# Intermolecular Forces & Physical Properties of Solutions Syllabus 2014

## I. MAJOR TOPICS

**Intermolecular Forces:**

a. Types of Intermolecular Forces
b. Viscosity, Surface Tension, and Capillary Action
c. Unique Characteristics of Water
d. Different type of Crystal Structures
e. Phase Changes - Heating Curves & Vapor Pressure
f. Phase Diagrams
g. Structures of Solids
h. Bonding in Solids

**Physical Properties of Solutions:**

i. Different Types of Solutions
j. The Solution Process - Molecular Level
k. Concentration units and Conversions

## II. OBJECTIVES/GUIDELINES:

**Intermolecular Forces:**

1. Understand the difference between inTERmolecular forces and inTRAmolecular forces.
2. Be familiar with the three major types of intermolecular forces in terms of relative strength and common examples of each: Ion-dipole, Van der Waals Forces (Dipole-dipole, dipole-induced dipole, **Dispersion forces), and **hydrogen bonding.
3. Be especially familiar with hydrogen bonding and the reasons for why it's an especially strong dipole-dipole interaction. Be able to explain the water and ethanol demo.
4. Understand the relationship between boiling point and strength of intermolecular forces. Be able to put compounds in order of increasing boiling points based on their intermolecular forces.
5. Be able to explain the concept of surface tension and comment on the relationship between high surface tension and intermolecular forces. Be able to site particular illustrations of surface tension such as water droplets beading up on the surface of a freshly waxed car, etc.
6. Be able to compare the menisci of water and mercury based on capillary action (cohesive or metallic bonding vs. adhesive forces).
7. Be familiar with factors that affect surface tension. Think back to the razor blade on water - how did soap and NaCl affect the surface tension of water. How did the surface tension of hexane compare to water?
8. Be familiar with the unique properties of water and how these properties play a role in the natural world around us (water cycle, climate, etc).
9. Most importantly, be able to explain why solid water is less dense than liquid water and the consequences of this pertaining to life on Earth.
10. Understand the differences between a crystalline solid and an amorphous solid.
11. What is Coulomb's Law? What factors determine how strong an ionic or metallic bond will be?
12. Be able to distinguish between the 4 different types of crystals: Ionic, Covalent, Molecular, and Amorphous.
13. Be familiar with all the phase changes and what they look like on a heating/cooling curve.
14. Know what $\Delta H_{vap}$ and $\Delta H_{fus}$ refer to.
15. Understand what vapor pressure is and it's relationship to temperature and intermolecular forces.
16. Have a conceptual understanding of dynamic equilibrium and how it related to vapor pressure. (remember the parking structure).
17. Understand what it means for a substance to boil in terms of vapor pressure.
18. Know what the critical temperature and critical pressure refer to. How can you explain a substance's critical temperature in terms of intermolecular forces? What is a supercritical "fluid"?
19. Be able to create heating/cooling curve calculations/diagrams. For example, be able to determine how much energy is necessary to convert 10 grams of water from -26°C to 156°C taking into account $\Delta H_{vap}$, $\Delta H_{fus}$, and specific heat values. Be careful with your units!

**Physical Properties of Solutions:**

20. Be familiar with the different classifications of solutions. What does it mean for a solution to be super-saturated?
21. Understand how to break down the solution process into three distinct steps and how they sum together to determine $\Delta H_{soln}$. Based on $\Delta H_{soln}$, be able to support the fact that "like dissolves like."
22. Be able to consider pairs of compounds and decide if they will be make a solution based on their intermolecular forces.
23. Be able to define each of the four different concentration units and be able to convert between them (just like the homework problems).
24. Be familiar with how colligative properties are affected by solutes that dissociate into ions (Van't Hoff factor).
25. What is a colloid? Why is a colloid a unique type of solution?
26. Be familiar with common colloids and be able to explain why they fall under this category.
27. Know what the Tyndall effect is and how it can be tested.
28. Understand how big, huge milk proteins are able to be soluble in water. In your explanation, be able to use the terms hydrophobic and hydrophilic.

