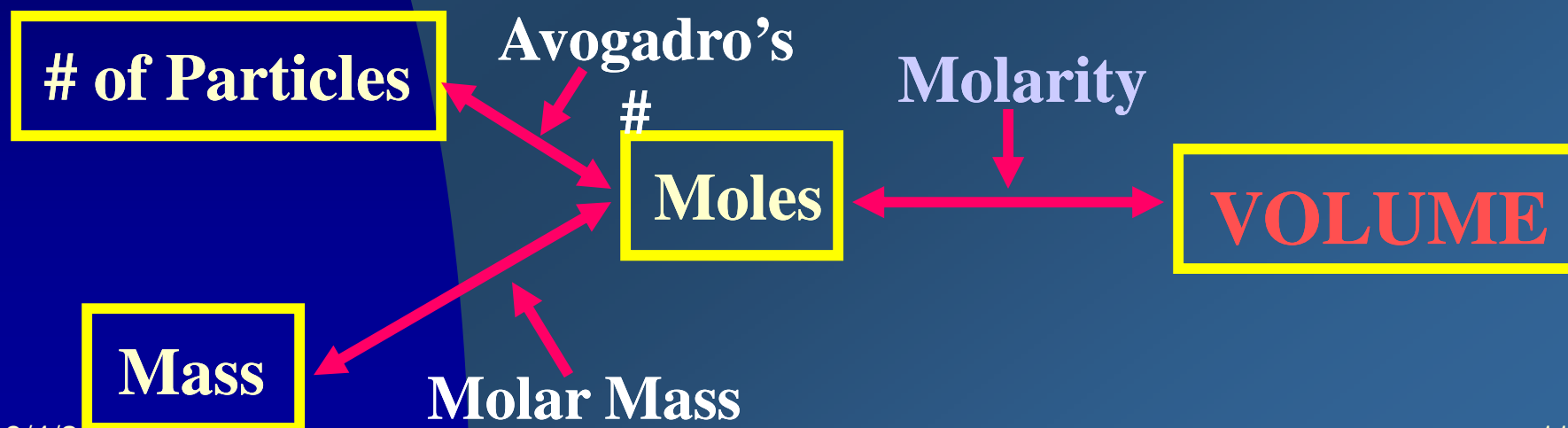
$$0.03091 \text{ mol AgNO}_3$$

C. From the moles of AgNO_3 and the mL of solution, calculate the molarity of the solution.

$$\left[\frac{0.03091 \text{ mol AgNO}_3}{175 \text{ mL soln}} \right] \left[\frac{1000 \text{ mL soln}}{1 \text{ L soln}} \right]$$

$$= 0.177 \text{ M AgNO}_3$$

(“When in doubt, calculate moles”)



How to make a 1.00 M solution of NaCl (aq)

4. Molality (m) (*Temperature Independent*)

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

Practice Problem: What is the molality of a saturated solution of NaCl in water at 0°C? NaCl's solubility is 35.7 g/100 mL at 0°C. Water's density is 1.000 g/mL at 0°C.

35.7 g NaCl	1 mL H₂O	10³ g H₂O	1 mol NaCl
100 mL H₂O	1.000 g H₂O	kg H ₂ O	58.44 g NaCl

$$\frac{6.11 \text{ mol NaCl}}{\text{kg H}_2\text{O}} = \boxed{6.11 \text{ m}}$$

5. Mole Fraction (X) (*Temperature Independent*)

$$\frac{\text{moles solute}}{\text{moles solute} + \text{moles solvent}} = \frac{\text{moles solute}}{\text{total moles}}$$

Practice Problem: If 26.2 g of H₂O (g) is mixed with 43.7 g of O₂ (g), what is the mole fraction of water vapor in this gaseous solution?

$$\frac{\cancel{26.2 \text{ g H}_2\text{O}}}{\cancel{18.015 \text{ g H}_2\text{O}}} \left| \frac{1 \text{ mol H}_2\text{O}}{\cancel{18.015 \text{ g H}_2\text{O}}} \right. = 1.45\bar{4} \text{ mol H}_2\text{O}$$

$$\frac{\cancel{43.7 \text{ g O}_2}}{\cancel{31.9988 \text{ g O}_2}} \left| \frac{1 \text{ mol O}_2}{\cancel{31.9988 \text{ g O}_2}} \right. = 1.36\bar{6} \text{ mol O}_2$$

$$X_{\text{water}} = \frac{1.454 \text{ mol H}_2\text{O}}{1.454 \text{ mol H}_2\text{O} + 1.366 \text{ mol O}_2}$$

$$X_{\text{water}} = \frac{1.454 \text{ mol H}_2\text{O}}{2.818 \text{ mol H}_2\text{O}} = 0.516 \text{ mol fraction}$$

6. Fractions by mass

$(a/b)(100)$ = percent = “parts per hundred” (pph)

$(a/b)(10^3)$ = parts per thousand

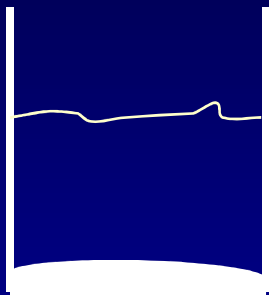
$(a/b)(10^6)$ = parts per million (ppm)

$(a/b)(10^9)$ = parts per billion (ppb)

$(a/b)(10^{12})$ = parts per trillion (ppt)

Solubility - The amount of **solute** that will dissolve in a given quantity of **solvent** at a given temperature.

