CHEMISTRY NOTES – CHAPTERS 17 AND 18

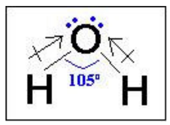
Water and Aqueous Systems - Solutions

Goals : To gain an understanding of :

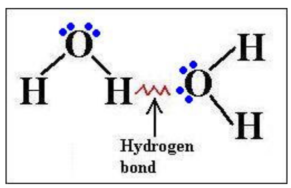
1. The chemical and physical properties of water.
2. Aqueous solutions and their concentrations.
3. Electrolytes and nonelectrolytes.
4. Colligative properties of solutions.

NOTES:

The water molecule is formed by the covalent bonding of two hydrogen atoms and one oxygen atom. It is a bent triatomic molecule because of two unshared pairs of electrons causing a bond angle of about 105º. It is a polar molecule because the oxygen atom has a greater electronegativity than the hydrogen atoms.



Hydrogen bonds form between water molecules because of the attraction of the positively charged hydrogen of one molecule attracting and bonding to the unshared pairs of electrons of another water molecule.



The polarity of the water molecule and the resulting hydrogen bonds create the following properties of water :

- liquid state - water has a lower molar mass (18.02 g/mol) than many other atmospheric gases (CO_2 - 44.01 g/mol, O_2 - 32.00 g/mol, N_2 - 28.02 g/mol), yet it is a liquid and not a gas due to the polarity and bonding of water molecules
- an excellent solvent - the polarity of the water molecule causes it to bond to many solvents (particularly ionic and polar molecular) and dissolve them
- high surface tension - the inward force that minimizes the surface area of a liquid as water molecules attract and bond to each other - causes water droplets to be as spherical as possible and small objects such as needles to be able to "float" on water (they really do not float - it is not the buoyant force holding them up - it is the surface tension. (its kind of like cheerleaders when they catch or hold up another cheer leader by interlocking their hands).
- low vapor pressure - because the water molecules are H-bonded to each other few escape to create a vapor pressure
- high specific heat - it takes a lot of energy to change the temperature of water because much of the heat energy put into water goes to breaking the hydrogen bonds and not just increasing the average kinetic energy (temperature) of the water molecules.
- high heat of vaporization - it takes a lot of energy to cause water to vaporize
- high boiling point - because of the bonding of water molecules to each other it takes a lot of energy to get enough water molecules to escape and create a pressure equal to the external pressure on the water.

Surface tension - the inward force that minimizes the surface area of a liquid as water molecules attract and bond to each other.

Surfactants (from *surf*ace *act*ing *a*gents) are substances which interfere with the hydrogen bonding that occurs between water molecules. These are also called "wetting agents" because they allow water molecules to more easily separate and then attach to or absorb into other substances.

Surfactants are also produced by our bodies to keep the lungs from collapsing. imagine how the insides of a wet plastic bag like to stick together. Now imagine the alveoli (tiny air sacs of lungs). they have to be moist to allow gases to dissolve and be transported. Why don't they stick together? Because of the surfactants produced by the body. Our bodies do not produce the surfactants until late during pregnancy which can create a syndrome called Respiratory Distress Syndrome in premature babies.

Two examples of surfactants are detergents and soaps. These substance help wet other substances as explained above and are used to clean materials. Soaps and detergents have hydrophobic (nonpolar end which may bond or attract to nonpolar substances such as oils) and a hydrophilic end (polar and so will attract and bond to water molecules) which bond dirt and oils to water in the process of cleaning.

Water, like most substances, becomes more dense as it cools. Water is most dense at 4 ºC and then expands as it becomes colder and freezes at 0 ºC. As water freezes the molecules arrange themselves into a crystalline structure which occupies more space making ice less dense than liquid water.