Complete the following problems from your Brown, LeMay & Bursten chemistry text. Show all of your work! (No Work = No Credit). The answers to the odd numbered problems are in the back of your text. It is your responsibility to get yourself in an academic position to answer ALL of these problems. If needed – PLEASE ASK ME FOR HELP!

**Problem Set #18**: problems 11.5, 11.6, 11.8, 11.14, 11.18, 11.20, 11.26

**Due Date:**________________________________

**Problem Set #19**: problems 11.27, 11.30, 11.35, 11.41, 11.42, 11.50, 11.58, 11.62

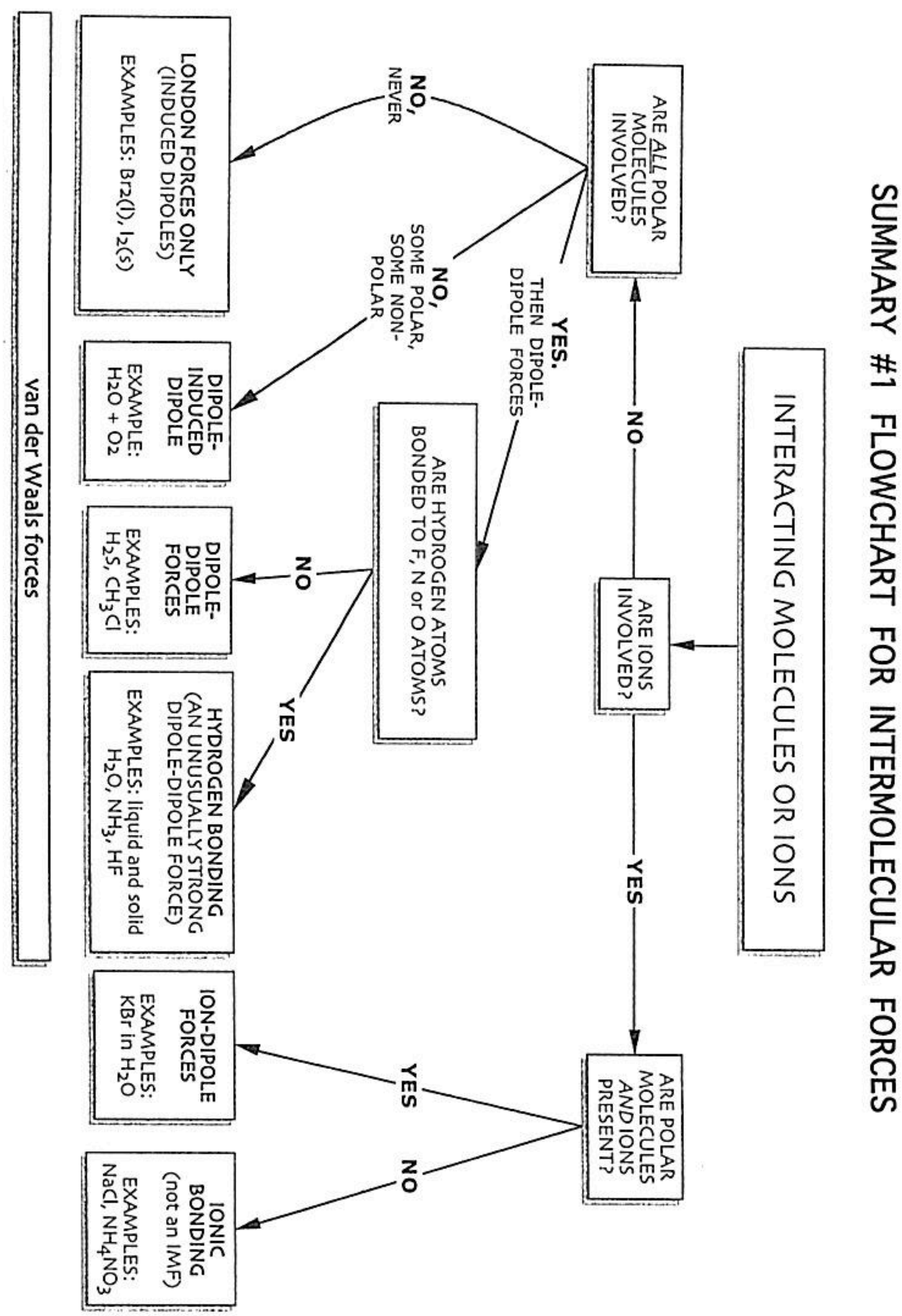
**Due Date:**________________________________