Saturated Solution - A solution that contains an amount of **solute** that is equal to its solubility.

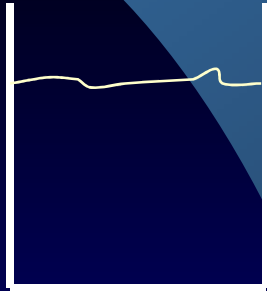


*The dissolved solute in the solution is in **dynamic equilibrium** with the undissolved solute (the precipitate).*

Dissolved solute \rightleftharpoons Precipitate

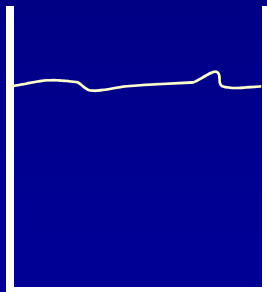
*The rate of dissolving is equal to the rate of precipitation.
This is a **dynamic equilibrium**.*

Unsaturated Solution - A solution that contains an amount of *solute* that is less than its solubility.



*All of the **solute** present is dissolved in an unsaturated solution.*

Supersaturated Solution - A solution that contains an amount of *solute* that is more than its solubility. This is a “metastable” state (“without stability”).



*If disturbed in anyway, the excess solute will precipitate out of solution and a **saturated solution** will result.*

Supersaturated Solutions

9/4/2008

18

The Thermodynamics of Dissolving

Why is n-hexane soluble in CCl_4 but water is not??

Stay Tuned!!

The Thermodynamics of Dissolving

The Dissolving Process: In order to predict whether or not a material will dissolve in another material, one must know the *enthalpy* change and the *entropy* change that will occur upon dissolving and the temperature.

Entropy - the *disorder* of a system. Most dissolving processes involve an *increase* in entropy. That is, the disorder present in the separate solute and solvent is lower than it is in the solution. This *favours* dissolving.

Enthalpy - the *total energy* of a system. Some dissolving processes are *endothermic*. That is, the solution has a higher enthalpy than the separate solute and solvent - $\Delta H > 0$! These processes do *NOT* favor dissolving.

The Thermodynamics of Dissolving

Enthalpy - Some dissolving processes are *exothermic*. That is, the solution has a lower enthalpy than the separate solute and solvent - $\Delta H < 0$! These processes do *favor* dissolving.

Summary: To predict whether or not a solute will dissolve *spontaneously* in a given solvent, one must evaluate both the *enthalpy* change that will occur and the *entropy* change that will occur during dissolving. The temperature must also be taken into consideration.

Since the entropy almost always increases during dissolving, we can concentrate, for now, on the *enthalpy* change.

Spontaneous Mixing

The enthalpy of solution is determined by the nature of the intermolecular attractive forces!

