

AP Chemistry Exam Review

(<https://www.youtube.com/watch?v=30LgKHJokmw>) Video link to the explanations



KNOW THESE “BASICS” FOR EACH TOPIC (OR MULTIPLE TOPICS)

RTFQ then ATFQ !!!

Reading = “boxing in” important information, underlining important information, making brief notes (or notations) around the question, graph, and illustrations.

EQUILIBRIUM, K , K_c , K_p , K_{sp} (2:30)

- At Equilibrium, the rate of the products being formed is equal to the rate of reactants being formed. However, the INITIAL rate (only reactants with NO products) is the fastest rate of the reaction... The rate will slow until equilibrium is established.

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

$$K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

Equilibrium Constants

K_c (molar concentrations)

K_p (gas pressures)

K_a (weak acid)

K_b (weak base)

K_w (water)

- When a system at Equilibrium is subjected to a stress, the system will shift so as to relieve the stress. Le Chatelier !!!

What is NOT given: (you may want to put in calculator)

K_{sp} (weak salt, solubility product (sp))

$$K_{sp} = [B]^b [C]^c, \text{ where } a A \leftrightarrow b B + c C$$

- Shift Away from the Addition, Shift Toward the Removal.
 - Heat: Endothermic (Heat acts as a Reactant), Exothermic (Heat acts as a Product)
 - When the temperature changes, Equilibrium constant, K , CHANGES!!!
 - Concentration
 - Changes in concentration will not change Equilibrium constant, but will cause a shift.
 - Change in Volume of the container (Pressure)
 - Changes in pressures will not change the Equilibrium constant UNLESS there are the same number of gas particles on the reactant and product sides – this would VERY LIKELY change the temperature as a result of the volume change.
- Adding solids, liquids, or inert gases (any gas that will not react with the particles in the reaction system) WILL NOT impose a stress on a system at equilibrium.

Interesting Equilibrium Fact: When a system at equilibrium has a reactant or product added to the reaction (stress imposed), its concentration (or partial pressure) will be greater than when it was at equilibrium – once equilibrium is re-established, yet the Equilibrium Constant, K , (ratio of products over reactants) does not change (remains the same) as long as the temperature does not change.

Also, A true equilibrium reaction does not go to completion. Completion indicates that all the reactants have formed products.

EQUILIBRIUM (CON'T)

Q vs K

“Q” is the Equilibrium Quotient – this quotient indicates the concentrations or pressures at “some point” in the reaction.

$Q > K$, Equilibrium Quotient has a greater value than the Equilibrium Constant. This indicates that there are too many PRODUCTS and too few REACTANTS. The reaction needs to shift to form more reactants and reduce the amount of products.

$$K = \frac{[product]}{[reactant]}$$

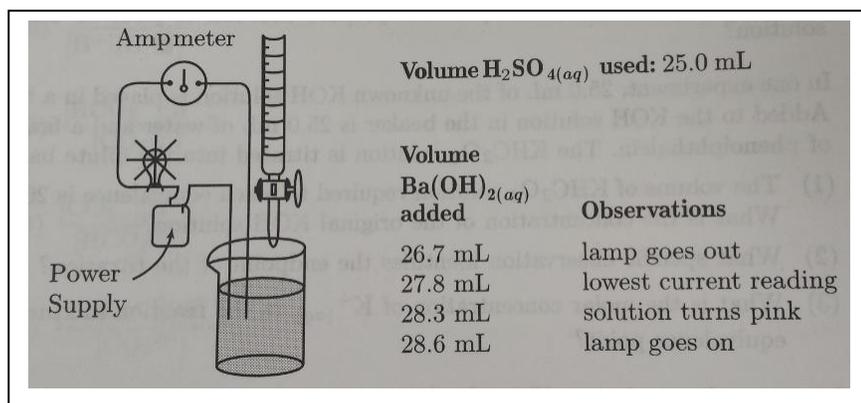
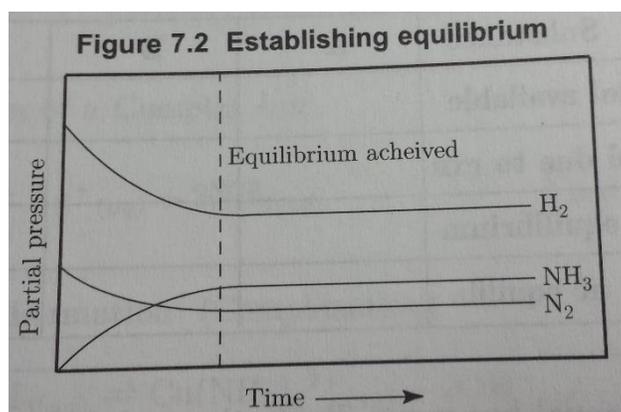
Q > K
↓
↑
Shift to form more reactants