**Problem set #20**: problems 13.2, 13.4, 13.7, 13.13, 13.14, 13.18, 13.20

**Due Date:**________________________________

**Problem Set #21**: problems 13.23, 13.24, 13.28, 13.29, 13.31, 13.32, 13.38

**Due Date:**________________________________



# Notes #37 INTERMOLECULAR FORCES – AP CHEMISTRY

A. **Kinetic Molecular Theory** revisited. . . .
- a gas, liquid and solid . . . all at the same temperature

**How do the phases compare with respect to average Kinetic Energy?_______________________________________________.

**At room temp, why aren't all substances in the same phase?

B. *Inter*molecular Forces vs. *Intra*molecular Forces

1. What are they?

a. **INTERMOLECULAR FORCES** -attractive forces between molecules. These determine many bulk properties of substances such as boiling pt, melting pt., etc.

**The greater the intermolecular forces the ____________________________ the melting and boiling pts, why?

b. **INTRAMOLECULAR FORCES** – attractive forces within molecules. These are the forces that lead to bonding.

**Generally, intermolecular forces are much ____________________than intramolecular forces. It takes a lot less energy to vaporize water molecules (41kJ) than it takes to break the H-O bonds (930kJ) in water.

2. There are three types of INTERMOLECULAR FORCES.

a. **ION-DIPOLE FORCES**- a strong electrostatic attractive force between an ion and a polar molecule.

EX: ions hydrated by water molecules.

$$NaCl(s) \text{ -------------------> } Na^{+} \quad + \quad Cl^{-}$$

b. **van der Waals Forces**- Forces that exist among molecules (no ions involved). There are three different types of van der Waals Forces:

i. **Dipole-Dipole Forces**- a medium strength electrostatic attractive force between two ________________ molecules,

EX:

The strength of a dipole-dipole force depends on the _________________________________of the dipole.

The stronger the dipole the ______________________________________________ the dipole-dipole attraction.

ii. **Dipole-Induced Dipole Forces**- The (+) partially charged poles of a polar molecule attract the electrons in a non-polar molecule, causing the e' distribution to be unsymmetrical- creating or "inducing" a pole in a non-polar molecule.

Ex: Oxygen (which is a _________________________molecule) dissolves in water (polar) so fish can breathe!

iii. **Dispersion Forces**- (London Forces) weak attractive forces that arises as a result of temporary dipoles induced in atoms of molecules. [Kind of like Dipole-Induced Dipole w/out the influence of a polar molecule].

Electrons are always moving randomly in an atom or molecule. Basically, at any instant an atom/molecule may become momentarily polar when there is an unsymmetrical distribution of electrons - resulting in the formation of a dipole. This lopsided atom's dipole will then attract the electrons of another atom (inducing a dipole). The effect will be the attraction of electrons from these resulting dipoles. Etc, etc, etc. ↺ The effect can be quite large! Since **ALL** molecules and atoms have electrons, **ALL** molecules and atoms have dispersion forces!!

However, dispersion forces are the **ONLY** intermolecular forces that affect _________________________and noble gases. It is the dispersion forces that allow non-polar molecules and noble gases to ever become liquids.