Water Dissolving Sodium Chloride Crystals

9/4/2008

24



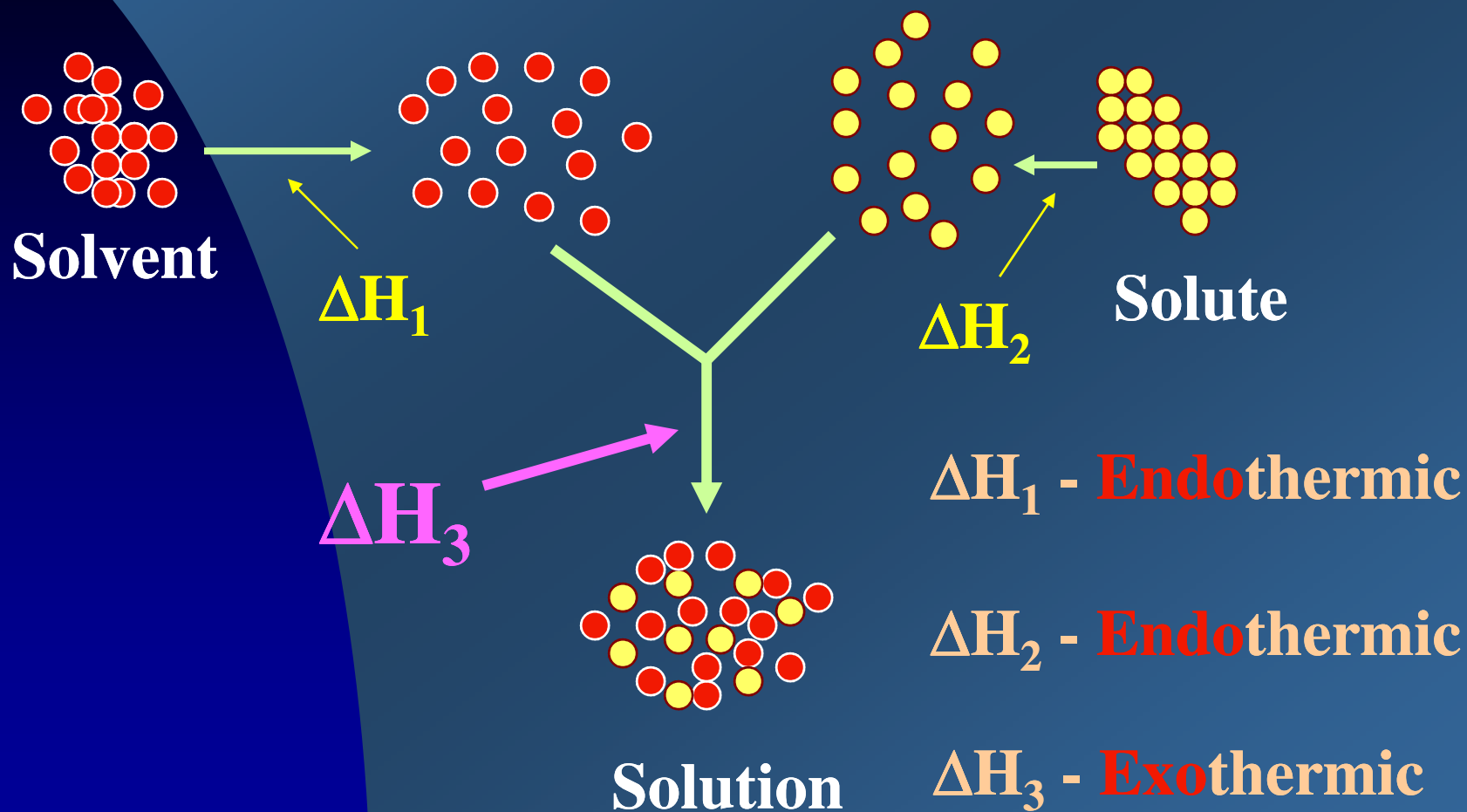
Gas

Liquid

Solution

The Thermodynamics of Dissolving

The Dissolving Process:



*Calculating
Enthalpy of
Solution*

9/4/2008

27

Calculating Enthalpy of Solution



Heat of Hydration

Ion-Dipole Interactions

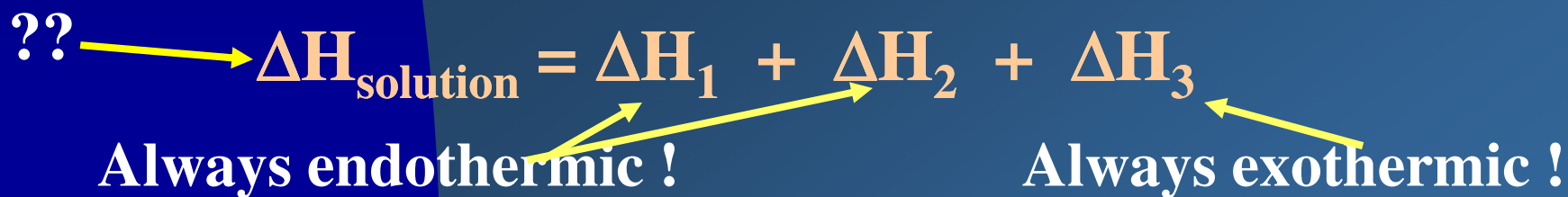
- When ions dissolve in water they become **hydrated**
- Each ion is surrounded by water molecules

The Thermodynamics of Dissolving

9/4/2008

31

The Thermodynamics of Dissolving



The Thermodynamics of Dissolving

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

If $\Delta H_3 > \Delta H_1 + \Delta H_2$ then $\Delta H_{\text{solution}} < 0$

If $\Delta H_3 < \Delta H_1 + \Delta H_2$ then $\Delta H_{\text{solution}} > 0$

If $\Delta H_3 = \Delta H_1 + \Delta H_2$ then $\Delta H_{\text{solution}} = 0$

If the solvent is a liquid, then ΔH_1 is the heat of vaporization for the liquid. If the solute is a solid, then ΔH_2 is the heat of sublimation of the solid.

How could you determine the value of ΔH_3 ?

The Thermodynamics of Dissolving

The Role of Intermolecular Attractive Forces:

For a dissolving process to be **exothermic**, the **solvent-solute** attractive forces must be greater than the **solute-solute** and the **solvent-solvent** attractive forces combined.

Can a substance dissolve if the process is **endothermic**?

Substance	$\Delta H_{\text{solution}}$
CO ₂ (g)	-19.41 kJ/mol
LiCl (s)	- 37.03 kJ/mol
KCl (s)	+ 17.22 kJ/mol
NaCl (s)	+ 3.88 kJ/mol
H ₂ SO ₄ (l)	- 95.28 kJ/mol

The Thermodynamics of Dissolving

1. Ethyl alcohol (*ethanol*) is miscible with both water and *n*-hexane. However, water is not miscible with *n*-hexane. *Why??*

$\text{H}_2\text{O} - \text{H}_2\text{O}$ interaction - *H-bonding and London Forces*

$\text{CH}_3\text{CH}_2\text{OH} - \text{CH}_3\text{CH}_2\text{OH}$ - *H-bonding and London Forces*

$\text{H}_2\text{O} - \text{CH}_3\text{CH}_2\text{OH}$ - *H-bonding and London Forces*

The interactions are all of similar strength. Thus, entropy is the predominating factor in solubility.

The Thermodynamics of Dissolving



H-bonding in Ethanol and Water

The Thermodynamics of Dissolving

2. AgCl is Not very soluble in water. *Why??*

H₂O - H₂O interaction - *H-bonding and London Forces*

Ag⁺ - Cl⁻ interaction - *Very strong ion-ion attraction
(Crystal Lattice Energy)*

**H₂O - Ag⁺
and H₂O - Cl⁻ interactions** - *Relatively weak ion-dipole
attractions*

The water-ion attractions are not strong enough to overcome the water-water and ion-ion attractions. Thus, enthalpy is the predominating factor in solubility.

The Thermodynamics of Dissolving

3. *n*-Heptane and *n*-hexane are completely miscible in each other. *Why??*

heptane - heptane interactions - *London Forces*

hexane - hexane interactions - *London Forces*

heptane - hexane interactions - *London Forces*

The attractive forces are very similar. Thus entropy is the predominant factor.