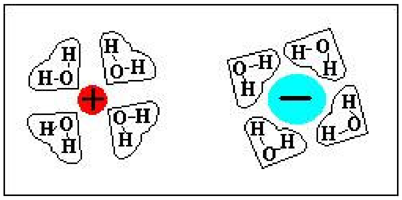
Aqueous solutions are solutions in which water is the solvent.

Solvents are substances which dissolve other substances.

Solutes are the substances which are dissolved in a solution.

Because water is polar the substances that are most likely to dissolve will have charges to bond them to water molecules. This makes ionic substances and polar covalent compounds the most soluble substances in water.

Solvation is the process by which charged solute particles become surrounded by water molecules attracted to the charge of the solute particle. Water then forms a shell like structure around the solute particle (cations or anions or polar molecular compounds) holding it in solution.



Not all ionic compounds will dissolve in water. Those ionic compounds in which the ions within the compound have a greater attraction for each other than for the charged water molecules will not dissolve in water.

Many molecular substances will not dissolve in water because they are nonpolar and will not be attracted to water molecules due to their lack of charge.

Some substances will form crystals which have water molecules as part of their crystalline structure. These water molecules are called the water of hydration or water of crystallization. Hydrates (substances containing the water of hydration) have formulas such as $CuSO_4{\cdot}5H_2O$ where the dot in the formula represents the water of hydration bonded to the substance in the crystal.

Anhydrous substances are substances which lack the water of hydration. Crystalline hydrated compounds, if heated, will lose the water of hydration and become non-crystalline anhydrous compounds.

When a hydrated compound loses its water of hydration it is said to effloresce (the phenomenon is called efflorescence). This will happen when the water vapor pressure inside the crystal is greater than the water vapor outside of the crystal. The water vapor pressure in side a crystal can easily be raised by heating it which will drive out the water molecules. When the water molecules are lost the compound will become anhydrous and non-crystalline.

Hygroscopic compounds are compounds with a very low vapor pressure and will absorb water vapor from the surrounding environment. Two types of hygroscopic compounds are desiccants and deliquescents. Desiccants absorb moisture from the air, but remain in the solid state. Deliquescent compounds absorb enough moisture to form a liquid solution.

Dessicants are used as drying agents. These are contained in the small capsules or packages found in packaging and are used to prevent moisture from damaging sensitive items such as computers and other electronic equipment in storage. A common desiccant is anhydrous silica gel. It will become hydrated as it absorbs moisture in the packaging.

Electrolytes are substances which will conduct electricity when dissolved or melted. All ionic compounds are electrolytes (the ions will carry the current) and some very polar molecular compounds may be electrolytes.

Nonelectrolytes are substances which do not conduct electricity when dissolved or melted. Molecular compounds are generally nonelectrolytes.

Weak electrolytes form few ions in solution and conduct electricity poorly. In weak electrolytes most of the solute exists in the undissociated form. Strong electrolytes form many ions in solution and conduct electricity well. In strong electrolytes most of the solute will dissociate into ions.

Solutions are homogeneous mixtures in which the solute will not settle out solution over time. In solutions the solute exists as atoms, ions or small molecules. solutions can exist as solids (alloys), liquids or gases.

Suspensions are temporary mixtures of one material suspended in another. suspensions are heterogeneous and have relatively large particles which will settle out of suspension over time.

Colloids are permanent mixtures in which the particles are of intermediate size between solutions and suspensions and are borderline homogeneous.

The dispersal medium of a colloid is the substance in which the particles are dispersed and the dispersed phase is the particles. For example smoke is a colloid. the air is the dispersal medium and the smoke particles are the dispersed phase.

Colloids and suspensions produce the Tyndall effect (the scattering of light as it passes through the medium). Solutions do not exhibit the Tyndall effect because the particles are smaller than the wavelengths of visible light and therefore do not scatter the light.

Brownian motion is the random, vibrating motion of colloidal particles. If you were to view colloidal particles under the microscope you would see them vibrate and jitter around. this is due to their being struck by the particles of the dispersal medium (usually water).