THE TAKE HOME QUESTIONS (you should be able to answer these):
Q. What factors affect the strength of dispersion forces (or dipole-induced dipole forces)?
A. The extent to which the electron distribution can be distorted to make a dipole. HINT: This is referred to as the polarizability of the molecule.

Q. What sort of atoms do you think would be the most easily polarized?
A.

PROBLEM 1: Put the following molecules in order of INCREASING boiling points.

$CBr_4$, $CI_4$, $CH_4$, $CF_4$, $CCl_4$

c. **HYDROGEN BONDING**- An especially strong type of dipole-dipole force between a hydrogen atom in a really polar bond (such as N-H, O-H, F-H) and the **O, N, or F** of another molecule.

Ex: Hydrogen bonding among H-F molecules.

*It's the ________________________ on the O, N, and F that attract the partial (+) charge on the H.

Q. What factors cause hydrogen bonding to be such a strong dipole-dipole force?
A1. The significant electronegativity difference between the H and the O, F, or N.
A2. The small size of the O, F, and N atoms (small atomic radii) allow dipoles to get close enough together, creating a stronger attractive force.

PROBLEM 2: Which of the following species are capable of H-bonding to themselves?

KF, HI, $NH_3$, $C_2H_6$, $CH_3OCH_3$, $CH_3COOH$

*HINT*: Ask yourself, "Does this molecule have an H-F, H-O, or H-N bond?"

Take a look at the overhead. Notice the effect hydrogen bonding has according to this graph.
*Normally, we would expect IMF's to increase as we went down a family because . . . .

*However, hydrogen bonding is an unusually strong force and it leads to unusually high boiling points.

PROBLEM 3: Consider the following molecules: $N_2$, $NH_3$, $H_2$, NO
(1) Indicate the different types of intermolecular forces that exist in each of the above molecules.

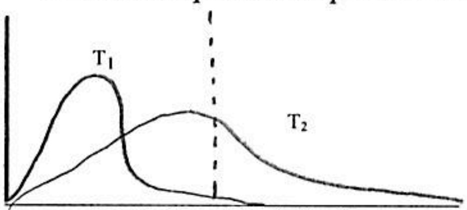
(2) Put the following molecules in order of increasing boiling points. Explain your reasoning.

# Notes #38 PHASE CHANGES AP Chemistry

**Now that we have discussed the properties of each of the states of matter, let's talk about transformations from one physical state to another. PHASE CHANGES!

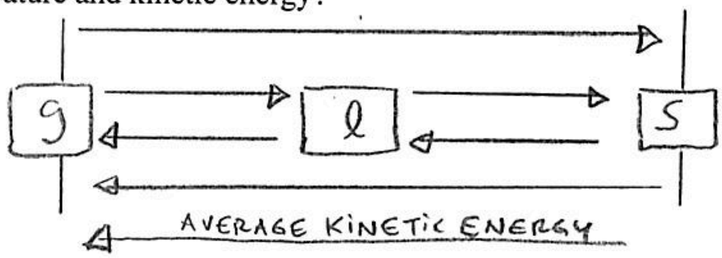
## A. LIQUID-VAPOR EQUILIBRIUM

1. **Evaporation/Vaporization** – conversion of a liquid to a gas.
It's an ________________________process, so energy has to be________________________.
The energy needed to vaporize 1 mole of a liquid is the molar heat of vaporization ____________________(kJ/mol). . . always________________________.
How is the process dependent on temperature and kinetic energy?