The Thermodynamics of Dissolving

4. *n*-Hexane and water are immiscible (not miscible).

Why??

H₂O - H₂O interaction - *H-bonding and London Forces*

hexane - hexane interactions - *London Forces*

H₂O - hexane interactions - *London Forces*

The water-hexane interactions are much weaker than the water-water interactions. Thus enthalpy (which is very positive - endothermic) is the predominant factor.

The Thermodynamics of Dissolving

The previous four examples demonstrate the useful principle that *“like dissolves like”*. In other words, materials that have similar structures and polarity such as *n*-heptane and *n*-hexane or water and ethanol will tend to be soluble or miscible in one another.



n-hexane



n-heptane



water



ethanol

The Thermodynamics of Dissolving

Some Useful Thermodynamic Concepts - *Read Chapter 19*

1. The enthalpy change accompanying the dissolving process can be thought of as the *change in total energy*. The total energy consists of the *energy available for doing work* (Free Energy, ΔG) and the *energy NOT available for doing work* (entropy, $T\Delta S$).

$$\Delta H_{\text{solution}} = \Delta G_{\text{solution}} + T\Delta S_{\text{solution}}$$

For a change to take place spontaneously, ΔG of the process must be less than 0.

The Thermodynamics of Dissolving

$$\Delta H_{\text{solution}} = \Delta G_{\text{solution}} + T\Delta S_{\text{solution}}$$

$$\Delta G_{\text{solution}} = \Delta H_{\text{solution}} - T\Delta S_{\text{solution}}$$

If $\Delta G_{\text{solution}} > 0$ at temperature T, the solute will **NOT** spontaneously dissolve.

If $\Delta G_{\text{solution}} < 0$ at temperature T, the solute will spontaneously dissolve.

Case 1: (-) (-) (+)

$$\Delta G_{\text{solution}} = \Delta H_{\text{solution}} - T\Delta S_{\text{solution}}$$

Spontaneous at all temperatures exothermic more disorder

The Thermodynamics of Dissolving

Case 2: (+) (+) (-)
 $\Delta G_{\text{solution}} = \Delta H_{\text{solution}} - T\Delta S_{\text{solution}}$
Non-spontaneous **endothermic** **less disorder**
at all temperatures

Case 3: (-) (-)
 $\Delta G_{\text{solution}} = \Delta H_{\text{solution}} - T\Delta S_{\text{solution}}$
exothermic **less disorder**

At higher temperatures, the entropy term dominates and the dissolving process is **non-spontaneous**. At lower temperatures, the enthalpy term dominates and the dissolving is **spontaneous**. Thus, this type of system would have higher solubility *at lower temperatures*.

The Thermodynamics of Dissolving

Case 4:

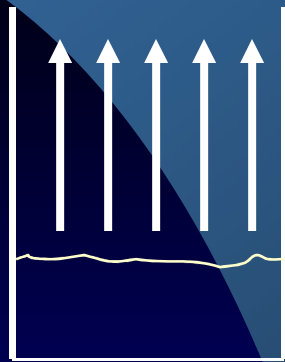
$$\Delta G_{\text{solution}} = \overset{(+)}{\Delta H_{\text{solution}}} - T \overset{(+)}{\Delta S_{\text{solution}}}$$

endothermic **more disorder**

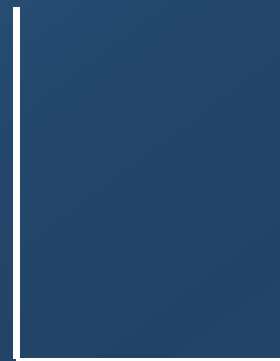
At higher temperatures, the entropy term dominates and the dissolving process is **spontaneous**.
At lower temperatures, the enthalpy term dominates and the dissolving is **non-spontaneous**.

Dynamic Equilibrium: When two opposing processes are occurring at exactly the same rate, a state of dynamic equilibrium has been achieved.

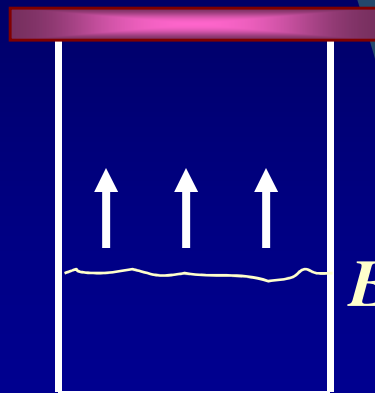
Dynamic Equilibrium



Evaporation

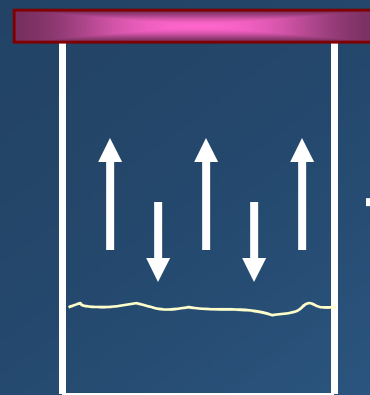


Open System
(No Equilibrium)

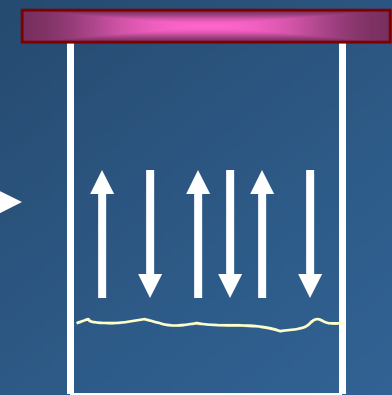


Evaporation

Liquid \rightarrow Gas
(No Equilibrium)



