AP CHEM SUMMER PACKET- Intro
The Content That Will Help You
Get A 5 On The AP Chem Exam

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By the 2017-2018 AP Chem Class*

If you have Mr. Tao, distract him with Fortnite or fishing.

Things You Need To Understand (and that we will try to explain…)

1. Nomenclature - what that one thing was called
2. Stoichiometry, Factor Label and moles - how to math
3. Basic Reactions and balancing with solubility rules
4. Significant figures
5. Kinetics - rate of reactions
6. Acids & Bases - pH, weak vs. strong
7. The Periodic Table - trends, what is what, ionization charges
8. The Atom - sublevels and all that jazz
9. Equilibrium
10. Gas Laws - how to math with gas
11. Intermolecular Forces - types and characteristics
12. Titration and Neutralization - death unit
13. Thermodynamics with Heat, Entropy, and Gibbs Free Energy
14. Electrochemistry - batteries and stuff
15. Spectroscopy- mass or light spec
16. Bonding, lewis dot, and VSEPR theory - characteristics of each bond
17. An in-depth understanding of Luis Amaro's videos (like and subscribe)

https://www.youtube.com/watch?v=mDKcfSm9poc

Aight, so because the curriculum changed, you guys know nothing. This stuff is basically rocket science for you normies. This study guide will teach you the basics and the rest we leave in the hands of Dr. Tao and Dr. Jo. You will also need to have an extensive understanding of Fortnite, fishing, and the philosophy of the Science Fish.

Binary Ionic Compounds (Type I)

- The cation is always first, followed by the anion
- A monatomic cation takes its name of the element
  - Ex. Na^+ is called sodium in the names of compounds containing this ion.
- A monatomic anion is named by taking the root of the element name and adding *-ide*
  - Ex. Cl^- ion is called chloride

Binary Ionic Compounds (Type II)

- Same rules as Type I, except use roman numerals for elements with multiple oxidation states
  - Ex. Iron Chloride can be represented as both $FeCl_2$ and $FeCl_3$ because of the multiple oxidation states of Iron (Fe^{2+} and Fe^{3+})
  - So, the systematic names of these compounds would be Iron(II) Chloride and Iron(III) Chloride, respectively

Binary Covalent Compounds (Type III)

- Formed between two nonmetals
- First element in formula is named first with its full name
- The second element is named and goes in as the anion
- Prefixes are used on both to denote the number of atoms present
  - Mono- (1), Di- (2), Tri- (3), Tetra- (4), etc.
- The prefix Mono- is never used on the first element
  - Ex. NO_2 would be named Nitrogen Dioxide

Polyatomic Ions

- The cation is always first, followed by the anion
- If the cation is a metal ion with a fixed charge, the name of the cation is the same as the regular element
  - Ex. Na^+ is just Sodium in the compound name
- If the cation is a metal ion with multiple oxidation states, the charge on the cation is indicated using a Roman numeral, in parentheses, following the name of the cation
  - Ex. Fe^{3+} would be Iron (III)
- If the anion is a monatomic ion, the anion is named by adding the suffix *-ide* to the root of the element name
  - Ex. I^- is iodide in the compound name

Acids and Bases

- Not Containing Oxygen
- H^+ cation becomes the prefix *hydro-*
- *-ic* gets added to the root name of the anion
  - Ex. HCl would be names Hydrochloric Acid

Naming Oxyacids

- H^+ cation becomes the prefix *hydro-*
- Suffixes are based on the ending of the original name of the oxyanion. If the name of the polyatomic anion ended with -ate, change it to -ic for the acid and if it ended with -ite, change it to -ous in the acid.
  - Ex. HNO_3 , containing the ion nitr**ate**, would be named Nitr**ic** Acid
  - Ex. HNO_2 , containing the ion nitr**ite**, would be named Nitr**ous** Acid

Examples and Answers:

- CuBr …… Copper(I) bromide
- $CuBr_2$ …..Copper (II) bromide
- FeO…...Iron(II) Oxide
- Fe_2O_3…..Iron(III) Oxide

Practice:

Formula	Anion	Acid name	Anion Name
HF	F^- is		
HCl	Cl^- is		
HBr	Br^- is		
HI	I^- is		
H_2S	S^{2-} is		
HNO_3	NO_3^- is		
$HC_2H_3O_2$	$C_2H_3O_2^-$is		
H_2SO_4	SO_4^{2-} is		
H_2CO_3	CO_3^{2-} is		
H_3PO_4	PO_4^{3-} is		

HClO	ClO^- is		
$HClO_2$	ClO_2^- is		
$HClO_3$	ClO_3^- is		
$HClO_4$	ClO_4^- is		
HIO_3	IO_3^- is		
HNO_2	NO_2^- is		
H_2SO_3	SO_3^{2-} is		

I. Write formulas for the following compounds:

1. aluminum bromide
2. copper (I) oxide
3. hydrogen sulfide
4. iodic acid
5. nickel (III) nitride
6. potassium iodide
7. lead (II)
8. magnesium nitride
9. iron (II) fluoride
10. copper (II) bromide
11. silver nitride

II. Name the following compounds

1. H_2CO_3
2. Li_2S
3. NaBr
4. $ZnCl_2$
5. MgO
6. $PbCl_2$
7. $HC_2H_3O_2$
8. K_2CO_3
9. HCl

Stoichiometry

1. Factor Label (converting units)

Treat everything as a ratio and "cancel out" the units to convert.

Ex.

$$2 \text{ ft} \times \frac{12 \text{ inches}}{1\text{ft}} = 24 \text{ inches}$$

Density**:** the mass of an object divided by its volume. It can be used as a conversion factor between mass and volume

Practice:

1. ***What is the density (g/mL) of a sample of mineral oil if 250 mL has a mass of 0.23 kg?***

2. ***Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.71002 g***

Knowing that 13.6 g mercury = 1 mL mercury…

3. ***Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.***

4. ***If you have a 2.130 mL sample of acetic acid with mass 0.002234 kg, what is the density in kg/L?***

5. ***What is the volume of 100.0 g of air if its density is 1.3 g/L?***

6. ***What is the density (g/mL) of a sample of mineral oil if 250 mL has a mass of 0.23 kg? Choose the following….***
   - 0.0092 g/mL
   - 0.92 g/mL
   - 5.8 g/mL
   - 10.9 g/mL
   - 57.5 g/mL

<u>Temperature</u>: a measure of the degree of hotness of an object. It is related to the average molecular kinetic energy of the object. Temperature is not a measure of heat, but heat flows from a hot object to a colder one if the two are in contact.

There are three temperature scales

1. <u>Fahrenheit</u> - mainly used in U.S.A
2. <u>Celsius</u> - formerly centigrade, used by the rest of the world
3. <u>Kelvin</u> - the SI unit for temperature, used mainly by scientists.

Temperature Conversion:

Water can be used as the reference to determine the temperature scales:

- Freezing point of water = 0 ^oC = 32 ^oF = 273.15 K
- Boiling point of water = 100 ^oC = 212 ^oF = 373.15 K

^oF = (9/5)x + 32

Kelvin (K) = ^oC + 273

^oC = $^5/_9$ x (^oF - 32)

Converting kelvin (K) to Celsius (^oC)

^oC = K - 273

<u>Practice</u>

1. ***Gallium is a metal that can melt in your hand at 302.93 K. What is the temperature in C?***

2. ***Body temperature is 98.6 F. What is the temperature in C?***

3. ***Room temperature is often used in calculations as 300 K. What is the temperature in Fahrenheit?***

MATH CONVERSION CHART – WEIGHT (US)

METRIC CONVERSIONS					
1 gram	=	1000 milligrams	1g	=	1000 mg
1 kilogram	=	1000 grams	1 kg	=	1000 g
1 tonne (1 megagram)	=	1000 kilograms	1 tonne (1 Mg)	=	1000 kg
STANDARD CONVERSIONS					
1 ounce	=	16 drams	1 oz	=	16 dr
1 pound	=	16 ounces	1 lb	=	16 oz
1 hundredweight	=	100 pounds	1 cwt	=	100 lb
1 ton (short)	=	20 hundredweight	1 ton	=	20 cwt
1 ton (short)	=	2000 pounds	1 ton	=	2000 lb
METRIC -> STANDARD CONVERSIONS					
1 gram	=	0.035274 ounces	1 g	=	0.035274 oz
1 kilogram	=	2.20462 pounds	1 kg	=	2.20462 lb
1 kilogram	=	35.27396 ounces	1 kg	=	35.27396 oz
1 tonne	=	1.10231 ton, short	1 tonne	=	1.10231 ton, short
STANDARD -> METRIC CONVERSIONS					
1 ounce	=	28.34952 grams	1 oz	=	28.34952 g
1 pound	=	453.59237 grams	1 lb	=	453.59237 g
1 pound	=	0.45359 kilograms	1 lb	=	0.45359 kg
1 hundredweight	=	50.8023 kilograms	1 cwt	=	50.8023 kg

Practice:

1. ***Convert 4.29 ng into grams.***

2. ***Convert 12 gallons into L.***

3. ***Convert 10 mL to L.***

4. ***Convert 82947 L to nanoliters.***

5. ***Convert 7 pounds to grams.***

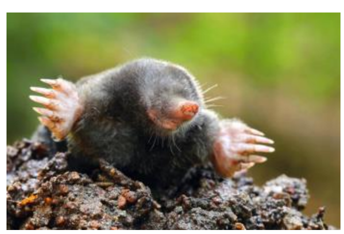
6. ***Convert 40 kg to g.***

Shazil Usmani

2. Moles

What is a mole?

A small, burrowing, insectivorous mammal with dark velvety fur, a long muzzle, and very small eyes.



What is a **chemistry** mole? The amount of mass in one amu of an element. Your periodic table has all the masses in one mole of each atom. For example, if you look at carbon on your periodic table, you'll see that one mole of carbon is 12.01 grams. You're going to use moles in basically every unit.

1 mole = 6.02 x 10^{23} atoms = Atomic mass on periodic table

Ex. How many atoms are in 50.0 grams of Sodium?

$$50 \text{ grams} * \frac{1 \text{ mole Na}}{23 \text{ grams}} * \frac{6.02 * 10^{23} \text{ atoms}}{1 \text{ mole Na}} = 1.31 * 10^{24} \text{ atoms of Na}$$

(NOTE: 50.0 grams only has 3 significant figures. Make sure that your final answer also only has 3 significant figures. Review sig fig rules below.)