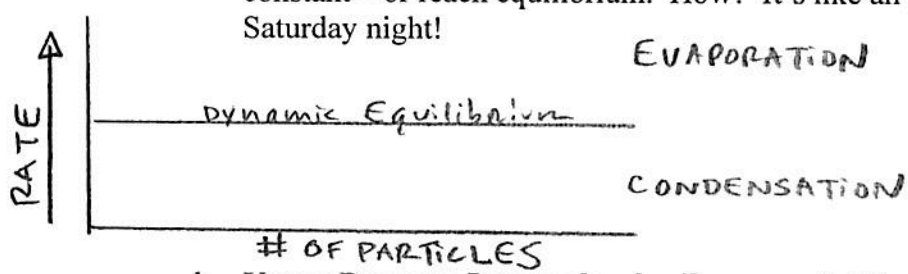

Kinetic Energy Distribution Curve

EVAP




2. **Vapor Pressure** – every liquid has a certain number of particles that are in the gas phase. These gaseous particles create a vapor pressure.
**even at $T_1$

a. How does it work on a molecular level? Imagine yourself placing a liquid in a closed container. The amount of liquid will at first decrease but it will eventually become constant – or reach equilibrium. How? It's like an Ann Arbor parking structure on a Saturday night!




b. **Vapor Pressure Intermolecular Forces and $\Delta H_{vap.}$** Vapor pressure and $\Delta H_{vap}$ are relative measures of the strength of an intermolecular force.

****A low vapor pressure and a high $\Delta H_{vap}$ imply _______________ intermolecular forces. Why?**

Ex: Which of the following would have the lowest vapor pressure? $Cl_2$, $Br_2$, $I_2$

Which of the following would have the lowest $\Delta H_{vap}$? $H_2O$, $H_2S$, $H_2Se$

c. **Vapor pressure and temp**: As you increase the temp., vapor press. will___________.

d. **Vapor pressure and boiling pt**: The boiling point is the temperature at which the vapor pressure of a liquid is EQUAL to the _________________. Could be *any* pressure.

The NORMAL boiling point is the temperature in which liquid boils when the external pressure is _________________________. What's going on at the molecular level?

Can water boil at room temperature? How?

3. **Critical Temperature and Pressure.**

a. There are two ways you can condense a gas 1) 2)

b. Every substance has a CRITICAL temperature ($T_c$) and CRITICAL pressure ($T_p$).

-<u>Critical Temperature</u>: Temperature above which the vapor cannot be liquefied no matter how much pressure is applied. Beyond this temperature, a "fluid" that has properties somewhere between a liquid and a gas (high density, but it's a gas) is formed.

-<u>Critical Pressure</u>: The pressure required to produce liquefaction at the critical temperature. Remember, once you go past the ($T_c$), you will <u>never</u> achieve liquefaction!

Ex:

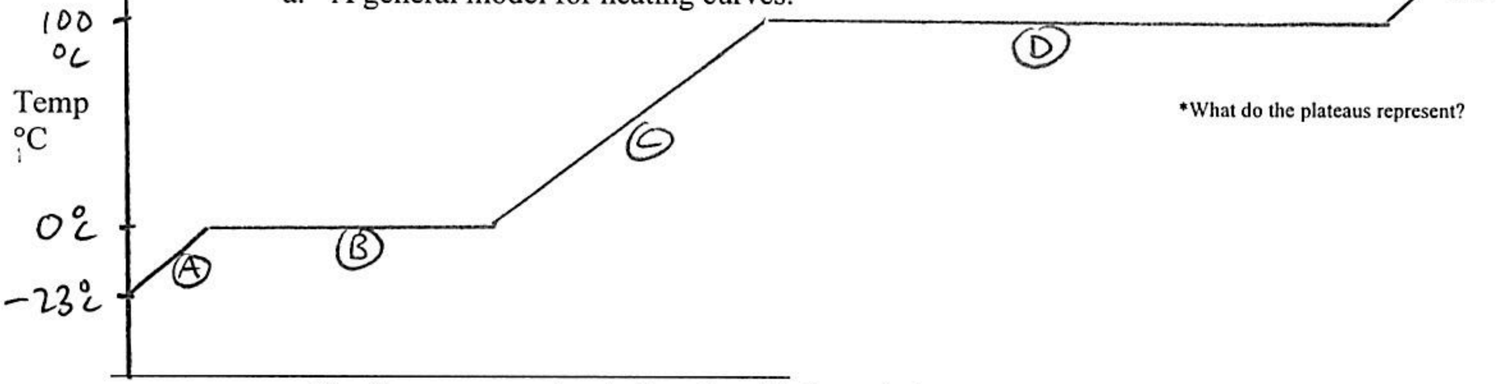
| Chemical | $T_c$ | $T_p$ |
|---|---|---|
| $H_2O$ | 374.4 °C | 219.5 atm |
| $O_2$ | -118.8 °C | 49.7 atm |
| $H_2$ | -239.9 °C | 12.8 atm |