Here are a couple of nice little java programs showing Brownian motion
http://www.geocities.com/SiliconValley/Vista/6175/brownian.html or http://www.genevue.com/A_Diffus/Jav1_2.html

An emulsion is a colloid in which both the dispersed phase and the dispersal medium are liquids. An emulsifying agent is a substance which acts to cause an emulsion to form. For example oil and water are not miscible (will not dissolve in each other). Soap is an emulsifying agent which acts as a bridge to hold the oil and water together to form an emulsion.

Qualitatively we can describe three types of solutions: saturated, unsaturated and supersaturated. Saturated solutions hold as much solute as possible under the given conditions. Unsaturated solutions do not hold as much solute as possible under the given conditions. Supersaturated solutions hold more solute than they can normally hold under the given conditions. As far as examples go your on your own for saturated and unsaturated. An example of a supersaturated solution is in the air. Under certain conditions air can become supersaturated with water vapor. the water vapor will not condense and form droplets due to the lack of condensation nuclei (particles for the water vapor to condense on). These clouds can then be "seeded" with a substance like silver iodide to cause the water vapor to condense and form rain.

Three factors which affect the rate at which a solid solute dissolves are temperature, agitation, and particle size.

- Temperature - In general, the rate of dissolving of solid solutes increases with increased temperature. A few solid solutes will dissolve faster at colder temperatures. This is due to the increased kinetic energy of the particles breaking the bonds holding the solute together and intermixing the particles of solute and solvent together.
- Agitation - stirring or shaking the mixture - this increases the mixing of solute and solvent particles together bringing the solute into contact with more solvent.
- Particle size - The smaller the particle size, the greater the surface area of the solute. the greater the surface area of the solute, the more contact the solute has with the solvent and the faster the solute will dissolve.

Gases are more soluble with colder temperatures and therefore their rate of dissolution will increase with colder temperatures.

Solubility refers to the ability (not rate) of a solute to dissolve in a particular solvent. Solubility is usually describe in units such as 5.0 g solvent per liter of solute at 25 ºC. The solubility of a solute depends on the nature of the solute and solvent. If the ions of an ionic compound have a greater attraction for each other than for the poles of the water molecule the compound will not be soluble in water.

There is a rule in chemistry called "like dissolves like," meaning that polar solvents will dissolve polar compounds and nonpolar solvents will dissolve nonpolar solutes. That is why water will dissolve most ionic compounds and polar molecular compounds, but not nonpolar compounds.

Here is a common place example (at least if you are like me and like spicy Mexican food). After a spicy salsa your mouth burns and while a drink of water feels good, it does not remove the burning sensation. That is because the burning sensation is caused by the chemical capsaicin, a nonpolar solid. It will not dissolve in water, but will dissolve in a nonpolar solvent - something with oil or fat in it. The sour cream will help, or better yet, some good ice cream - not the fat free stuff.

Miscible liquids are liquids which will dissolve in each other. An example of two miscible liquids are water and alcohol (a polar molecular compound). Immiscible liquids will not dissolve in each other. An example of two immiscible liquids are water and oil (nonpolar molecular compound). Oil will not dissolve in water because the water molecules bond to each other more strongly than they do the oil molecules. The bonds between the oil molecules and water molecules are actually stronger than the bonds between the oil molecules which is why oil spreads out to such a thin layer over water.

Two factors which affect the solubility of gases are temperature and pressure.

- Temperature - gases will decrease in solubility with an increase in temperature. gases are generally small, nonpolar molecules. With increased temperature they have an increased kinetic energy which allows them to escape the liquid.
- Pressure - gases will increase in solubility with an increase in pressure and vice versa.

Both of the above factors can be easy to remember if you think of what happens to a cold soda. As you soon as you open it (decrease in pressure) gases (mostly carbon dioxide) start escaping. If allowed to become warm the soda becomes flat due to the loss of dissolved carbon dioxide.