Liquid \rightleftharpoons Gas
(No Equilibrium)

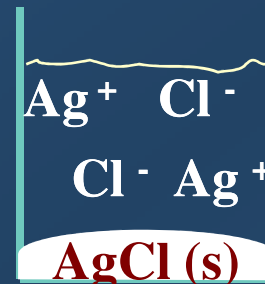


Liquid \rightleftharpoons Gas
(Equilibrium)

Dynamic Equilibrium



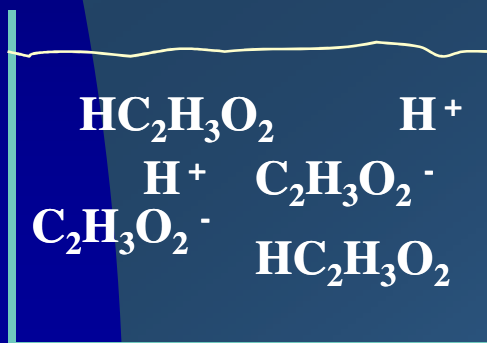
Chemical
Equilibrium



Rate of Precipitation = Rate of Dissolving



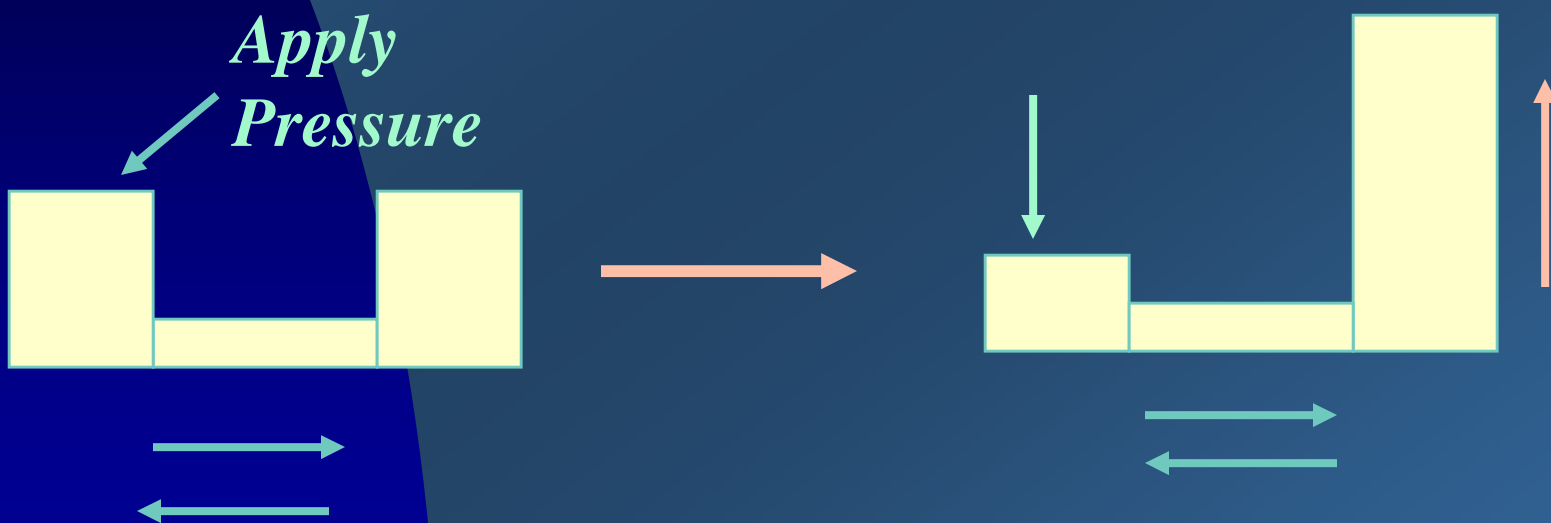
Rate of dissociation (ionization) = Rate of Association



Dynamic Equilibrium

The net result of a dynamic equilibrium is that no change in the system is evident.

Le Chatelier's Principle - *If a change is made in a system at equilibrium, the equilibrium will shift in such a way so as to reduce the effect of the change.*



Pressure applied to the system at equilibrium caused it to shift until a new equilibrium was established.

Solubility of Solids, Liquids, and Gases

Using Le Chatelier's Principle to predict the temperature effect on solubility:

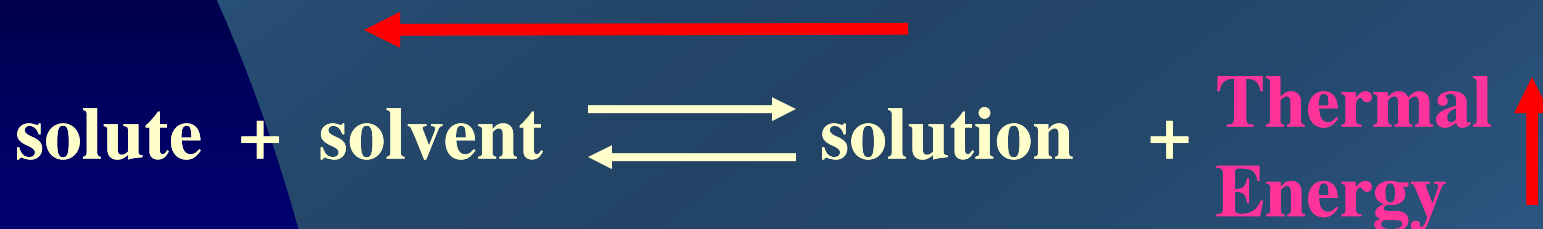
1. If a solute has an **endo**thermic enthalpy of solution ($\Delta H_{\text{solution}} > 0$), it will **TEND** to have a higher solubility at higher temperatures.