<u>Practice:</u>

1. ***How many moles are in 50 grams of nickel?***

2. ***How many molecules are in 4 moles of copper? How many kilograms?***

3. ***How many moles are in 7.03 x 10^{40} atoms of Na?***

3. Types of Reactions

- Synthesis: The combination of two elements to make a compound. A + B → AB

Ex. $2H_2$ (g) + O_2 (g) → $2H_2O$ (l)

- Decomposition: Breaking down a compound into its components.
  AB → A + B

Ex. H_2CO_3 (aq) → CO_2 (g) + H_2O (l)

- Single Displacement: Replacing a metal/cation in a compound with another metal cation as a element. A + BC → B + AC

Ex. Li (s) + NaCl (s) → LiCL (s) + Na (s)

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- Double Displacement: Swapping the metal/cation in the reactants. AB + CD → AD + CB

Ex. 2KI (aq) + $Pb(NO_2)$ (aq) → PbI_2 (s) + $2KNO_2$ (aq)

- Combustion: A reaction between combustible material and an oxidizer to produce an oxidized product. The most common reaction is with a hydrocarbon and oxygen to produce CO_2 and H_2O.
  C_xH_y + O_2 → CO_2(g) + H_2O(g)

Ex. CH_4 (g) + $2O_2$ (g) → CO_2(g) + $2H_2O$(g)

- You will also apply double displacement to precipitate reactions. In a precipitate reaction, two solutions react to form a precipitate and ions in solution. In order to identify which elements form ions and which elements form precipitate, you must know your solubility rules.

Basic Solubility Rules (memorize them or you're screwed)

1. All nitrates (NO_3^{-1}) are soluble. (Ex. $NaNO_3$)
2. All alkali (Group 1) metals and ammonium salts (NH_4^+) are soluble. (Ex. KCl, NH_4Cl)
3. Most halides (Group 17) are soluble (Not with Ag, Hg, and Pb).
4. Most sulfates (SO_4^{2+}) are soluble (Not with Ca, Ba, Hg and Pb).
5. Most hydroxides (OH^-) are **slightly soluble***. NaOH and KOH are soluble. $Ba(OH)_2$ and $Ca(OH)_2$ are **marginally soluble**.
6. S^{2-}, CO_3^{2-}, CrO_4^{2-} and PO_4^{3-} salts are **slightly soluble** unless they apply to the second rule.

*These will still form precipitate.

There are three main ways to describe precipitate reactions: complete balanced reactions, complete ionic reactions and net ionic reactions.

- Complete balanced reactions give the overall reaction, but not the reactants and products as they exist in reaction.
  AB(aq) + CD(aq) → AD(s or l) + CB(aq)

Ex. $2KI\ (aq) + Pb(NO_2)\ (aq) \rightarrow PbI_2\ (s) + 2KNO_2\ (aq)$

- Complete ionic reactions represents all reactants and products as they exist in reaction.
  $A^+(aq) + B^-(aq) + C^+(aq) + D^-(aq) \rightarrow AD(s\ or\ l) + C^+(aq) + D^-(aq)$

Ex. $2K^+\ (aq) + 2I^-\ (aq) + Pb^{2+}\ (aq) + 2NO_2^-\ (aq) \rightarrow PbI_2\ (s) + 2K^+\ (aq) + 2NO_2^-\ (aq)$

- Net ionic reactions represent only the components that produce precipitate; spectator ions (start and end in the reaction as ions, not involved in the reaction) are not included.
  $A^+(aq) + D^-(aq) \rightarrow AD(s\ or\ l)$

Ex. $Pb^{2+}\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s)$

There will also be equilibrium reactions, in **which both the reactants and products are produced at equal rates**. These reactions will use a double arrow ($\leftrightharpoons$), instead of a single arrow ($\rightarrow$).

States in Reactions:

- (s): solid
- (l): liquid
- (g): gas
- (aq): aqueous or hydrated or in water

<u>Practice:</u>

In the following reactions, identify the precipitate and write the complete and net ionic equations including states.

1. ***$Pb(NO_3)_2 + KI \rightarrow PBI_2 + KNO_3$***

2. ***$HCl + NaOH \rightarrow NaCl + H_2O$***

- ***This precipitation reaction is also what kind of reaction?***

3. ***$AgNO_3 + LiCl \rightarrow AgCl + LiNO_3$***

4. Balancing

Just like regular equations, chemical equations have to be equal on both sides. We balance equations so that it is consistent with the Law of Conservation of mass. Let's look at an example:

$$2C_4H_{10} + 13O_2 \rightarrow 10H_2O + 8CO_2$$

In this example you can see that both sides of the equation are balanced. On the reactant side (left) there are 8 moles of Carbon (2 x 4) and that is consistent with the product side (right), where there are also 8 moles of Carbon (8 x 1). The same applies to all of the other elements.

Practice:

1. ***Balance*** $NaBr + Cl_2 \rightarrow NaCl + Br_2$

2. ***Balance*** $C_9H_{20} + O_2 \rightarrow CO_2 + H_2O$

5. Limiting Reagent and Moles to Moles

Limiting Reagent is the idea that you have more than one reactant, so the one reactant that you have less of will run out first. In order words, you can only produce as much product as the limiting reactant can make.

Ex. $2H_2 + O_2 \rightarrow 2H_2O$

(Side Note: Hydrogen, Nitrogen, Oxygen, Fluorine, Chlorine, Bromine, and Iodine are the 7 elements that have a subscript of two in their elemental form. These are called Diatomic elements)

Before I continue with limiting reagent, I'm going to take a moment to explain converting moles. If you look at the equation above, you can see that the ratio between hydrogen to water is 2:2. In other words, for every 2 moles of Hydrogen you have, you will make 2 moles of water.

Practice:

1. ***If you have 8 moles of H_2SO_4 and 2.5 moles of NaOH, which is the limiting reagent?***

2. ***What is the limiting reagent in a reaction of 50 g Li and 50 g Cl?***

3. ***With 9 g of Mg (s) and 9 g of O_2 (g), how much MgO can be produced?***

SIG FIGGGGGGGS! (Frashure)

Significant Figures**:** You'll need to know this for like, every single problem you do, but it will be helpful to know and keep in mind when taking tests to prevent you from rounding too much. These babies are all about that exact number.

Here are the rules...

1. Zeros in between two numbers are always significant. Both 4308 and 40.05 contain four significant figures.
2. Zeros that follow without a decimal point are not significant. Thus, 470,000 only has two significant figures.
3. Trailing zeros that aren't needed to hold the decimal point are significant. For example, 4.00 has three significant figures.

*If a number has a dot behind it, such as 400. °K, that dot counts as a decimal point

Okay you got that? How many sig figs are in the following numbers?

1. 4
2. 367000
3. 5,960
4. 500000.

For **addition and subtraction**, look at the decimal portion of the numbers ONLY! Here is how ya do it:

1. Count the number of significant figures in the decimal portion of each number in the problem. (The digits to the left of the decimal place are not used to determine the number of decimal places in the final answer.)
2. Add or subtract in the normal fashion.
3. Round the answer to the LEAST number of places in the decimal portion of any number in the problem.

The following rule applies for **multiplication and division**:

1. The LEAST number of significant figures in any number of the problem determines the number of significant figures in the answer.

2. This means you MUST know how to recognize significant figures in order to use this rule.

For multiplication or division, remember, an answer can NEVER be MORE precise than your least precise measurement.

Ex. 2.5 x 3.42

The answer to this problem would be 8.6 (which was rounded from the calculator reading of 8.55). Why?
2.5 has two significant figures while 3.42 has three. Two significant figures is less precise than three, so the answer has two significant figures.

Important Note!!!
The AP exam will always use sig figs. For the free response, they will take off points for incorrect sig figs.

<u>Practice</u>

1. The number 535.602 rounded to 3 significant figures is:

2. 1.0 liter is:

3. If a can of soup contains 22.0 oz (ounces) of soup, how many grams of soup is that? (1 lb = 16 oz, 1 lb = 454 g). use correct sig figs

4. What is the metric unit for volume?

5. A sample contains 430 mg of mercury. How many significant figures are in the number?

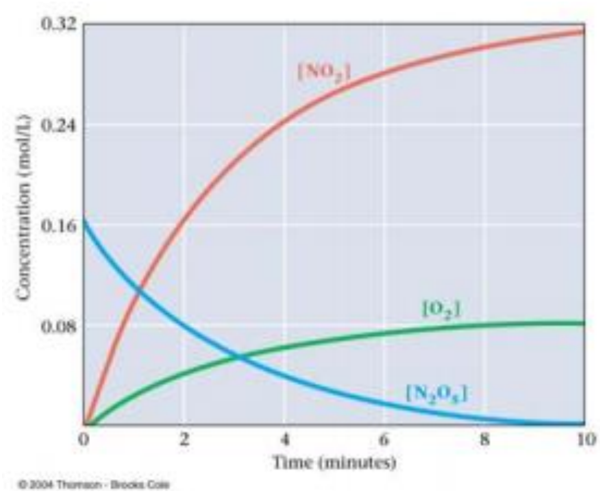
7. A length of glass tubing is 0.525 m. How many inches long is the tubing? (2.54 cm = 1 inch)

Kinetics (Barry)

Consider the following decomposition reaction: $N_2O_5(g) \rightarrow 2\ NO_2(g) + \frac{1}{2}\ O_2(g)$

As this reaction proceeds over time (towards completion), the concentration of N_2O_5 decreases and the concentration of NO_2 and O_2 increases. Reactant molecules are consumed and product molecules are formed as a result. These three concentrations are constantly changing until eventually they reach a plateau, they reach a chemical equilibrium. Notice that each substance changes concentration at a different rate. When one mole of N_2O_5 decomposes, two moles of NO_2 are formed and one-half mole of O_2 is formed.

We can determine the reaction rate by monitoring a change in concentration (of a reactant or product). The most useful unit to use for rate is molarity. Since volume is constant, molarity and moles are directly proportional.



Rate (the speed of a reaction) is not constant, it changes with time.