Thermal pollution is heat pollution caused by man's activities. An example would be the release of hot water used to cool a power plant. If bodies of water experience thermal pollution the increase in temperature causes a decrease in the solubility of oxygen gas which can be fatal if the concentration of oxygen becomes low enough. This can occur naturally as well. If shallow bodies of water in places like Florida become too warm there can be a die off of animal life due to the lack of oxygen in the water.

Henry's law states that the solubility (S) of a gas in a liquid at a given temperature is directly proportional to the pressure (P) of the gas above the liquid. The equation for Henry's law is $S_1/P_1 = S_2/P_2$.

Chromatography is a process of separation the components of a mixture based on their physical and chemical properties. Chromatography is a very important tool for analytical chemists. The simplest type of chromatography is paper chromatography. In paper chromatography a small amount of a mixture is put near the bottom of a piece of chromatography paper and allowed to dry. The bottom of the paper then is placed in water. As water moves up the paper by capillary action it will pick up the materials in the mixture. The particles of the mixture will separate out due to their attractions for the paper, the solvent and by their mass.

Concentration is a measurement of the amount of a solute that is dissolved in a given amount of a solvent. Two qualitative descriptions of concentration are dilute solutions and concentrated solutions.

Dilute solutions are solutions which have a small amount of solute dissolved as compared to the amount of solvent.

Concentrated solutions have a lot of solute dissolved as compared to the amount of solvent.

Molarity (M) is concentration expressed as moles of solute per liter of solution (mol/L). It is really nothing more than a ratio of one thing to another (moles of solute to liters of solution). Molarity is calculated by dividing the number of moles of solute by the numbers of liters of **solution,** not solvent. Here is an example of calculating molarity.

What is the molarity (M) of a solution which has 2.5 moles of solute in enough water to form .75 L of solution?

M = mol/L = 2.5 mol/ .75 L = 3.3 M

Sometimes you may need to calculate the number of moles first as shown here.

What is the molarity of a solution formed by dissolving 5.5 g of NaCl in enough water to form 1.2 L of solution?

Starting with the given ratio we will convert g into moles and then divide to solve for molarity :

$$\left(\frac{5.5\,\cancel{g}}{1.2\text{ L}}\right)\left(\frac{1\text{ mol}}{58.44\,\cancel{g}}\right)_{g\text{ - mol}} = \frac{.078\text{ mol}}{1\text{ L}} = .078\text{ M}$$

You will also need to be able to calculate moles of solute (and then grams of solute) and volume of solution from the equation for molarity.

To solve for moles of solute rearrange the equation M = mol/L to get mol = L x M. Mass can then be solved for by multiplying by the molar mass. Here is an example.

How many grams of calcium chloride are present in .50 L of a .50 M (mol/L) solution of calcium chloride?

g solute = L x M x molar mass

$$= (.50\,\cancel{L})(.50\,\cancel{mol}/\cancel{L})(110.98\text{ g}/\cancel{mol}) = .28\text{g}$$

To solve for volume of solution we can rearrange the equation for molarity to get L = mol/M. Here is an example.

How many liters are needed to form a .80 M (mol/L) solution with .50 mol of solute?

$$L = \text{mol/M} = .50\,\cancel{mol}/.80\,\cancel{mol}/L = .63\text{ L}$$

Often, in the laboratory, certain molar concentrations need to be prepared from existing concentrated stock solutions. This is called dilution and the formula to calculate molarity from existing solutions by dilution is : $\mathbf{M_1 \times V_1 = M_2 \times V_2}$. Here is a sample problem using dilution:

To what volume should .45 L of a .50M stock solution be diluted to produce a .050 M solution?