What conclusion can you make about these trends and IMFs?

c. Why do these trends occur? Think of it this way. . . IMFs in a substance are constant. It's the kinetic energy of the substance that changes. So this means . . . . .(?)

4. **Heating Curves** . . .The energies associated with a substance changing phase and heating up or cooling down.

a. A general model for heating curves.




<u>Kinetic energy</u> (gradually and steadily increasing)

time

b. **Using thermochemical equations.** These aid in calculating the heat given off or absorbed during a change of state process.

Ex: How much energy does it take to convert 0.500 kg of ice at -20 °C to steam at 250 °C?

Specific heats: ice = 2.1 J/g°C, water = 4.2 J/g°C, steam = 2.0 J/g°C, $\Delta H_{vap}$ = 40.7 kJ/mol, $\Delta H_{fus}$ = 6.02 kJ/mol

# Notes #39 Properties of Liquids and Solids AP Chemistry 2014

## A. PROPERTIES OF LIQUIDS

1. **Surface Tension**- due to the increase in the attractive forces between molecules at the surface of the liquid compared to the forces between molecules in the center, or bulk, of the liquid. Why does this increased attraction occur *only* at the surface?

HIGH SURFACE TENSION = ________________IMFs

a. Water has an especially high surface tension because of______________________________________.

This explains the following observations:

i. Water "beads up" on a freshly waxed car.

ii. You can fill a glass of water above the rim of a glass.

b. **Capillary Action**- because of high intermolecular forces (surface tension), a liquid can spontaneously rise in a capillary tube (very small diameter glass tube). This is caused by a delicate balance between COHESION and ADHESION.
-**COHESION**- the intermolecular attraction between LIKE molecules (water & water).
-**ADHESION**- the intermolecular attraction between UNLIKE molecules (water & glass).

EXPLAIN THE FOLLOWING ILLUSTRATIONS OF WATER AND MERCURY!

2. **Viscosity**- A measure of a fluids resistance to flow.
Viscosity is used with motor oils.

HIGH VISCOSITY = ________________IMFs
Let us look at some viscosities of common liquids. . .

## B. The wonder of Water:

** Unlike any other compound, each O in a $H_2O$ molecule can make _______ hydrogen bonds (other atoms can only hydrogen bond once). Each O can H-bond using both of its lone pairs. This fact has a couple very important consequences:

1. Because the hydrogen bonding attractive forces are so great, water has an extremely high boiling point and high surface tension for its size.

2. Water, for its size, has an extremely high specific heat (4.184J/C°.g). Big affects on climate.....Absorbs heat in the summer, releases heat in the winter. In other words, prevents rapid temperature changes.

3. Ice floats - Ice is _______________ dense than liquid water.
When water molecules get close enough to become ice, they arrange themselves in a highly ordered 3-D network in which each oxygen atom is tetrahedrally bonded to FOUR hydrogen atoms. ____ hydrogens are bound by covalent bonds and _____ hydrogens are bound by hydrogen bonds. Notice all the empty space in the 3-D structure of ice makes ice less dense! Imagine when ice starts to melt.....A bunch of water molecules have enough kinetic energy to break free of the intermolecular hydrogen bonds. These molecules become trapped in the crystalline cavities of the remaining 3-D ice structure. What happens to the density? It becomes ___________.
Actually, water is most dense at around 4°C. At temperatures above 4°C, the particles are moving too fast and move away from each other, lowering the density