This can be used to define **reaction rate**: ***rate = concentration/time***

Average rate: slope of secant line, change in concentration over a change in time

Instantaneous rate: slope of tangent line, change in concentration at an instance

Example:

Refer to the graph above. In a time span of 4 minutes, the concentration of $N_2O_5(g)$ goes from 0.160 M to 0.040 M. Calculate the average rate.

$$\text{rate} = \frac{0.040M - 0.160M}{4\text{ min} - 0\text{ mins}} = -0.030 \text{ mol L}^{-1}\text{ min}^{-1}$$

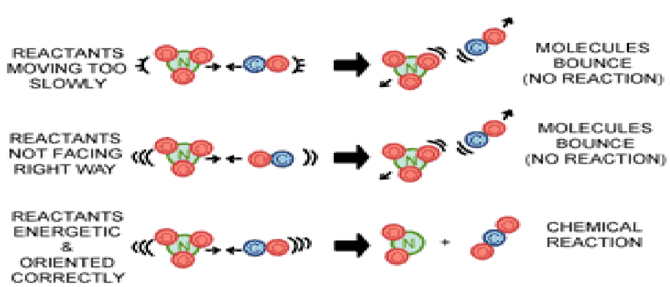
Rate of Δ[reactant] is always negative and rate of Δ[product] is always positive.

Factors that affect reactions rates:

- **Physical state of the reactants** - solids react much more slowly than liquids or aqueous solutions (can be explained using IMF)
- **Concentration of the reactants** - more reactant molecules present, the more collisions occur per unit time, the faster reaction proceeds, the greater its rate
- **Temperature of the reaction** - temperature is proportional to the average kinetic energy of the molecules, higher temperatures, faster molecules, the more likely they are to collide and the more energetic the collisions become
- **Presence of a catalyst** - a catalyst is used to increase the rate of a reaction without being consumed by it. This occurs by altering the reaction pathway, by doing so lowers the activation energy needed for the reaction to proceed, therefore, more collisions are successful and the reaction rate is increased.

Brief Summary of this thing called the Collision Theory:

- Particles must collide.
- Most often, only two particles may collide at one time.
- Proper orientation of colliding molecules is essential so that molecules come into contact with each other to form products.
- Collisions must occur with enough energy - effective energy



MAJOR CONCEPT!!!!!!!

*When delving into the topic of chemical kinetics, especially rate laws, remember this: The order of a reaction must be **determined experimentally**; it cannot be deduced from the coefficients in the balanced equation!*

Reaction mechanism: is a series of steps that represents the pathway of the reaction, they *predict* what takes place at each stage of an overall chemical reaction. Most reaction mechanisms involve two or more elementary steps. Some steps may be fast and other steps may be slow. A reaction can only proceed as fast as its rate determining step - the slowest step in a mechanism. Elementary steps are the individual steps that constitute a reaction mechanism, they can be either uni,bi, or tri molecular (in AP). These prefixes are determined by the number of reacting molecules in the step.

Bimolecular: 2 A -> B rate = $k[A]^2$, in and only in mechanisms, the number of reacting molecules is equivalent to the rate order (will discuss this in chapter 12)

Consider this reaction mechanism:

$$2NO \xrightarrow{fast} N_2O_2 \qquad \text{Elementary step 1}$$

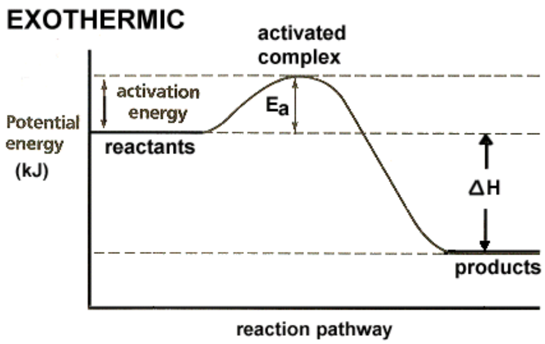
$$N_2O_2 + H_2 \xrightarrow{slow} N_2O + H_2O \qquad \text{Elementary step 2}$$

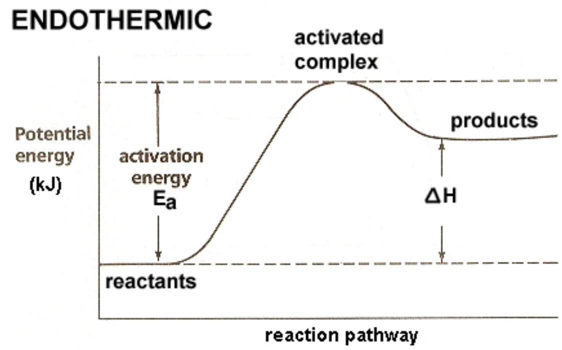
$$N_2O + H_2 \xrightarrow{fast} N_2 + H_2O \qquad \text{Elementary step 3}$$

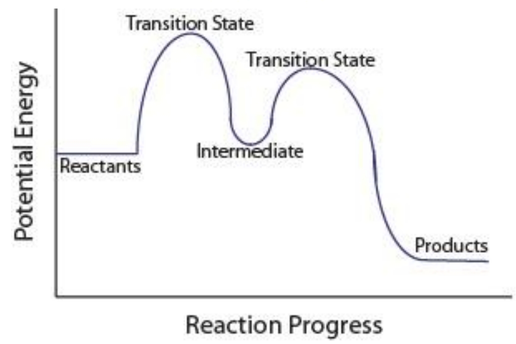
The overall rate of the reaction cannot exceed the rate of the slowest step The rate of the slowest step will be approximately equal to the rate of the overall reaction. The slowest step in a mechanism is usually the one that has the highest activation energy.

Sometimes the mechanism will involve **intermediates**. If a chemical appears as a product in one step and then again as a reactant in another step, and it can be crossed out, then it is an intermediate. Sometimes the mechanism will involve a **catalyst**. If a chemical appears as a reactant in one step and then again as a product in another step, and it can be crossed out, then it is a catalyst.

Potential Energy Diagrams - Get Familiar to seeing these kinds of graphs in kinetics. These diagrams can give us information on whether a reaction is exothermic or endothermic, activation energy (and complex).







<u>Practice</u>: (I will teach you all how to do these, but you can give them a try now)

Answer the following, while considering the mentioned mechanism:

$$2\mathrm{NO} \xrightarrow{fast} \mathrm{N_2O_2} \qquad \text{Elementary step 1}$$

$$\mathrm{N_2O_2} + \mathrm{H_2} \xrightarrow{slow} \mathrm{N_2O} + \mathrm{H_2O} \qquad \text{Elementary step 2}$$

$$\mathrm{N_2O} + \mathrm{H_2} \xrightarrow{fast} \mathrm{N_2} + \mathrm{H_2O} \qquad \text{Elementary step 3}$$

1. ***What is the overall reaction (reaction excluding the intermediates and catalyst)?***

2. ***What is the rate determining step, and the rate law (slow step)?***

3. ***What are the intermediates in this reaction?***

4. Is there a catalyst used in the reaction?

5. Draw the potential energy diagram for the mechanism:

Practice (Continued):

1.) Answer the following while considering the mentioned mechanism:

$2NO_2(g) \rightarrow NO_3(g) + NO(g)$ (slow)

$CO(g) + NO_3(g) \rightarrow CO_2(g) + NO_2(g)$ (fast)

1. What is the overall reaction (reaction excluding the intermediates and catalyst)?

2. What is the rate determining step, and the rate law?

3. What are the intermediates in this reaction?

4. Is there a catalyst used in the reaction?

5. Draw the potential energy diagram for the mechanism:

2.) Answer the following considering the mentioned mechanism: ***(Malcolm Jennings)***

k1 $2NO \leftrightharpoons H_2O_2$ ***(fast equilibrium)***

k2 $H_2O_2 + H_2 \rightarrow N_2O + H_2O$ ***(slow)***

k3 $N2O + H_2 \rightarrow N_2 + H_2O$ ***(fast)***

1.) What is the overall reaction (reaction excluding the intermediates and catalyst)?

2.) Determine the rate law

3.) Identify the intermediates

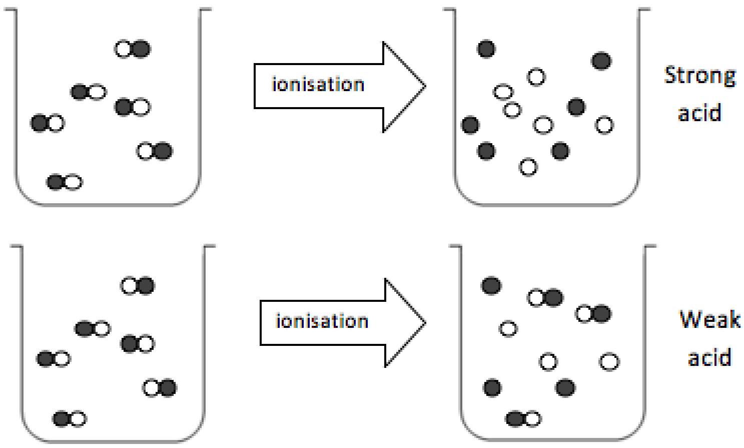
4.) Is there any catalyst present?

5.) Draw potential energy diagram for mechanism

(Malcolm Jennings)

Acids and Bases (Eshbaugh)

What's an Acid? What's a Base? These are the questions you may be asking yourself, if not you probably already know the basics. But do you know there is a difference between a strong and weak acid, one that is more than just how diluted the acid or base is?



Boom, the answer is here. A strong acid, or base, has the tendency to completely or almost completely dissociated into its ions, while a weak acid will just kinda chill and not.

Brønsted-Lowry Acids and Bases: Danish Chemist, Johannes Bronsted, and English Chemist, Thomas Lowry have given us a more general definition of what acids and bases are. According to them, **an acid is a hydrogen donor**, while **a base is a hydrogen receiver.**

Example: adding Hydrochloric Acid (HCl) to water will cause the HCl to donate its proton (H+) to water, making it a Brønsted-Lowry Acid.

Characteristics of Acids and Bases:

Acids: 1 - 6.9999 pH, they taste sour, they increase H^+ in water.