Rearranging to solve for V_2 we get : $V_2 = \dfrac{M_1 \times V_1}{M_2}$

$$V_2 = \frac{(.50\,\cancel{M})(.45\text{ L})}{.050\,\cancel{M}} = 4.5\text{ L}$$

You do not need to have units in liters for dilution problems, it will work equally well with any volume unit. Do be sure volume units are the same if you are calculating for molarity.

Percent concentration is another way to quantitatively express concentration. Percent concentration is the fraction of the solute to solvent ratio as expressed in percent. To calculate the percent composition you dived the amount of solute by the amount of solution and then multiply by 100%. The equation is (solute/solution) x 100%. Units are in percents with solute and solvent units indicated in parentheses. Percent composition can be a ratio of volume of solute to volume of solution or mass of solute to volume of solution.

Here is an example of a volume/volume (v/v) percent composition problem :

What is the percent composition (V/V) of a 250.0 mL solution which contains 5.5 mL of sodium chloride?

$$\% \text{ comp} = \left(\frac{\text{volume solute}}{\text{volume solvent}}\right) \text{x } 100\%$$

$$= \left(\frac{5.5 \cancel{\text{mL}}}{250.0 \cancel{\text{mL}}}\right) \text{x } 100\% = 2.2 \ \% \ (\text{v/v})$$

In the above problem units were the same for volume so no conversions were necessary and it was not necessary to indicate the units of volume in parentheses (% comp. would be the same if it were L/L or mL/mL).

Here is a mass/volume (m/v) percent composition problem :

What is the percent composition (m/v) of a solution made from dissolving 8.8 g of sodium chloride in enough water to form 120.0 mL of solution?

$$\% \text{ comp.} = \left(\frac{8.8 \text{ g}}{120.0 \text{ mL}}\right) \text{x } 100\% = 7.3\% \ (\text{g/mL})$$

Colligative properties are properties that depend on the number of solute particles dissolved in a given mass of solvent. they do not depend on the chemical natures of the solute or solvent. It is important here to remember the differences in the dissolving process of ionic and molecular compounds. 1.0 moles of aluminum chloride ($AlCl_3$) will produce 4.0 moles of solute particles (aluminum chloride will dissociate into an aluminum ion and three chloride ions) and 1.0 moles of glucose ($C_6H_{12}O_6$) will produce only one mole of solute particles. Therefore one mole of aluminum chloride dissolved will have 4 times the colligative effect that one mole of glucose will if they are dissolved in the same amount of water.

Three important colligative properties of aqueous solutions are:

- Vapor pressure lowering - a nonvolatile solute (one that will not vaporize) will lower the vapor pressure of water. This is due primarily to the process of solvation. As water molecules bond to the solvent it would require more energy to break that bond and allow the water molecule to vaporize to create vapor pressure.
- Freezing point depression - solute particles will lower the freezing point of water. This is due primarily to the solute particles interfering with the crystal formation of water as it freezes. A common example of this is putting salt on the roads to melt snow or ice or to prevent ice from forming. Another common example is the antifreeze in a car's engine (also raises the boiling point to prevent overheating in summer).
- Boiling point elevation - nonvolatile solute particles will raise the temperature of boiling water. The forces here are similar to that of vapor pressure lowering. As water molecules bond to solute particles more energy is needed to get them to vaporize and it then takes more heat energy to get the vapor pressure of water to equal the atmospheric pressure to cause the water to boil. A common use of this is when food is cooked salt is added to the water. This not only adds flavor but raises the temperature of the boiling water causing the food to cook faster (although, to a probably insignificant amount).

Molality (m) is concentration expressed as moles of solute per kilogram of solvent (not solution). To calculate molality you divide the number of moles of solute by the number of kilograms of solvent. Here is a sample problem. In this problem the solute is given in grams so it will need to be converted into moles and the mass of the solvent is given in grams which will need to be converted into kg.

What is the molality (m) of a solution which contains 20.2 g of calcium chloride ($CaCl_2$) in 800.0 g of water?