$Q < K$, Equilibrium Quotient has a lesser value than the Equilibrium Constant. This indicates that there are too many REACTANTS and too few PRODUCTS. The reaction needs to shift to form more products and reduce the amount of reactants.

$$K = \frac{[product]}{[reactant]}$$

Q < K
↑
↓
Shift to form more products

$Q = K$, Equilibrium Quotient is equal to the Equilibrium Constant. This indicates that the rate of forming products is equal to the rate of formation of the reactants.



Conductivity could be used in a titration. The conditions for this would be reacting an acid with a base where water and a **precipitate** are formed. Things to remember, only ions can conduct electricity (within a solution), precipitates do not conduct electricity. Pure Water does not conduct electricity. Therefore, at the equivalence point, you have equal moles of acid and base – which have formed PURE WATER and a Precipitate! pH is not the issue for a question like this!!! See graphic on right for a great example. Try to explain the observations.

ACIDS / BASES: (12:42)

Be able to determine $[H^{+1}]$ / pH **without** the use of a calculator:

Examples:

- pH = 4.00 has a $[H^{+1}]$ value = 1.00×10^{-4}
- pH = 5.50 has a $[H^{+1}]$ value between 1.00×10^{-5} & 1.00×10^{-6}
- $[H^{+1}] = 1.00 \times 10^{-11}$ has a pH value = 11.00
- $[H^{+1}] = 4.33 \times 10^{-8}$ has a pH value between 7.00 & 8.00

Neutralization – Deals with removing H^{+1} or OH^{-1} ions from a solution.

Equivalence Point: Similar to Neutralization. The Equivalence Point is a certain pH value when the acid (or base) that is being titrated is neutralized (moles of acid equal moles of base). On the graph to the right, there is no point given to represent the Equivalence point, BUT it would be at 40 mL and a pH of 8.00 (for this titration).

- **Starting with a Weak Acid**, the pH will be ABOVE 7.00 at the equivalence point. (the graph on the right represents this idea).
- **Starting with a Weak Base**, the pH will be BELOW 7.00 at the equivalence point.

Half-Equivalence Point: The VOLUME where the weak acid (or weak base) is half-way to the equivalence point. This is an IMPORTANT point on the titration curve (graph) – the pH will equal the pK_a (or pOH will equal the pK_b). On the graph to the right, Point “S”, represents the half-equivalence point.

Also! At the Half-Equivalence Point, that is when the $[HA] = [A^{-1}]$ or $[B] = [HB^{+1}]$

$M_a V_a = M_b V_b$ is a FAST calculation you can use to determine certain values at the Equivalence point. Also do NOT waste time converting volume to liters, use the volumes and molarities that are given.

KNOW the six strong acids!

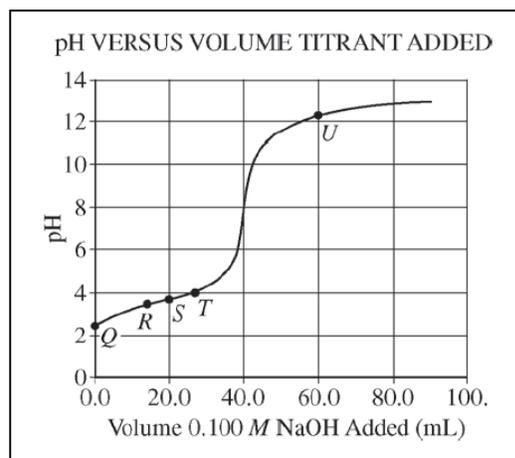
- HCl, HBr, HI, HNO_3 , $HClO_3$ (some argue $HClO_4$), and H_2SO_4 - KNOW the strong bases – Solubility Rules for hydroxide, OH^{-1}

$$K_w = [H^{+1}][OH^{-1}] = 1.0 \times 10^{-14} \text{ at } 25^{\circ}C$$
$$= K_a \times K_b$$
$$pH = -\log[H^{+1}], pOH = -\log[OH^{-1}]$$
$$14 = pH + pOH$$
$$pH = pK_a + \log \frac{[A^{-1}]}{[HA]}$$
$$pK_a = -\log K_a, pK_b = -\log K_b$$