Bases: 7.0001 - 14 pH, bitter, increases OH^-

Properties of Acids and Bases

ACIDS	BASES
taste **sour**	taste **bitter**
do not feel slippery	feel **slippery**
pH < 7	**pH > 7**
release hydrogen (**H^+**) ions in aqueous solution	release hydroxide (**OH^-**) ions in aqueous solution
corrode metals	do not corrode metals
react with metals to produce a compound and **hydrogen gas**	do not react with metals to produce a compound and hydrogen gas
turn litmus **red/pink**	turn litmus **blue**

- **Uses of Acids**

a) Acetic Acid = Vinegar

b) Citric Acid = lemons, limes, & oranges. It is in many sour candies such as lemonhead & sour patch.

c) Ascorbic acid = Vitamin C which your body needs to function.

d) Sulfuric acid is used in the production of fertilizers, steel, paints, and plastics.

e) Car batteries.

- **Uses of Bases.**

a). Bases give soaps, ammonia, and many other cleaning products some of their useful properties.

b) The OH- ions interact strongly with certain substances, such as dirt and grease. c) Chalk and oven cleaner are examples of familiar products that contain bases.

d) Your blood is a basic solution.

Acid – Base Reactions

a) A reaction between an acid and a base is called neutralization. An acid base mixture is not as acidic or basic as the individual starting solutions. **(Manish Khadka)**

Periodic Trends (Frashure) (Eshbaugh)

Let's break it down. I don't know if you guys covered this in honors chem, but not all elements act alike (shocking, i know). There are some trends you need to keep in mind when trying to decide how elements react or interact with one another, or themself.

- **Atomic Radius**
    - The atomic radius is the distance from the atomic nucleus to the outermost stable electron orbital in an atom that is at equilibrium. The atomic radius tends to decrease as one progresses across a period because the effective nuclear charge increases, thereby attracting the orbiting electrons and lessening the radius. The atomic radius usually increases while going down a group due to the addition of a new energy level (shell). However, diagonally, the number of protons has a larger effect than the sizeable radius.
- **Ionization Potential**
    - The ionization potential (or the first ionization energy) is the minimum energy required to remove one electron from each atom in a mole of atoms in the gaseous state. Trend-wise, ionization potentials tend to increase while one progresses across a period because the greater number of protons (higher nuclear charge) attract the orbiting electrons more strongly, thereby increasing the energy required to remove one of the electrons. There will be an increase of ionization energy from left to right of a given period and a decrease from top to bottom.
- **Electron Affinity**
    - The electron affinity of an atom can be described either as the energy gained by an atom when an electron is added to it, or conversely as the energy required to detach an electron from a singly-charged anion. As one progresses from left to right across a period, the electron affinity will increase, due to the larger attraction from the nucleus, and the atom "wanting" the electron more as it reaches maximum stability. Down a group, the electron affinity decreases because of a large increase in the atomic radius, electron-electron repulsion and the shielding effect of inner electrons against the valence electrons of the atom.

- **Electronegativity**
    - Electronegativity is a measure of the ability of an atom or molecule to attract pairs of electrons in the context of a chemical bond. The type of bond formed is largely determined by the difference in electronegativity between the atoms involved, using the Pauling scale. Trend-wise, as one moves from left to right across a period in the periodic table, the electronegativity increases due to the stronger attraction that the atoms obtain as the nuclear charge increases. Moving down a group, the electronegativity decreases due to the longer distance between the nucleus and the valence electron shell, thereby decreasing the attraction, making the atom have less of an attraction for electrons or protons.

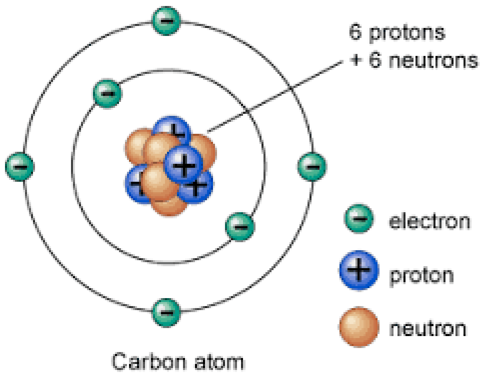
Practice:

1) Which of these has the largest atomic radius?
    a) Aluminum
    b) Fluorine
    c) Calcium
    d) Potassium
    e) Sulfur
2) Which of these have the lowest Electron Affinity?
    a) Polonium
    b) Sulfur
    c) Tellurium
    d) Selenium
3) What's the difference between Electronegativity and Electron Affinity?

4) Why does Fluorine have a higher ionization energy than Iodine?

<u>The Atom</u> <u>(Daimen Pierce)</u>

I hope you understand the basics of what an atom consists of. If you don't then this will save your life. An atom is the most basic and fundamental concept to understand in AP Chemistry.



An Atom consists of three units: **Electrons, Protons, and Neutrons.**

<u>Protons</u>: positively charged and are in the nucleus of the atom (the center).

<u>Neutrons</u>: are also in the nucleus of the atom and are a neutral charge.

<u>Electrons</u>: orbit the nucleus and are negatively charged.

The radius of an atom is as far as the farthest energy level of electrons are, the valence shell of electrons. Electrons orbit a nucleus because they are attracted to the positively charged protons. Opposites attract!

- An element is an atom that is independent from other elements based on its protons. Naturally, an atom will have an equal number of electron to protons. Elements are also classified by how many protons they have by their atomic number. For example hydrogen has one proton and as an atomic number of one.
- The characteristics of an atom are determined by how many protons it has. The number of neutrons and electrons in an element can vary put the number of protons will always remain the same in an atom.
- An isotope is an element that has more or less neutrons than the average (the one given in the periodic table, its Atomic Weight measured in AMU's or atomic mass units). Protons and neutrons both weigh 1 AMU and electrons, while technically having mass, have no significant weight (0 AMU).

- An isoelectronic atom is one that has a full octet (8 electrons in the valence shell). These atoms are now negatively charged because there are more electrons than protons and have properties more closely linked to the Noble Gases (which naturally have a full octet)

- Atoms can donate or receive electrons from other atoms in a chemical reaction in order to obtain a full octet (this is an atom's dream, they want this because it is more stable)

Practice (No isotopes- just elements at their most basic state)

Use this link for an interactive periodic table: [***https://www.ptable.com/***](https://www.ptable.com/)

1. ***What element has 6 protons?_____***
2. ***What element has 17 protons?_____***
3. ***Which elements have a natural full octet?_____***
4. ***What element has 8 electrons?_____***
5. ***What element has 79 electrons?_____***
6. ***If an Oxygen atom has an atomic mass of 17, how many neutrons does it have?_____***
7. ***If an atom has an atomic mass of 7 and has 4 neutrons, what element is it?_____***

(Daimen Pierce)

Equilibrium

What is It?

When you have a system made up of a bunch of molecules, those molecules sometimes combine. That's the idea of a chemical reaction. A chemical reaction sometimes starts at one point and moves to another. Now imagine the reaction finished and you have a pile of new chemicals. Guess what? Some of those chemicals want to go through a reverse chemical reaction and become the original molecules again. This is what equilibrium is- The point where the formed chemicals (Called Products) wants to re-react to form the initial chemicals (called Reactants)

ICE Boxes: A way to determine concentrations, moles or pressures using applied stoichiometry. ****************MUY IMPORTANTE*************

ICE is a simple acronym for the titles of the first column of the table.

- **I** stands for **initial**. This row contains the initial concentrations or pressures of products and reactants.
- **C** stands for the **change** in concentration. This is the concentration change required for the reaction to reach equilibrium. It is the difference between the equilibrium and initial rows. The concentrations in this row are, unlike the other rows, expressed with either an appropriate positive (+) or negative (-) sign and a variable; this is because this row represents an increase or decrease (or no change) in concentration.
- **E** is for the concentration when the reaction is at **equilibrium**. This is the summation of the initial and change rows. Once this row is completed, its contents can be plugged into the equilibrium constant equation to solve for K_c.

Equilibrium equation:

$$aA_{(g)} + bB_{(g)} \rightleftharpoons cC_{(g)} + dD_{(g)}$$

$$K_c = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

What does K represent?

- K provides information on how the rxn will react
- If the **K is less than 1** the reaction will be favored towards the left, thus forming more reactants

If the **K is greater than 1** the reaction will then be favored towards the right, forming more products

Exercise 1 **Writing Equilibrium Expressions**

Write the equilibrium expression for the following reaction:

$$4\ NH_3(g) + 7\ O_2(g) \rightleftharpoons 4\ NO_2(g) + 6\ H_2O(g)$$

$$K = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7}$$

<u>Practice</u>: this is what an ICE Box looks like. Check youtube for how to fill this out more, but essentially the x values are equivalent to their coefficient. **The 'E' in the ICE box will be a combination of 'I' and 'C' rows.**

Reaction: $CO(g) + H_2O(g) \Longleftrightarrow CO_2(g) + H_2(g)$

I	0M	0.0500M	0.100M	0.200M
C	-x	-x	+x	+x
E		0.05-x		

So in this situation, Carbon Monoxide reacts with Water Vapor in order to create Carbon Dioxide and Hydrogen Gas, but the hydrogen gas also reacts with the Carbon Dioxide to produce Carbon Monoxide and Water Vapor

1. ***What are the Reactants in the example above?***

2. ***What are the Products in the example above?***

3. ***If the "E" row for H_2O is 0.05-x, what is it for the rest of the compounds (again, combine "I" and "C" to make "E"?***

Polyatomic Ions

Polyatomic ions are ions that consist of more than one kind of atom. (Ions form when an atom either gains or loses valence electrons in pursuit of filling its valence shell.) The following table lists the most common polyatomic ions, grouped by ionic charge. They'll pop up frequently and annoy you until you simply buckle down and memorize them.