Solution :

$$\underbrace{\left(\frac{20.2 \cancel{\text{g}} \, CaCl_2}{800.0 \cancel{\text{g}}}\right)}_{\text{original ratio}} \underbrace{\left(\frac{1 \text{ mol}}{110.98 \cancel{\text{g}}}\right)}_{\text{g solute to mol solute}} \underbrace{\left(\frac{1000 \cancel{\text{g}}}{1 \text{ kg}}\right)}_{\text{g solvent to kg solvent}} = \frac{.228 \text{ mol}}{1 \text{ kg}} = .228 \text{ m}$$

Mole fraction is yet another way of expressing concentration. Mole fraction is the ratio of moles of the parts of a solution (solute or solvent) to the total number of moles of the solution (solute and solvent). The mole fraction of the solute is the moles of solute divided by the total number *of moles of the solution. The mole fraction of the solvent is the number of moles of the solvent divided by the total number of moles of the solution. Here is an example (Moles are given here, but you will usually have to calculate the number of moles first).*

What are the mole fractions of solute and solvent of a solution made up of 2.2 moles of lithium fluoride and 150.0 moles of water (total moles of solution then is 152.2 moles)?

$$\text{mole fraction of solute} = \frac{2.2\ \cancel{mol}\ LiF}{152.2\ \cancel{mol}} = .014$$

$$\text{mole fraction of solvent} = \frac{150.0\ \cancel{mol}\ H_2O}{152.2\ \cancel{mol}} = .9855$$

The sum of the mole fractions should be very close to wholeness (1). It may be a little off due to the use of significant figures. Note also that there are no units for mole fraction as moles cancel out in the equation.

Boiling point elevation and freezing point depression are directly related to the molal concentration by a constant. these are the molal freezing point constant (K_f = -1.86 º/m) and the molal boiling point constant (K_b = .512 ºC/m). What this means is that for every molal concentration of a nonvolatile solute the freezing point drops by 1.86 ºC and the boiling point raises by .512ºC. To calculate the amount of change in the freezing point or boiling point elevation we simply multiply the constant times the molal concentration ($\triangle T_f = K_f m$ or $\triangle Tb = K_b m$ where $\triangle T_f$ and $\triangle T_b$ are the changes in the freezing and boiling points). Here is an example using a molecular solute (one that does not dissociate into ions upon dissolving).

What is the freezing point of an aqueous 1.2 m solution of sucrose, $C_6H_{12}O_6$, (K_f = -1.86 °C/m)?

$\triangle T_f = K_f m$

$= (-1.86\ °C/\cancel{m})(1.2\ \cancel{m}) = -2.2\ °C$

Pure water freezes at 0 °C and since a solute lowers the freezing point the final answer is subtracted from 0 °C.

Freezing point = 0 °C - 2.2°C = -2.2 °C

Here is an example using an ionic solute - one that will dissociate into ions upon dissolving.

What is the freezing point of a 1.2 m solution of sodiun nitrate ($NaNO_3$)?

The molality of the final solution will be twice that of the given molality since sodium nitrate will dissociate into two ions (a sodium ion and a nitrate ion).

$\triangle T_f = K_f m$

$= (-1.86\ °C/\cancel{m})(2.4\cancel{m}) = -4.5\ °C$

The freezing point will be -4.5 °C.

Types of Mixtures

I. Solutions: A homogeneous mixture of two or more substances in a single phase

A. Soluble
  1. Capable of being dissolved (miscible and immiscible)

B. Solvent
  1. The dissolving medium in a solution

C. Solute
  1. The dissolved substance in a solution

D. Types of solutions
  1. Gaseous mixtures
    a. Air is a solution
  2. Solid solutions
    a. Metal alloys
  3. Liquid solutions
    a. Liquid dissolved in a liquid (alcohol in water)
    b. Solid dissolved in a liquid (salt water)