↑ Thermal Energy



Solubility of Solids, Liquids, and Gases

2. If a solute has an **exothermic** enthalpy of solution ($\Delta H_{\text{solution}} < 0$), it will **TEND** to have a lower solubility at higher temperatures.



Note: *The enthalpy of solution is somewhat dependent on the concentration of the solute. Therefore, the effect of temperature on the $\Delta H_{\text{solution}}$ is measured at the saturation point.*

Solubility of Solids, Liquids, and Gases

The Effects of Temperature on Solubility - *Most but not all liquids and solids increase in solubility as the temperature is increased.*

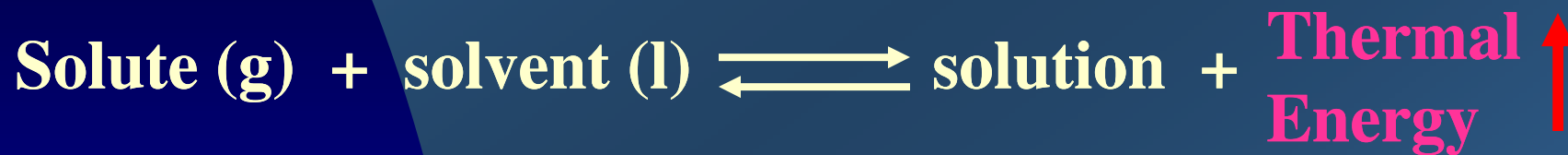
All gases, on the other hand, decrease in solubility with increasing temperature. Why???

Gas	0°C	25°C	50°C
CO ₂	0.33*	0.145*	0.076*
O ₂	0.0069*	0.0039*	0.001 22*
SO ₂	23*	9.4*	4.3*

* $\frac{\text{g gas}}{100 \text{ g H}_2\text{O}}$

Solubility of Solids, Liquids, and Gases

*Le Chatelier's Principle is useful for predicting the effects of temperature on the solubility of gases in liquids because these are all (or nearly all) **exothermic** changes.*



*Le Chatelier's Principle is not useful for predicting the effects of temperature on the solubility of solids and liquids. **WHY??***

Solubility of Solids and Liquids

*All solid and liquid solutes that dissolve **exothermically** should be less soluble at higher temperatures, according to Le Chatelier's Principle.*

BUT...

*If the solution process is **exothermic AND** produces a higher state of disorder (higher entropy), an increase in temperature will **favor** dissolving!*

$$\begin{array}{ccccccc} (-) & & (-) & & (+) & & \\ \Delta G_{\text{solution}} & = & \Delta H_{\text{solution}} & - & T\Delta S_{\text{solution}} & & \end{array}$$

Spontaneous

*Higher temperatures make the solution process **MORE** spontaneous!*

Solubility of Gases

*Then if the dissolving of gases is exothermic, why do their solubilities **ALWAYS DECREASE** with increasing temperature???*

Because when gases dissolve, the entropy actually decreases!

$$\overset{(+)}{\Delta G}_{\text{solution}} = \overset{(-)}{\Delta H}_{\text{solution}} - \overset{(-)}{T\Delta S}_{\text{solution}}$$

Thus, at higher temperatures, the entropy term dominates and $\Delta G_{\text{solution}}$ becomes more positive! This is why Le Chatelier's Principle works for the solubility of gases in liquids.

Solubility of Gases

The Effects of Temperature on the Solubility of Gases:

Why???

Solubility of Gases

Recall the thermodynamics of dissolving:

Very minimal
in the case of
gases!! 


$$\Delta S < 0$$

Solubility of Solids and Liquids

9/4/2008

56

Solubility of Gases

The Effects of Pressure on the Solubility of Gases:

Henry's Law - *The solubility of a **sparingly soluble** gas is directly proportional to the partial pressure of the gas above the liquid.*