Common Polyatomic Ions

- NH_4^+ ammonium
- OH^- hydroxide
- CN^- cyanide
- SO_4^{2-} sulfate
- O_2^{2-} peroxide
- HSO_4^- hydrogen sulfate
- $C_2H_3O_2^-$ acetate
- SO_3^{2-} sulfite
- ClO_4^- perchlorate
- CO_3^{2-} carbonate
- NO_3^- nitrate
- ClO_3^- chlorate
- HCO_3^- hydrogen carbonate
- NO_2^- nitrite
- ClO_2^- chlorite
- $C_2O_4^{2-}$ oxalate
- PO_4^{3-} phosphate
- ClO^- hypochlorite
- HPO_4^{2-} hydrogen phosphate
- CrO_4^{2-} chromate
- $H_2PO_4^-$ dihydrogen phosphate
- $Cr_2O_7^{2-}$ dichromate
- H_3O^+ hydronium
- PO_3^{3-} phosphite
- MnO_4^- permanganate

Gas Laws

The Ideal Gas Law states that one mole of an ideal gas will occupy a volume of 22.4 liters at STP (Standard Temperature and Pressure, 273K and 1atm).

The Ideal Gas Law also states that the volume (V) occupied by n moles of any gas has a pressure (P) at temperature (T) in Kelvin. The relationship for these variables, P V = n R T, where (R) is known as the gas constant, is called the ideal gas law or equation of state.

- Volume (V) measured in Liters
- Pressure (P) measured in atm
- Temperature (T) measured in Kelvin
- Moles represented by (n)
- Gas Constant R is .08206 L atm mol-1K-1

MAKE SURE ALL VARIABLES ARE CONVERTED TO CORRECT UNITS
REFER TO PREVIOUS PAGES FOR UNIT CONVERSIONS

In addition, the relationships between Pressure, Volume, and Temperature are found in 3 equations derived from the Ideal Gas Law.

AN INVERSE RELATIONSHIP MEANS THAT WHEN ONE VARIABLE INCREASES, THE OTHER DECREASES
A DIRECT RELATIONSHIP MEANS THAT VARIABLES INCREASE AND DECREASE TOGETHER

- Boyle's Law states that the pressure and volume of a gas have an **inverse** relationship, where P1V1=P2V2

- Charles's Law states that the volume of a gas is **directly** proportional to the temperature, where V1/T1 = V2/T2
- *Gay Lussac's Law* states that the pressure of a gas is **inversely** related to the temperature, where P1/T1 = P2/T2

THESE LAWS ARE ONLY USED WHEN THE MISSING VARIABLE IS CONSTANT

Practice: (SARAH)

BOYLE'S LAW

Sulfur dioxide (SO2), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants. Consider a 1.53-L sample of gaseous SO2 at a pressure of 5.6 × 10^3 Pa. If the pressure is changed to 1.5 × 10^4 Pa at a constant temperature, what will be the new volume of the gas ?

CHARLES' LAW

A sample of gas at 15ºC and 1 atm has a volume of 2.58 L. What volume will this gas occupy at 38ºC and 1 atm ? (MUST USE KELVIN)

GAY-LUSSAC'S LAW

A 30.0 L sample of nitrogen inside a rigid, metal container at 20.0 °C is placed inside an oven whose temperature is 50.0 °C. The pressure inside the container at 20.0 °C was at 3.00 atm. What is the pressure of the nitrogen after its temperature is increased to 50.0 °C?

IDEAL GAS LAW

A sample of hydrogen gas (H2) has a volume of 8.56 L at a temperature of 0ºC and a pressure of 1.5 atm. Calculate the moles of H2 molecules present in this gas sample.

Intermolecular Forces (Jake Wellek)

Intermolecular Forces (IMF)- interactions between two molecules, rather than within the molecule. There are four different IMF's.

- London Dispersion Force (LDF): The forces that exist among **noble gas** atoms and **nonpolar molecules**. As the electrons move about the nucleus, a momentary nonsymmetrical electron distribution can produce a temporary dipole arrangement of charge

- Dipole-Dipole attraction (dip-dip): Molecules with polar bonds often behave in an electric field as if they had a center of positive charge and a center of negative charge. Molecules with dipole moments (**polar covalent bonds**) can attract each other electrostatically by lining up so that the positive and negative ends are close to each other.

- Hydrogen Bonding: Unusually strong dipole-dipole attractions that occur among molecules in which **hydrogen** is bonded to a highly electronegative atom (**Nitrogen, Oxygen, or Fluorine**).

- Ion Dipole: an attractive force that results from the electrostatic attraction between an **ion** and a neutral molecule that has a dipole (**polar covalent bond**).

Type of IMFA	Involves	Occurs Between...	Strength of Attraction	Effect on Boiling/Melting Points	Effect on Freezing Points
London/Dispersion	Temporary Dipoles	Nonpolar molecules	Low	Low elevation (↑BP)	(↓↓↓↓FP)
Dipole-Dipole	Permanent dipoles	Polar molecules	Medium	Medium elevation (↑↑BP)	(↓↓↓FP)
Hydrogen Bonds	Permanent dipoles between H & F,O, N	Polar molecules	Medium-High	Medium-High elevation (↑↑↑BP)	(↓↓FP)
Ion-Dipole	Full ion and dipole	Polar molecules	High	High elevation (↑↑↑↑BP)	(↓FP)

Practice:

What types of intermolecular forces are acting in the following phases of matter?

1. ***N_2 (g)?***
   a. ***Ion-Ion***
   b. ***Ion-Dipole***
   c. ***Dipole-Dipole***
   d. ***London Dispersion***

2. ***NaCl (s)?***
   a. ***Ion-Ion***
   b. ***Ion-Dipole***
   c. ***Dipole-Dipole***
   d. ***London Dispersion***

3. ***H_2O (l)?***
   a. ***Ion-Ion***
   b. ***Ion-Dipole***
   c. ***Dipole-Dipole***
   d. ***Dispersion***

4. NH_3 (s)?

a. ***Ion-Ion***
b. ***Ion-Dipole***
c. ***Dipole-Dipole***
d. ***Dispersion***

MULTIPLE CHOICE. Choose the one alternative that best completes the statement or answers the question.

1. ***The principal source of the difference in the normal boiling points of ICl (97eC; molecular mass 162 amu) and Br_2 (59eC; molecular mass 160 amu) is __________.***
   a. ***ICl has greater strength of hydrogen bonding than Br_2***
   b. ***ICl has stronger dipole-dipole interactions than Br_2***
   c. ***ICl has stronger dispersion forces than Br_2***
   d. ***The I-Cl bond in ICl is stronger than the Br-Br bond in Br_2***
   e. ***ICl has a greater molecular mass than Br_2***

2. ***Which one of the following derivations of ethane has the highest boiling point?***
   a. C_2I_6
   b. C_2Br_6
   c. C_2F_6
   d. C_2Cl_6
   e. C_2H_6

3. ***What is the predominant intermolecular force in CBr4?***
   a. ***Hydrogen-bonding***
   b. ***Ion-dipole attraction***
   c. ***Ionic bonding***
   d. ***Dipole-dipole attraction***
   e. ***London-dispersion forces***

Titration and Neutralization

Titration: adding a base or acid to a solution of unknown concentration to find said concentration. The unknown solution will have an **indicator** in it that will change the solution's color at the endpoint (when the titration is done). During titration, you always titrate stuff with the opposite. You can't titrate an acid with an acid, it would just increase the acidity and you'd get an big, fat L.

In order to understand how this jazz works, you need this equation.

$$M = moles/liters$$

The **M** is **molarity**, a measurement of concentration. There are other measurements, like molality (grams/liters), but they all suck because they don't use moles, so molarity is all you really need. To refer to concentration, we put stuff in brackets, like **$[Na^+]$**.

Equivalence Point: when moles of acid equals moles of base, NOT NECESSARILY THE END POINT. Sometimes the end point and equivalence point are the same, but don't assume that.

The equation to find this point is…

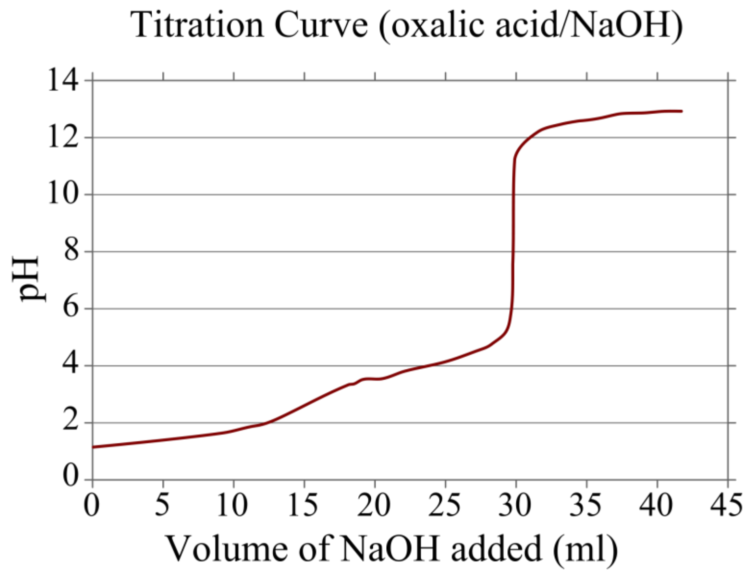
$$M_1V_1 = M_2V_2$$

Molarity times volume is just moles! Magical. Don't forget this equation 'cause it's really half the acid base and titration unit. You're screwed if you do.

They'll probably ask you to find pH and or $[H_3O^+]$, so have ANOTHER equation.

$pH = -\log([H_3O^+])$ $1 \times 10^{14} = [OH^-][H_3O^+]$ $pOH = -\log([OH^-])$ $[H_3O^+] = 10^{-pH}$

Now, you'll also often see a graph when there's a titration problem. Looks kinda like this.



The equivalence point is where the graph has a zero slope (vertical line) (DERIVATIVE!!!!), or looks like it's a straight line. These graphs are always of pH and volume so don't get anything mixed up.

<u>Neutralization</u>: when a equal moles of a strong acid and base react, they neutralize each other to form H_2O and a salt.

<u>Salt</u>: any ionic compound.

TITRATION DOESN'T ALWAYS have to be a NEUTRALIZATION REACTION!!!!!!!!

If you have a weak acid being titrated with a strong base, the pH at the equivalence point will be basic. If you have a weak base being titrated with a strong base, the pH will be acidic.

<u>Buffers</u>: a thingumy that maintains the pH of a solution. It's made of a Brønsted-Lowry acid and its conjugate base. When acid, like HCl, is added, the base in the buffer neutralizes it. When base, like NaOH, is added, the acid neutralizes it. **However!!!** A buffer cannot be made of a strong acid or a strong base!!!!!!! Tao **will** quiz you on what is and isn't a buffer.