## C. **SOLIDS**

1. Solids can be divided into TWO categories:

a. **Amorphous Solids** - solids that lack a well-defined arrangement. Some don't consider amorphous solids real solids but rather liquids cooled to such low temperatures that they resist flow (high viscosity). They are really solids. Don't believe these kooks!
EX:

b. **Crystalline Solids** - solids that have a rigid, well-defined three-dimensional arrangement. All the atoms, ions, and/or molecules occupy specific places.
EX:

2. There are **FOUR** different Types of Crystals.

1. **IONIC CRYSTALS -**

- COMPOSITION: crystals of ____________ compounds. Really STRONG electrostatic attractive forces hold the ions together into their rigid lattice or crystal structures.

- PHYSICAL CHARACTERISTICS: Hard, brittle, HIGH melting point, poor conductivity of heat and electricity.
- EXAMPLES: Any ionic compound....NaCl, CsCl, $ZnSO_4$, etc
-

SIDE NOTE: Strength of an ionic bond is represented by COULOMB'S LAW: Ionic Bond Strength $\propto \frac{(+q)(-q)}{d^2}$

- Bigger charges = ________________ ionic bond strength - Smaller ionic radii = ____________ ionic bond strength

EX: Circle the ionic compound with the HIGHER melting point in each pair.

NaCl vs $MgCl_2$ $BaCl_2$ vs $MgCl_2$

2. **COVALENT CRYSTALS**

- COMPOSITION: crystals composed of a network of______________________atom.
Really strong covalent bondshold these lattice structures together.

- PHYSICAL CHARACTERISTICS: Hard, HIGH melting point, usually poor conductor of heat and electricity

- EXAMPLES: Carbon's allotropes - Graphite and Diamond, Quartz ($SiO_2$), Tungsten carbide

3. **MOLECULAR CRYSTALS**

- COMPOSITION - crystals of ____________________________compounds. Relatively weak intermolecular forces - van der Waal's forces (Dispersion forces, dipole-dipole, H-bonding) - hold these crystals together.

- PHYSICAL CHARACTERISTICS: Soft, LOW melting points (< 100°C), poor conductors of heat and electricity.

- EXAMPLES: Ice, Sulfur dioxide ($SO_2$), $P_4$, $S_8$ (Another way to think about it. . . .If it's a molecule, it will form a molecular crystal).

4. **METALLIC CRYSTALS**

- COMPOSITION: crystals composed of _______________ atoms. The most simple crystals just because every lattice point is occupied by an atom of the same metal. METALLIC bonds hold these lattice structures together.

** METALLIC bonds - the bonding electrons are delocalized over the entire crystal.
"It is almost as if a bunch of metal cations are immersed in a sea of delocalized valence electrons."

** Since metals have such high densities, _________________________ and ___________________________ crystal structures are preferred over simple cubic.

- PHYSICAL CHARACTERISTICS: Soft to hard, low to high melting points, good conductors of heat and electricity. (think about how these properties can be explained by delocalized electrons)

| Type of Solid | Form of Unit Particles | Forces Between Particles | Properties | Examples |
|---|---|---|---|---|
| Molecular | Atoms or molecules | London dispersion, dipole-dipole forces, hydrogen bonds | Fairly soft, low to moderately high melting point, poor thermal and electrical conduction | Argon, Ar; methane, $CH_4$; sucrose, $C_{12}H_{22}O_{11}$; Dry Ice™, $CO_2$ |
| Covalent-network | Atoms connected in a network of covalent bonds | Covalent bonds | Very hard, very high melting point, often poor thermal and electrical conduction | Diamond, C; quartz, $SiO_2$ |
| Ionic | Positive and negative ions | Electrostatic attractions | Hard and brittle, high melting point, poor thermal and electrical conduction | Typical salts—for example, NaCl, $Ca(NO_3)_2$ |
| Metallic | Atoms | Metallic bonds | Soft to very hard, low to very high melting point, excellent thermal and electrical conduction, malleable and ductile | All metallic elements—for example, Cu, Fe, Al, Pt |