$$S \propto P$$

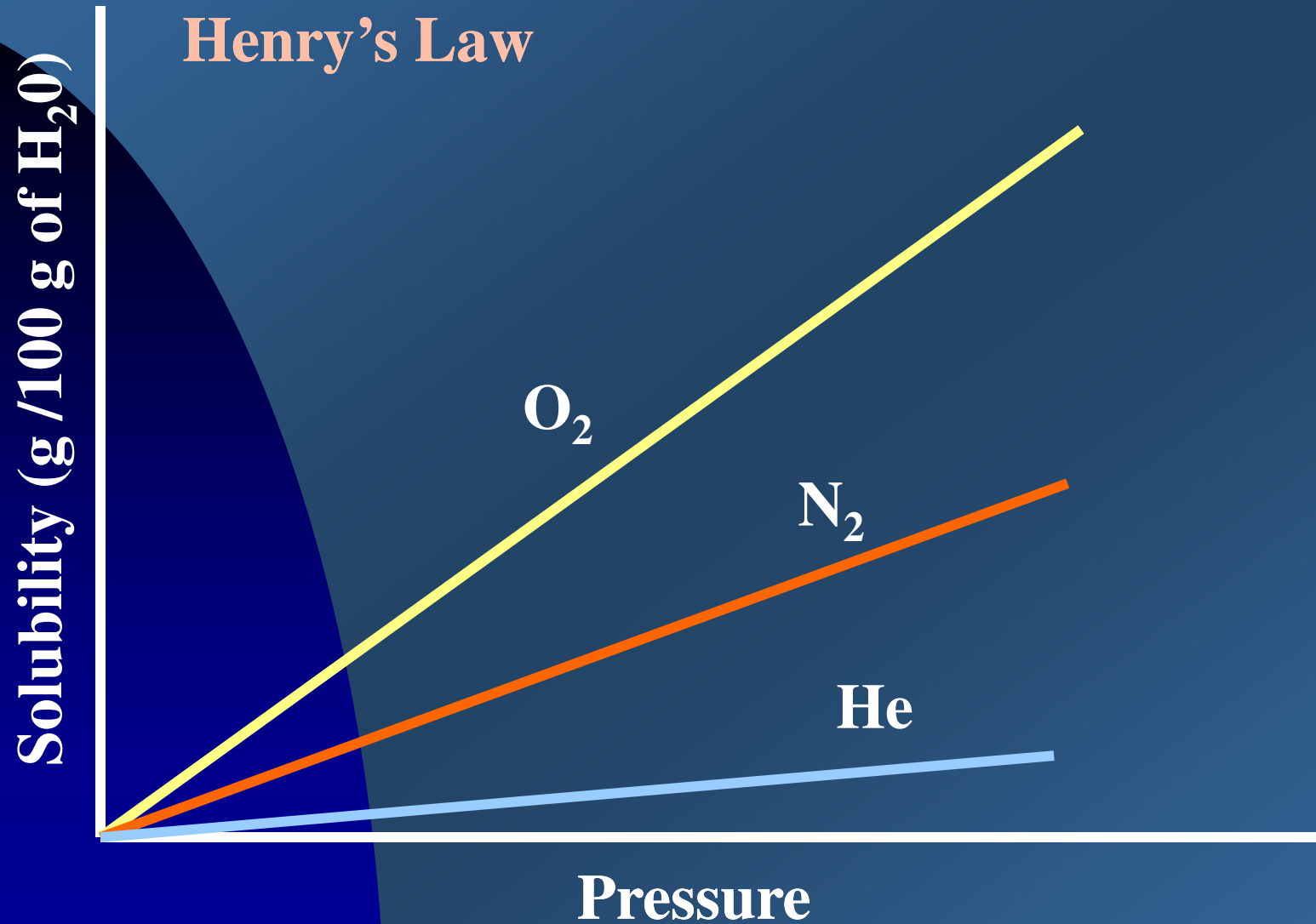
$$S = kP$$

Le Chatelier's Principle can be used to predict the effects of pressure on the solubility of a gas in a liquid.



Solubility of Gases

Henry's Law



Solubility of Gases

What happens when you open a soft drink?

Colligative Properties of Solutions

Definition: A property of solutions that depends only on the concentration (number) of solute particles and not on the nature of the solute.

The same number of particles of different solutes in a given solvent will produce the same change in a *colligative property*.

- *Vapor Pressure Lowering (VPL)*
- *Boiling Point Elevation (ΔT_{bp})*
- *Freezing Point Depression (ΔT_{fp})*
- *Osmotic Pressure (Π)*

Colligative Properties of Solutions

Vapor Pressure Lowering



Non-volatile solute

Figure on Page 542

*Demonstrates that the
Vapor pressure of the
Solution is LOWER than
That of the pure solvent.*

Colligative Properties of Solutions

Ideal Solution: An solution in which the component molecules are subject to forces that are **Identical** to those they would experience in the pure state.

A. The volume of an ideal solution is the sum of its pure component volumes; there is no expansion or contraction on mixing.

B. The $\Delta H_{\text{solution}} = 0$; mixing is neither **exo**thermic nor **endo**thermic.

C. The vapor pressure above an ideal solution is given by Dalton's Law of Partial Pressures.

Colligative Properties of Solutions

Raoult's Law: The vapor pressure of a solution that contains a non-volatile solute is directly proportional to the mole fraction of the solvent in the solution.

$$P_A \propto X_A \qquad P_A = X_A P_A^0$$

“Ideal solutions” containing more than one volatile component also obey Raoult's Law.

$$P_A = X_A P_A^0 \qquad P_B = X_B P_B^0$$

$$P_{\text{total}} = P_A + P_B = X_A P_A^0 + X_B P_B^0$$

Colligative Properties of Solutions

The vapor pressure of water is 23.76 mm Hg at 25°C. A nonvolatile, nonelectrolyte that dissolves in water is glucose .

Calculate the vapor pressure of the solution at 25°C when 7.888 g of glucose, $C_6H_{12}O_6$ (180.158 g/mol), are dissolved in 232.3 g of water.

Water's Molar Mass =18.0153 g/mol.

$$P_A \propto X_A$$

$$P_A = X_A P_A^0$$

Colligative Properties of Solutions

Boiling Point Elevation: Because of VPL, a solution containing a non-volatile solute will boil at a **higher** temperature than the pure solvent.

$$\Delta T_{bp} \propto \text{molality of the solute}$$

$$\Delta T_{bp} = k_{bp} m$$

k_{bp} - *boiling point elevation constant*

<u>Solvent</u>	<u>K_{bp}</u>
H ₂ O	0.512 °C/molal
Ethanol	1.22 °C/molal
Cyclohexane	2.79 °C/molal
HC ₂ H ₃ O ₂	3.07 °C/molal

Colligative Properties of Solutions

Practice Problem: Calculate the boiling point of a solution containing 0.600 kg of CHCl_3 and 42.0 g of eucalyptol, $\text{C}_{10}\text{H}_{18}\text{O}$. Chloroforms b.p. = 61.2°C and its $k_{\text{bp}} = 3.63^\circ\text{C/molal}$.