II. Suspensions

A. Suspension

1. The particles in a solvent are so large that they settle out unless the mixture is constantly stirred or agitated
2. Particles in a suspension are on the order of 1000 nm in diameter
3. Particles in a suspension can be filtered out

III. Colloids

A. Particle size

1. 1 nm to 1000 nm in diameter
   a. Larger than particles in solution
   b. Smaller than particles in suspensions
2. Particles remain suspended by the movement of surrounding molecules
3. Particles are not easily filtered

B. Tyndall Effect

1. Light is scattered when passing through a colloid

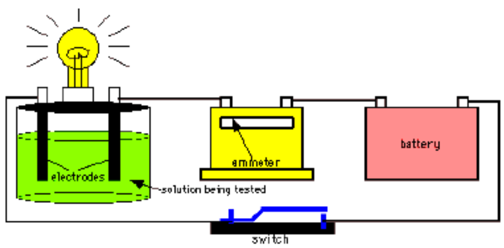
IV. Electrolytes vs. Nonelectrolytes

A. Electrolyte

1. A substance that dissolves in water to give a solution that conducts electric current

B. Nonelectrolyte

1. A substance that dissolves in water to give a solution that does not conduct an electric current



C. Measuring Conductivity

1. Good conductors (strong electrolytes)
   a. Lamp glows brightly
   b. Ammeter registers a substantial current
2. Moderate conductors (weak electrolytes)
   a. Lamp is dull
   b. Ammeter registers a small current
3. Nonconductors
   a. Lamp does not glow
   b. Ammeter does not register a current at all

The Solution Process

I. Factors Affecting the Rate of Dissolution

A. Increasing the Surface Area of the Solute

1. Finely divided substances dissolve more rapidly

B. Agitating a Solution

1. Stirring or shaking brings solvent into contact with more solute particles
2. Added energy temporarily increases solubility

C. Heating

1. Heating always increasing the rate of dissolution of solids in liquids by increasing the KE of particles

II. Solubility

A. Solution Equilibrium

1. The physical state in which the opposing processes of dissolution and crystallization of a solute occur at equal rates

B. Saturation Levels

1. Saturated solution
   a. A solution that contains the maximum amount of dissolved solute

2. Unsaturated solutions
    a. A solution that contains less solute than a saturated solution under the existing conditions
3. Supersaturated Solutions
    a. A solution that contains more dissolved solute than a saturated solution contains under the same conditions

C. Solubility Values
1. The solubility of a substance is the amount of that substance required to form a saturated solution with a specific amount of solvent at a specified temperature
2. The rate at which a substance dissolves does not alter the substances solubility

III. Solute-Solvent Interactions

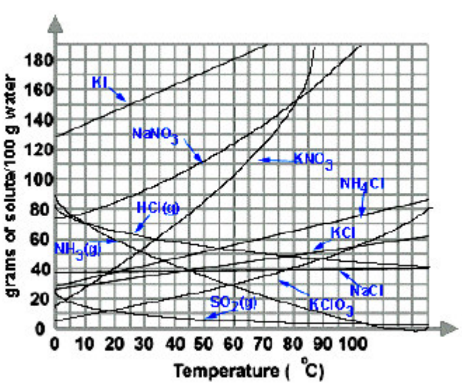
A. "Like dissolves like"
1. Polar substances dissolve in polar solvents
2. Nonpolar substances dissolve in nonpolar solvents

B. Dissolving Ionic Compounds in Aqueous Solutions
1. Electropositive hydrogen of the water molecule is attracted to negatively charged ions
2. Electronegative oxygen of the water molecule is attracted to positively charged ions
3. Hydration
    a. The solution process with water as the solvent
4. Hydrates
    a. Ionic substances that incorporate water molecules into their structure during the recrystallization process $CuSO_4 \cdot 5H_2O$
    b. the " ·" means that the water is chemically attached