You'll have units on this stuff in AP Chem, so don't worry too much about it. Just remember the equations and definitions for now, so it doesn't matter if you don't totally understand it now. Tao or Jo will explain further, and if you don't understand them, use the packets they hand out and the book. And if you don't understand them then, bribe Tao with fishing hooks and Fortnite v-bucks. Jo will accept food for payment as well.

<u>Practice</u>:

1. ***How many moles HCl is in a 10 mL solution of of .6 M HCl?***

2. ***What is the molarity of a 15 mL solution with 0.01 moles of KOH?***

3. ***What is the volume needed to make 4 M HNO_3 with 2 moles of acid?***

4. ***You're titrating HCl with 0.1 M NaOH. What is the concentration of HCl if the equivalence point is at 10 mL?***

5. ***How much 0.25 M NaOH is necessary to neutralize 75 mL of .7987 M HI?***

Thermochemistry I: ΔH and Heating Curves (Reyes)

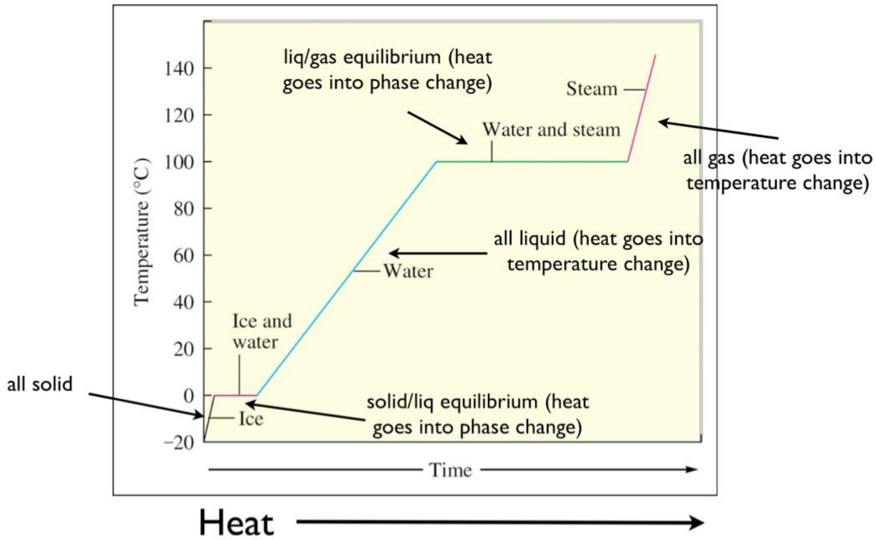
- Enthalpy (H)- flow of energy (heat exchange) at constant pressure when two systems are in contact
- ΔH is a state function
  - ΔH = q at constant pressure (atmospheric pressure)
- ΔH represents the change in enthalpy of a system in a reaction
  - When ΔH is on the right side of the chemical equation, the reaction is exo
  - When ΔH is on the left side of the chemical equation, the reaction is endo
- Breaking bonds is an endothermic process
  - ΔH = +
- Forming bonds is an exothermic process
  - ΔH = -
- Enthalpy can be calculated from several sources such as:
  - Stoichiometry
  - Calorimetry
  - From tables of standard values
  - Hess's Law
  - Bond energies
- $q = m\ Cp\ \Delta T$ *mcat*
  - q = heat energy
  - m = mass
  - Cp and s = specific heat capacity
    - *Cp and s are the same, but s is the main variable used*
      - $q = m\ Cp\ \Delta T$ and $q = m\ s\ \Delta T$ are the same
  - $\Delta T = T_{final} - T_{initial}$
- Cp H_2O = 4.18 J/ g°C
- 1 calorie = 4.18 Joules
- 1 Calorie/kilocalorie/food calorie = 1000 calories

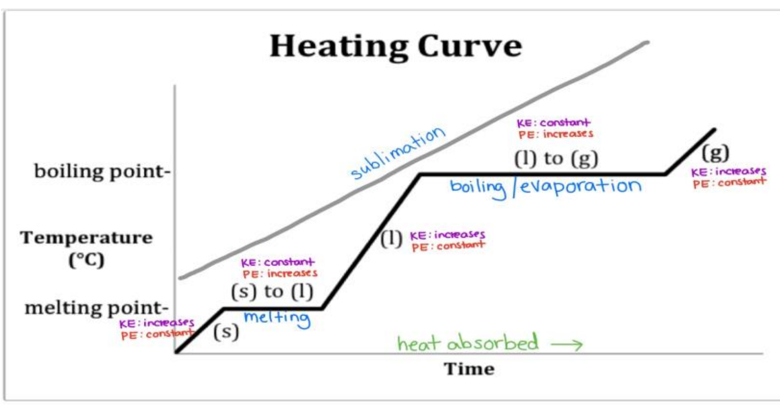
Heating Curves:

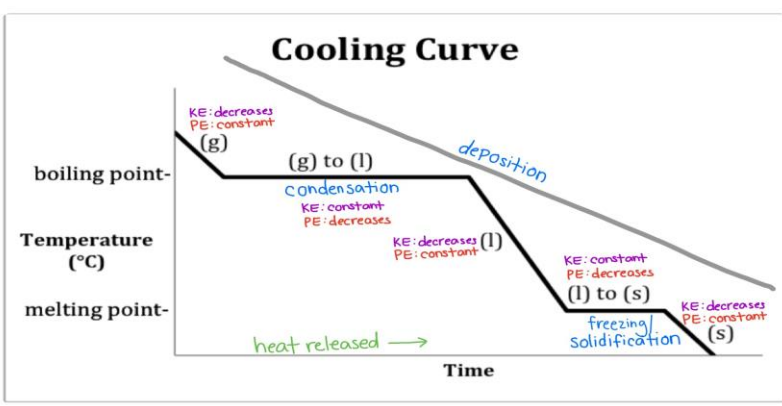
The energy coming in results in higher potential energy not higher kinetic energy. Breaking up the IMF between the molecules leads to a high potential energy.

This can be easily seen in a heating curve that plots the temperature of a system as a function of the heat flow into the system. Initially the system is a solid, then it has a melting transition, then it is a liquid, then has a vaporization transition, and then it is a gas.

The first diagram below shows the heating curve for water. The temperature of the system is plotted as function of time. Heat is flowing at a constant rate, thus time can be interpreted as heat.


liq/gas equilibrium (heat goes into phase change)
Steam
Water and steam
all gas (heat goes into temperature change)
all liquid (heat goes into temperature change)
Water
Ice and water
all solid
solid/liq equilibrium (heat goes into phase change)
Ice
Temperature (°C)
140
120
100
80
60
40
20
0
−20
Time
Heat


Heating Curve
boiling point-
Temperature (°C)
melting point-
sublimation
KE: constant
PE: increases
(l) to (g)
boiling / evaporation
(g)
KE: increases
PE: constant
(l)
KE: increases
PE: constant
KE: constant
PE: increases
(s) to (l)
melting
KE: increases
PE: constant
(s)
heat absorbed →
Time


Cooling Curve
KE: decreases
PE: constant
(g)
(g) to (l)
boiling point-
condensation
KE: constant
PE: decreases
deposition
Temperature (°C)
KE: decreases
PE: constant
(l)
KE: constant
PE: decreases
(l) to (s)
melting point-
freezing/ solidification
KE: decreases
PE: constant
(s)
heat released →
Time

Problems:

1. A 20.8 g piece of iron loses 1086.5 joules of heat energy, and its temperature changes from 100 celsius to 25 celsius. Calculate the specific heat capacity of iron.

2. To what temperature will a 50.0 g piece of glass raise if it absorbs 5555 joules of heat and its specific heat capacity is 0.500 J/g°C? The initial temperature of the glass is 20.0°C.

3. What is the specific heat capacity (J/g°C) of silver metal if 55.00 g of the metal absorbs 47.3 calories of heat and temperature rises 15.0°C?

4. 100.0 mL of 37.0°C water is cooled until its temperature is 4°C. If the specific heat of water is 4.18 J/g°C, calculate the amount of heat energy needed to cause this change in temperature. Density of water is 1 g/mL.

5. When 1 mole of methane (CH_4) is burned at constant pressure, 890 kJ/mol of energy is released as heat. Calculate ΔH for a process in which a 5.8 g sample of methane is burned at constant pressure.

6. In a coffee cup calorimeter, 100.0 mL of 1.0 M NaOH and 100.0 mL of 1.0 M HCl are mixed. Both solutions were originally at 24.6°C. After the reaction, the final temperature is 31.3°C. Assuming that all solutions have a density of 1.0 g/cm^3 and a specific heat capacity of 4.18 J/g°C, calculate the enthalpy change for the neutralization of HCl by NaOh. Assume that no heat is lost to the surroundings or the calorimeter.

Hess's law states that the total enthalpy change during a chemical reaction is the same whether the reaction is made in one step or in several steps.