What is NOT given: (you may want to put in calculator)

$$[H^{+1}] = \text{Antilog}(-pH), \text{ or } 10^{(-pH)}$$

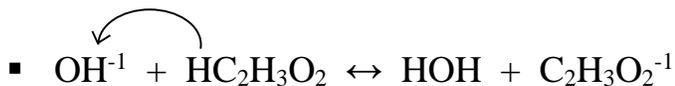
$$[OH^{-1}] = \text{Antilog}(-pOH), \text{ or } 10^{(-pOH)}$$



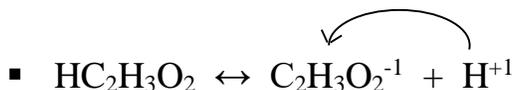
Li	Know what salts will do to a pure water solution when added – Hydrolysis of Salts	Cl
Na		Br
K		I
Rb		NO_3
Cs		ClO_3
Fr		ClO_4
	Salts that start with a metal (on the left) will behave as a BASE when placed in pure water.	
	Salts that end with a non-metal (on the right) will behave as an Acid when placed in pure water.	
	Salts that start with a metal (on left) and end with a non-metal (on the right) will NOT CHANGE the pH when placed in pure water.	

Buffer Solutions / Bronsted Lowery Acid-Base Pairs / Conjugate Acid-Base Pairs

- NO STRONG ACIDS make up a buffer solution!
- Weak acid with its salt (common ion), example: $\text{HC}_2\text{H}_3\text{O}_2$ and $\text{NaC}_2\text{H}_3\text{O}_2$
 - If strong Base is added, the $\text{HC}_2\text{H}_3\text{O}_2$ will react with the hydroxide ions, OH^{-1} :

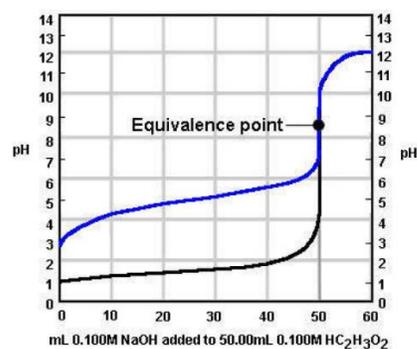
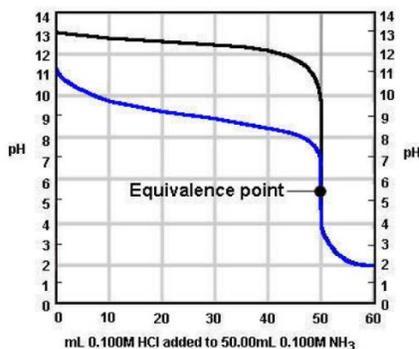
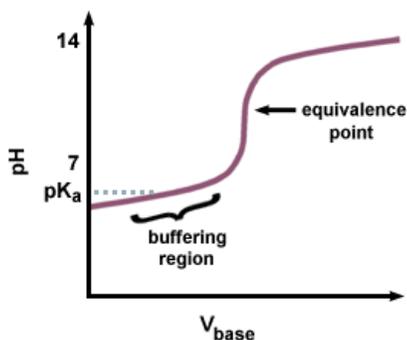


- Weak base with its salt (common ion)
 - If strong Acid is added, the $\text{C}_2\text{H}_3\text{O}_2^{-1}$ will react with the hydrogen ions, H^{+1} :