C. Nonpolar Solvents
1. Polar and ionic compounds are not soluble in nonpolar solvents
2. Fats, oils and many petroleum products are soluble in nonpolar solvents
3. Nonpolar solvents include CCl_4 and CH_4

D. Liquid Solutes and Solvents
1. Immiscible - Liquid solutes and solvents that are not soluble in each other
    a. Oil and water
2. Miscible - Liquids that dissolve freely in one another in any proportion
    a. Benzene and carbon tetrachloride (both nonpolar)
    b. Water and ethanol (both polar)

E. Effects of Temperature on Solubility
1. Solubility of solids increases with temperature
2. Solubility of gases decreases with temperature



Concentration of Solutions - A measure of the amount of solute in a given amount of solvent or solution

I. Molarity (*M*)

A. Molarity
1. The concentration of a solution expressed in moles of solute per liter of solution

B. Calculations Involving Molarity
1. Determining the molarity of a solution

$$Molarity(M) = \frac{Moles \quad of \quad solute}{Liters \quad of \quad solution}$$

Ions in Aqueous Solutions and Colligative Properties

I. Compounds in Aqueous Solutions

A. Dissociation

1. The separation of ions that occurs when an ionic compound dissolves
   a. One formula unit of NaCl produces two ions
   b. One mole of NaCl produces two moles of ions
      $NaCl\ (aq) \rightarrow Na^{+1}\ (aq) + Cl^{-1}\ (aq)$
   c. One formula unit of $CaCl_2$ produces three ions
   d. One mole of $CaCl_2$ produces three moles of ions
      $CaCl_2\ (aq) \rightarrow Ca^{2+}\ (aq) + 2Cl^{-1}\ (aq)$

II. Precipitation Reactions

A. Solubility Rules

1. No compound is completely insoluble
2. Compounds of very low solubility can be considered insoluble
3. Dissociation equations cannot be written for insoluble compounds

Table 14-1 **General Solubility Guidelines**
1. Most sodium, potassium, and ammonium compounds are soluble in water.
2. Most nitrates, acetates, and chlorates are soluble
3. Most chlorides are soluble, except those of silver, mercury(I), and lead. Lead(II) chloride is soluble in hot water
4. Most sulfates are soluble, except those of barium, strontium, and lead
5. Most carbonates, phosphates, and silicates, are insoluble, except those of sodium, potassium, and ammonium
6. Most sulfides are insoluble, except those of calcium, strontium, sodium, potassium, and ammonium

III. Ionization

A. Ionization

1. Ions are formed from solute molecules by the action of the solvent
2. Polar water molecules are attracted to polar solute molecules
   a. Electronegative oxygen of water is attracted to electropositive portion of a solute molecule
   b. Electropositive hydrogen of water is attracted to the electronegative portion of a solute molecule

Colligative Properties: Properties that depend on the concentration of solute particles but not on their identity

I. Vapor Pressure Lowering

A. Volatility

1. Nonvolatile substances
   a. Substances that have little or no tendency to become a gas under existing conditions (strong IMF)
2. Volatile substances
   a. Substances with a definite tendency to become gases under existing conditions (weak IMF)

B. Effect of Solutes on Vapor-Pressure

1. Any nonvolatile solute will lower the vapor pressure of a solution, having two noticeable effects
   a. Raising the boiling point of the solution
   b. Lowering the freezing point of the solution

II. Electrolytes and Colligative Properties

A. Electrolytes and Solution Concentration

1. Electrolytes dissociate or ionize to form two or more moles of particles for each mole of solute added to solution
   a. One mole of $BaCl_2$ produces 3 moles of ions in solution
      $BaCl_2 \rightarrow Ba^{2+} + 2Cl^{1-}$
2. Electrolytes in water solutions lower the freezing point nearly two, three, or more times as much as nonelectrolytes of the same molality
3. Electrolytes in water solutions raise the boiling point nearly two, or three, or more times as much as nonelectrolytes of same molality