*Hess's law also appears in other units, therefore it is important to know it and how to solve it

Hess's Law Practice: Calculate ΔH

1. $Sn\ (s) + 2\ Cl_2\ (g) \rightarrow SnCl_4\ (l)$ $\Delta H =$

   $Sn + Cl_2 \rightarrow SnCl_2$ $\Delta H = -325\ kJ$

   $SnCl_2 + Cl_2 \rightarrow SnCl_4$ $\Delta H = -186\ kJ$

2. $4Al = 3MnO_2 \rightarrow 2Al_2O_3 + 3\ Mn$ $\Delta H =$

   $4Al + 3O_2 \rightarrow 2Al_2O_3$ $\Delta H = -3352\ kJ$

   $Mn + O_2 \rightarrow Mn\ O_2$ $\Delta H = -521\ kJ$

3. $5Z + V \rightarrow 2X + 3Y$ $\Delta H =$

   $X + Y \rightarrow 2W$ $\Delta H = 100\ kJ$

   $2Z + V \rightarrow W$ $\Delta H = 400\ kJ$

   $V + 2Y + X \rightarrow Z$ $\Delta H = -75\ kJ$

Thermochemistry II: Spontaneity & Free Energy (Zhen Zhou & Caitlin Dan)

-Need to know (better read other stuff first because you won't understand anything)

- **Spontaneity**
- A process is said to be spontaneous if it occur without outside intervention can be any speed
    - Favors the formation of products under a certain set of conditions
    - The opposite which does not favor the formation of products is said to be nonspontaneous
- Entropy (ΔS) the driving force for a spontaneous process is an increase in entropy of the universe. Entropy can be view as a measure of randomness or disorder.
- **Gibb's Free Energy**
- Useful in dealing with the temperature dependence of spontaneity
- Symbolized by *G*
- $G = H - TS$ or $\Delta G = \Delta H - T\Delta S$
    - Good to memorize using something like "**G**oing **H**ome **T**o **S**leep"

- If $\mathbf{\Delta G_{sys} \leq 0}$, the process is *spontaneous*
- If $\mathbf{\Delta G_{sys} = 0}$, the system is at *equilibrium*
- If $\mathbf{\Delta G_{sys} > 0}$, the process is not *spontaneous*

- If $\mathbf{\Delta G_{sys} < 0}$, the process is *exothermic*
- If $\mathbf{\Delta G_{sys} > 0}$, the process is *endothermic*

- *ΔS positive*	- *ΔH negative*	Spontaneous at all temp
- *ΔS positive*	- *ΔH positive*	Spontaneous at high temp
- *ΔS negative*	- *ΔH negative*	Spontaneous at low temp
- *ΔS negative*	- *ΔH positive*	Not spontaneous at any temp

ENTROPY

Entropy is basically just another word for **disorder.** Use this to answer the questions below.

1. Melting ice is an example of an (increase/decrease) in entropy

2. $H_2O_{(g)} \rightarrow H_2O_{(l)}$ This is has an (increase/decrease) in entropy

3. Which has higher positional entropy?
   a) Gas H_2O or Water H_2O
   b) 100 atm gas or 1 atm gas

ELECTRO CHEM!

Alright, what's up beta males, we gotta introduce you to basic electrochemistry. There are only two big things you really need to know. Those being what parts are in each cell and the types of cells. You also need to know that voltage always moves from anode to cathode no matter what!

Parts

- **SALT bridge**
  - These are not always made of NaCl, but rather any ionic material.
  - These join the two halves of the electrolytic cell, while still keeping the mixtures separate.
  - This also completes the circuit, without this. No voltage could be made or would be stored
- **Anode**
  - This is where the oxidation reaction takes place, meaning this is the piece that has electrons stripped
  - This is made out of a solid metal with no charge(NOT A CATION)

    - During both Galvanic and electrolytic cells, the metal slowly diminishes in mass

- **Cathode**
    - This is where the reduction reaction takes place, meaning this is the piece that gains electrons
    - This is also made out of a solid metal
    - During both Galvanic and electrolytic cells, the metal slowly increases in mass
- **Ion Solutions**
    - These are ions that are found in solutions where the anode and cathode lay, usually where anode and cathode lie in their respective metal ion solution.
    - These drive the oxidation and reduction solutions that occur at the anode and cathode!

Galvanic (Voltaic)

- These cells create voltage from the flow of electrons they have
- These are spontaneous reactions, meaning your delta G value is negative.
- These have a positive voltage and positive E value
- Electrolytic
- These puppers are cells that require voltage to force a reaction. Usually like a battery that is in between the wire connecting the anode and cathode
- These cells are NOT spontaneous, meaning that their delta G value is positive.
- These have a negative voltage and negative E value

PRO TIPS

- For a Galvanic cell, you want the most positive E value, meaning that you will have the flip one of the reduction equations that is given to you
- For a Electrolytic cell, you want the most negative E value, meaning that you will have the flip one of the reduction equations that is given to you.
- The reaction that has electrons as a product after flipping is your anonde
- The reaction that has electrons as a reactant after flipping is your cathode

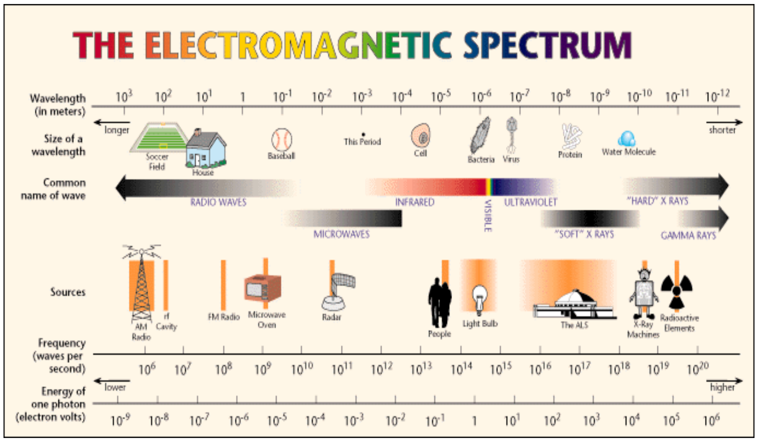
SPECTROSCOPY

What is it?

Spectroscopy is the study of the interaction of radiant energy and matter.

Background Knowledge:

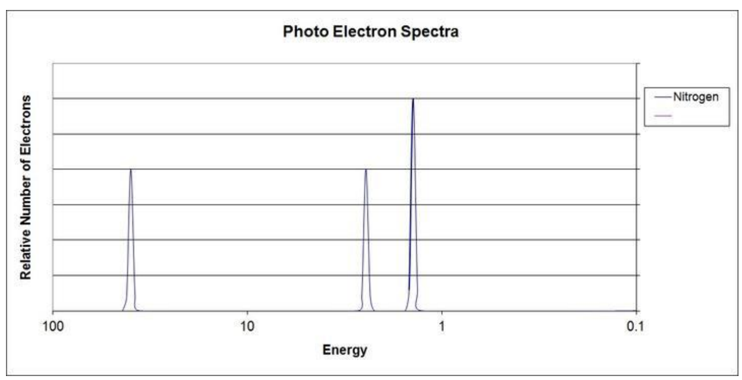
Before beginning to learn spectroscopy, you should review some electron configuration, basic algebra, and atomic theory.



Looking at the electromagnetic spectrum, we can identify different types of wavelengths. <u>The wavelength at which energy is absorbed gives us information about the particle.</u>

What is Photoelectron Spectroscopy?

When x-ray radiation is shone at a substance, it has so much energy that it will remove electrons. The particular wavelength of x-ray used to remove each electron within the substance gives scientists an idea of the number of electrons in shells and how tightly they are held to the nucleus and thus, the identity of an element. This is known as <u>photoelectron spectroscopy. We use it to determine the electronic structure of atoms and molecules.</u>



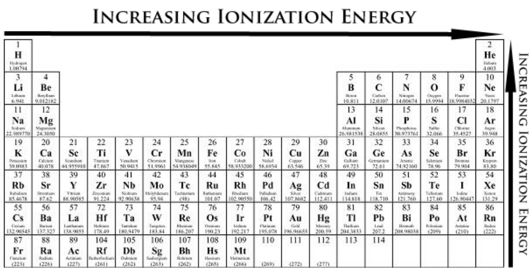
Photoelectron Spectra

This diagram is used to identify elements based on their relative electron count vs. electron binding energy.

Ionization Energy

The ionization energy is how much energy is required to remove an electron from a neutral atom in the gaseous phase, and typically has units of kilojoules or electron volts per mole. It usually follows a periodic trend shown below. Meaning, for example.

Elements such as fluorine hold great amounts of ionization energy while elements such as cesium do not.



$$\textbf{Frequency} = \frac{\textbf{Velocity}}{\textbf{Wavelength}}$$

$$f = \frac{v}{\lambda}$$

Unit of Frequency: Hz

Basics:

Frequency: The number of wavelengths within a given period

Wavelength: The length between two crests.

$$E = \frac{hc}{\lambda}$$

$E = energy$
$h = Planck's\ constant$
$c = speed\ of\ light$
$\lambda = wavelength$

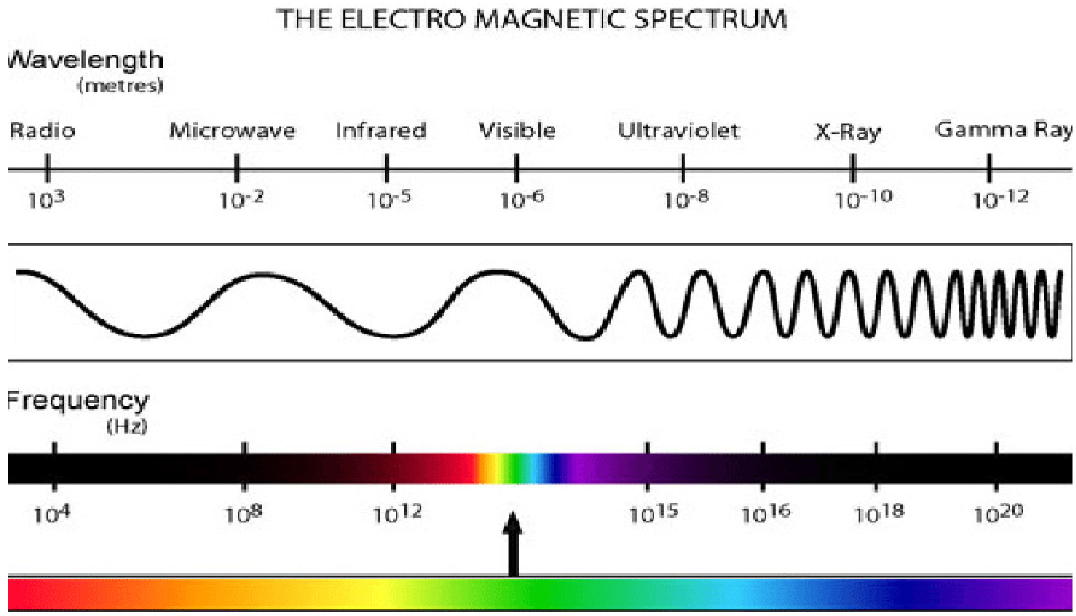
There are 2 main types of photoelectron spectroscopy, depending on the energy of the radiation used to eject electrons: Ultraviolet and X ray photoelectron spectroscopy.