# Notes #40 Colligative Properties/Solutions I

Ap Chemistry

A. TYPES OF SOLUTIONS:

1. What is a *solution*? It's a ________________________________________ of two or more substances.
2. Solutions can often be categorized by the states (solid, liquid gas) of the substances in the solution mixture.

| COMPONENT 1 | COMPONENT 2 | STATE OF SOLUTION | EXAMPLE |
|---|---|---|---|
| Gas | Gas | Gas | |
| Gas | Liquid | Liquid | |
| Gas | Solid | Solid | $H_2$ gas in palladium |
| Liquid | Liquid | Liquid | |
| Solid | Liquid | Liquid | |
| Solid | Solid | Solid | |

3. Solutions are also categorized by their capacity to dissolve a solution.
   a. Unsaturated Solution: solution that contains _______ solute than it has the capacity to dissolve (at a particular temp)
   b. Saturated Solution: solution that contains ______________amount of solute that will dissolve (at a particular temp).
   c. Supersaturated Solution: contains more solute than is present in a saturated solution.
      - supersaturated solutions are pretty unstable and are just waiting for some of the solute to "crash out" in the form of crystals.
      - How do you suppose you make a supersaturated solution?

B. MAKING' SOLUTIONS AND ENERGY

1. We can imagine the solution process occurring in THREE distinct steps.

2. We can think about the above scheme and better understand why "like dissolves like."

| EXAMPLE | $\Delta H_1$ | $\Delta H_2$ | $\Delta H_3$ | $\Delta H_{soln}$ | Soluble? |
|---|---|---|---|---|---|
| NaCl in water | | | | | |
| NaCl in hexane | | | | | |
| oil and water | | | | | |
| water and ethanol | | | | | |
| $CCl_4$ and benzene | | | | | |

3. Let's look at some actual numbers for the dissolving of salt in water.

STEP 1: NaCl ----------> $Na^+(g)$ + $Cl^-(g)$ **$\Delta H_1$ = 786 kJ/mol**

STEP 2 + 3: $H_2O$ (l) + $Na^+(g)$ + $Cl^-(g)$ ---------> $Na^+$ (aq) + $Cl^-$ (aq) **$\Delta H_2 + \Delta H_3 = \Delta H_{hydr}$ -783kJ/mol**

$\Delta H_{soln}$ =

Looking at $\Delta H_{soln}$ for NaCl in water, it's slightly ______________________ (as is true for many dissolving processes).

Why would a solute bother to dissolve if it's an endothermic reaction? What's the driving force?

This second driving force is the inherent tendency towards DISORDER in all natural events. Explain.

EXAMPLE: Would $I_2$ be more soluble in water or carbon disulfide, $CS_2$?

C. CONCENTRATION UNITS: We are familiar with measuring concentration in MOLARITY, however, there are other ways of measuring concentration, each different way having it's own benefits and limitations.

1. Percent by Mass: ratio of the mass of a solute to the mass of the entire solution, multiplied by 100%.

2. Mole Fraction ($X$): ratio of the moles of a solute to the moles of the entire solution.

3. Molarity ($M$): number of moles per 1 liter of solution

4. Molality ($m$): number of moles of solute per 1 kg (1000g) of solvent

EX 1: A solution is prepared by mixing 1.00 g ethanol ($C_2H_5OH$) with 100.0 g water to give a final volume of 101 mL. Calculate the molarity, mass percent, mole fraction and molality of ethanol in this solution.

Can you comment on when certain units of concentration would be preferable over others for certain tasks?

EX 2: (on a separate sheet of paper) You may also be ask to convert between the different concentration units. The electrolyte in automobile lead storage batteries is a 3.75 $M$ sulfuric acid solution that has a density of 1.230 g/mL. Calculate the mass percent and molality.