1. What is the solvent? CHCl_3
2. Calculate the molar mass of the solute?

$$\begin{array}{r} 10 \text{ C @ } 12.011 \text{ u} = 120.11 \text{ u} \\ 18 \text{ H @ } 1.00794 \text{ u} = 18.1429 \text{ u} \\ 1 \text{ O @ } 15.9994 \text{ u} = 15.9994 \text{ u} \\ \hline 1 \text{ C}_{10}\text{H}_{18}\text{O} = 154.25 \text{ u} \end{array}$$

Colligative Properties of Solutions

3. Calculate the molality of the solute.

$$\frac{42.0 \text{ g } \cancel{\text{C}_{10}\text{H}_{18}\text{O}}}{0.600 \text{ kg CHCl}_3} \left| \frac{1 \text{ mol } \text{C}_{10}\text{H}_{18}\text{O}}{154.25 \text{ g } \cancel{\text{C}_{10}\text{H}_{18}\text{O}}} \right. = 0.453\text{g m}$$

4. Calculate the ΔT_{bp} .

$$\Delta T_{\text{bp}} = k_{\text{bp}} m = (3.63 \text{ }^\circ\text{C/m})(0.453\text{g m}) = 1.64\text{7 }^\circ\text{C}$$

5. Calculate the boiling point of the solution.

$$T_{\text{bp solution}} = T_{\text{bp solvent}} + \Delta T_{\text{bp}} = 61.2 \text{ }^\circ\text{C} + 1.64\text{7 }^\circ\text{C}$$

$$= 62.8\text{ }^\circ\text{C}$$

Colligative Properties of Solutions

Freezing Point Depression: A solution containing a non-volatile solute will freeze at a lower temperature than the pure solvent.

$\Delta T_{\text{fp}} \propto$ molality of the solute

$$\Delta T_{\text{fp}} = k_{\text{fp}} m$$

k_{fp} - *freezing point depression constant*

<u>Solvent</u>	<u>K_{fp}</u>
H ₂ O	1.86 °C/molal
Ethanol	1.99 °C/molal
Cyclohexane	20.4 °C/molal
HC ₂ H ₃ O ₂	3.90 °C/molal

Colligative Properties of Solutions

Practice Problem: A 2.00 m solution of sugar in water was found to have a freezing point of -3.70°C . What is the k_{fp} for water based on these data?

$$\Delta T_{\text{fp}} = k_{\text{fp}} m \quad k_{\text{fp}} = \Delta T_{\text{fp}} / m$$

$$k_{\text{fp}} = [0.00^{\circ}\text{C} + (-3.70^{\circ}\text{C})]/2.00 \text{ m}$$

$$k_{\text{fp}} = -3.70^{\circ}\text{C}/2.00 \text{ m} = \boxed{-1.85^{\circ}\text{C}/\text{m}}$$

What would the k_{fp} be for a 1.23 m solution of sugar in water? It would be $-1.85^{\circ}\text{C}/\text{m}$. *The k_{fp} is constant for a given solvent.*

Colligative Properties of Solutions

If k_{bp} or k_{fp} of a solvent is known and the ΔT_{fp} or ΔT_{bp} is measured for a solution containing a known mass of solute, the molar mass of the solute can be calculated.

$$\Delta T_{fp} = k_{fp} m = k_{fp} \frac{\text{g solute}}{\text{molar mass of solute} \times \text{kg solvent}}$$

$$\text{molar mass of solute} = \frac{(k_{fp}) (\text{g solute})}{(\text{kg solvent})(\Delta T_{fp})}$$

“When in doubt, calculate moles!”

Phase Diagrams

Colligative Properties of Solutions

Osmotic Pressure

$$\pi V_{\text{solution}} = n_{\text{solute}} RT$$

*“When in doubt,
calculate moles!”*

Colligative Properties of Solutions

Osmotic Pressure is very important in living systems!

Colloids

Solutions: Homogeneous mixtures containing “dissolved” solute particles of molecular, atomic, or ionic size ~ <1000 pm

- ☞ **Transparent.**
- ☞ **Stable with respect to settling.**
- ☞ **Particles diffuse rapidly.**
- ☞ **Particles pass through ordinary filters.**
- ☞ **Solutions exhibit colligative properties.**

Colloidal Dispersions: Heterogeneous mixtures containing particles intermediate in size between solutions and suspensions

~ 100,000 pm > colloids > 1000 pm

Colloids

- ☞ Particles scatter light (Tyndall Effect).
- ☞ Stable with respect to settling.
- ☞ Particles diffuse slowly.
- ☞ Particles pass through ordinary filters.
- ☞ Dispersions do NOT exhibit colligative properties.

Suspensions: Heterogeneous mixtures containing relatively large dispersed particles ~ >100,000 pm

- ☞ Translucent to opaque
- ☞ Unstable with respect to settling.
- ☞ Particles diffuse very slowly.
- ☞ Particles are easily filtered out.
- ☞ Suspensions do NOT exhibit colligative properties.

Colligative Properties of Solution

Surface Active Agent - “Surfactant”: A material that consists of molecules having both a hydrophobic part and a hydrophilic part.