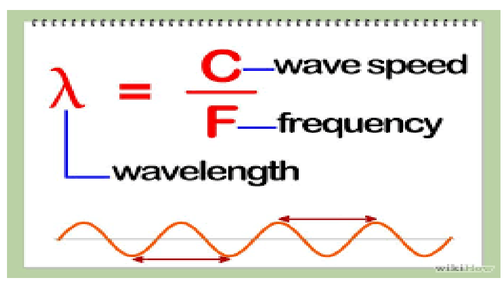
- <u>Ultraviolet:</u>
    - Illuminating a sample with ultraviolet (UV) light will typically ionize the material by ejecting valence electrons.
    - Due to shielding by the core electrons, the valence electrons feel a reduced attraction to the nucleus. Therefore, valence electrons require less energy to remove compared to core electrons.
- <u>X Ray</u>
    - Since x rays have a higher frequency than UV rays, they eject core electrons. Core electrons are inner-shell electrons that are closer to the nucleus, and thus require more energy to remove compared to valence electrons.
    - Once electrons are ejected from the sample, a detector is able to calculate the kinetic energies of the electrons, as well as the relative number of electrons with that kinetic energy. We can use this information to calculate the minimum energy required to remove electrons from different subshells within an atom.
    - This is called the *binding energy* of the electron, and the binding energies depend on the chemical structure and elemental composition of a sample.

Overview

- Photoelectron spectroscopy is a useful analytical tool used by chemists to determine the electronic structure of atoms and molecules.
- Photoelectron spectrometers ionize samples by bombarding them with high-energy radiation, such as UV or x-rays, and detecting the number and kinetic energy of ejected electrons.
- The frequency and energy of incident photons can be used to calculate the binding energy of the ejected electron using the following **equation : $BE= h\nu - KE_{electron}$**
- The PES spectrum is a graph of relative electron count vs. electron binding energy.
- The peaks in PES spectra correspond to the electrons in different subshells of the atom. The peaks with the lowest binding energy correspond to the valence electrons, while the peaks at the highest binding energy correspond to the inner-shell or core electrons.
- Problems
    1. A laser emits light of frequency 4.74 x 1014 sec-1. What is the wavelength of the light in nm?
    2. A certain electromagnetic wave has a wavelength of 625 nm.
        a. What is the frequency of the wave?
        b. What is the energy of the wave?
    3. A blue light has a frequency of about 7.5 x 1014 Hz.
        a. Calculate the wavelength
        b. Calculate energy in joules.

Jazmin Castelo 65 and the little bit on page 66





Wave Characteristics

- Wavelength ( ) - is the distance between two peaks or troughs in a wave
- Frequency ( ) - is the number of waves per second that pass a given point in space
- Speed (c) - all types of electromagnetic radiation travel at the speed of light

- 2.99 x 10^8 m/s
- c=^v

Velocity = (wavelength)(frequency)

What is the frequency of light that has a wavelength of 550 nm

Frequency of Electromagnetic Radiation

The brilliant red colors seen in fireworks are due to the emission of light with wavelengths around 650 mn when strontium salts such as Sr(NO3)2 and SrCO3 are heated. (This can be easily demonstrated in the lab by dissolving one of these salts in methanol that contains a little water and igniting the mixture in an evaporating dish) Calculate the frequency of red light of wavelength 6.50x10^2 nm

Bonding

The forces that hold atoms together in compounds are called **chemical bonds**.

Ionic Bonds - bonds formed by the attraction between oppositely charged ions. Occurs when there is a large difference between electronegativities. (Metal + Nonmetal)
Ionic compounds are formed when a metal reacts with a nonmetal. The energy of interaction between a pair of ions can be calculated using **Coulomb's Law**, which is,

$$E=(2.31 \times 10^{-19}\ J \cdot\ nm)\ (Q_1Q_2/r)$$

E = Energy (in joules)
R = Distance between the ion centers in nanometers
Q_1 and Q_2 = numerical ion charges

A negative value means there is attraction between the pair of ions.
A positive value means there is repulsion and no bond is formed

Lattice Energy - change in energy that it takes to completely separate a mole of a solid ionic compound into its gaseous ions.
Lattice energy increases with the increasing charge on ions and the decreasing size of ions

Covalent Bonding - bonds form from electron sharing by the nuclei.
Nonpolar Covalent Bond - when a pair of electrons is equally shared between two nonmetal atoms.
Polar Covalent Bond - when a pair of electrons is unequally shared between two nonmetal atoms. The greater the difference in electronegativity, the more polar the bond.

Lewis Structure - Show how valence electrons are arranged among the atoms in the molecule.
The most important requirement for the formation of a stable compound is that the atoms achieve noble gas electron configurations.
Duet Rule - Hydrogens form stable molecules when it shares two electrons. Hydrogen and Helium can only have 2 electrons in their valence shells
Octet Rule - When an atom is surrounded by 8 electrons, which fills the valence shell, giving it a same electron configuration as noble gases.

Steps for Writing Lewis Structures

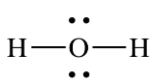
1.) Calculate sum of valence electrons in the atoms
2.) Use a pair of electrons to form a bond between each pair of bound atoms
3.) Arrange the remaining electrons to satisfy the duet rule for hydrogen and octet rule for second-row elements

Example:
H_2O
Step 1) H = 1 electron
H = 1 electron
O = 6 electrons
1+1+6 = 8 electrons total

Step 2) Draw the bonds

H — O — H

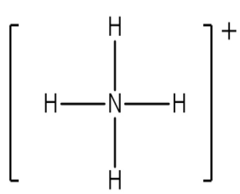
Step 3) Distribute remaining electrons (Make sure to keep track of your electrons)
Since two bonds were used, only 4 electrons need to be distributed because each bond uses 2 electrons



The electrons are added to the Oxygen atom to make it have a noble gas configuration.

Double bonds and triple bonds can be used in a Lewis structure, with the double bond containing 4 electrons and the triple bond containing 6 electrons.

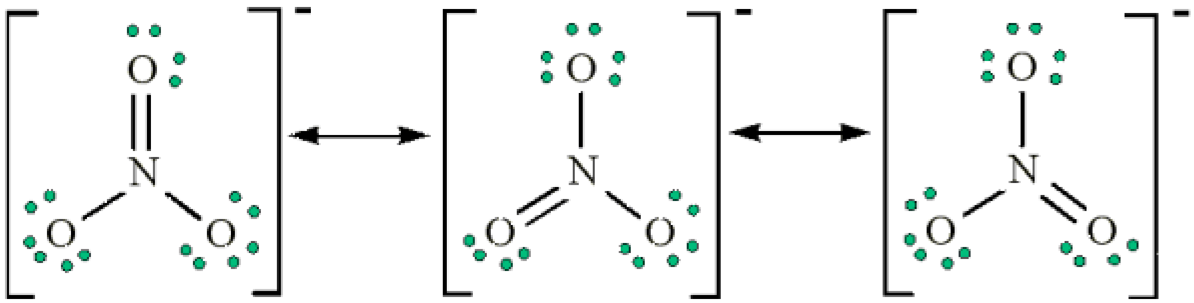
Brackets are needed to surround the Lewis structure if the molecule has a charge.
Example: NH_4^+



Exceptions to the octet rule include boron, which can have less than 8 electrons around it, meaning it does not have a complete octet.
Sulfur is also an exception as it can exceed the octet rule, making it have more than 8 electrons around it.

Resonance Structures

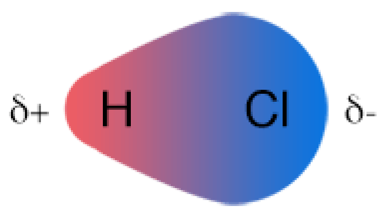
Resonance structures occur when there is more than one valid Lewis structure can be written for a particular molecule.
Example: NO_3^-

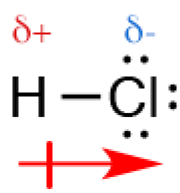


Dipoles - When two equal, but opposite charges are separated by a distance, a dipole forms.

Dipole moment - When a molecule has a center of positive charge and a center of negative charge, it would have a dipole moment. Dipole character of a molecule is represented by an arrow pointing to the negative charge center with the tail of the arrow indicating the positive charge center.
The atom in a molecule with the higher electronegativity will be partially negative, and the one with lower electronegativity will be partially positive.

Example:



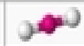
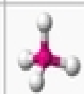
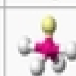
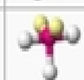
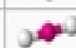
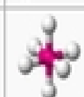
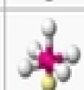
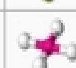
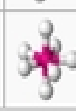


Chlorine has a higher electronegativity than Hydrogen, therefore it is partially negative and Hydrogen is partially positive making the arrow point to Chlorine.

Molecular Shapes

VSEPR - It is useful in predicting geometries of molecules formed from nonmetals. The bonding and nonbonding pairs of electrons around an atom will be positioned as far apart as possible.

Below is a chart of all the molecular geometries, including the bond angles and examples.

Bonding electron pairs	Lone pairs	Electron domains (Steric #)	Shape	Ideal bond angle (example's bond angle)	Example	Image
2	0	2	linear	180°	CO_2	
3	0	3	trigonal planar	120°	BF_3	
2	1	3	bent	120° (119°)	SO_2	
4	0	4	tetrahedral	109.5°	CH_4	
3	1	4	trigonal pyramidal	107°	NH_3	
2	2	4	angular	109.5° (104.5°)	H_2O	
5	0	5	trigonal bipyramidal	90°, 120°, 180°	PCl_5	
4	1	5	seesaw	180°, 120°, 90° (173.1°, 101.6°)	SF_4	
3	2	5	T-shaped	90°, 180° (87.5°, < 180°)	ClF_3	
2	3	5	linear	180°	XeF_2	
6	0	6	octahedral	90°, 180°	SF_6	
5	1	6	square pyramidal	90° (84.8°), 180°	BrF_5	
4	2	6	square planar	90°, 180°	XeF_4	
7	0	7	pentagonal bipyramidal	90°, 72°, 180°	IF_7	

Double and triple bonds have larger electron domains than single bonds, meaning they exert a greater repulsive force than single bonds, making their bond angles greater.

Practice:

Draw the Lewis dot and VSEPR structures for the following:

Molecule	Lewis structure	VSEPR
H_20		
CO_2		
CCl_4		
PCl_5		
NH_3		