- **Important note:** Good buffers contain a weak acid (or a weak base) and its common ion. The buffer will “absorb” EITHER a strong acid (H^{+1} ions) or a strong base (OH^{-1} ions)!
- Conjugate Acid-Base Pair: The acid will have one more hydrogen atom than the base (the acid is reacting with)!
 - Examples: $\text{NH}_4^{+1} / \text{NH}_3$, $\text{HC}_2\text{H}_3\text{O}_2 / \text{C}_2\text{H}_3\text{O}_2^{-1}$, $\text{H}_2\text{PO}_4^{-1} / \text{HPO}_4^{-2}$
 - Of the conjugate acid-base pairs above, which molecules are amphoteric?
 - Amphoteric – A molecule that can act as an acid (H^{+1} donor) or as a base (H^{+1} acceptor)

More Titration Graphs



KINETICS: (37:09)

- Know how to determine the order of a Reactant(s)
- Know how to determine Rate Laws
- Be able to determine the Rate Determining Step from a reaction mechanism (steps for a reaction, fast steps and the one slow step)
- Know this statement: “To determine the Rate Law, include ALL the reactants up to, and including the slow step, minus any intermediates.”

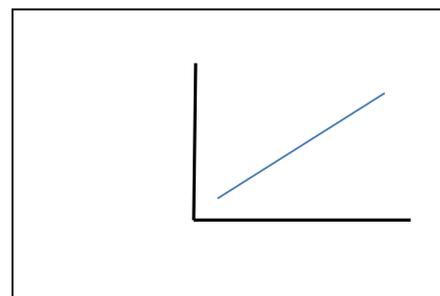
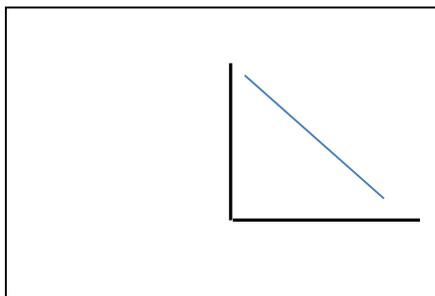
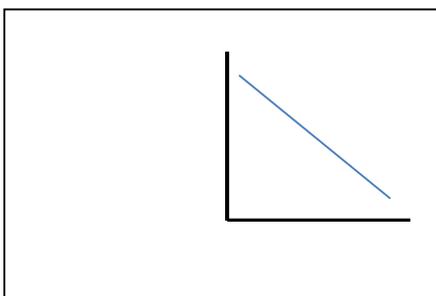
$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

- Know these Terms:
 - Reaction Mechanism (the steps in a reaction, all are fast except one slow step)
 - Reactants, Products, Intermediates, Catalysts (within a reaction mechanism)
 - Half-Life – and know how to determine it from looking at a graph! (hint: do not look at the time, look at the concentration (or percent) – you will start at time zero with ALL the material, from there look for HALF the amount... Then look at the time.
 - Rate determining step: The SLOWEST step in the reaction mechanism.

Common Mistakes:

- **Do not** try to use $\text{Rate} = k [A] [B]$ to determine “k” if you do NOT have a rate value. Use one of the equations in the equation box above!
- **Do not** use a “curve” to determine the order, know the graphs below - AND the y-axis (vertical axis) unit below! You need to focus on either the negative slope or positive slope – NO CURVES!



THERMODYNAMICS: (1:12:35)

Spontaneous = Thermodynamically Favorable

Non Spontaneous = Not Thermodynamically Favorable

Heat of Reaction:

- Can be calculated numerous ways:
 - Bond Energies (equation not given)
 - Σ Reactants minus Σ Products
 - Σ Bonds broken – Σ Bonds formed
 - Question 27 (on 2014 2nd mock exam is an “Over-Illustrated” example of Bond Energies. Energy is absorbed (endothermic) to break the bonds, and energy is lost (exothermic) to make the bonds!
 - Heat of formation of the reactants and products
 - Equation is given on equation sheet
 - Hess’s Equations (Hess’s Law of Heat of Summation)
 - Enthalpy change for any process is equal to the sum of the enthalpy changes for any set of STEPS that lead from initial to final conditions of the process

$$q = mc\Delta T$$

$$\Delta S^\circ = \Sigma S^\circ \text{ products} - \Sigma S^\circ \text{ reactants}$$

$$\Delta H^\circ = \Sigma \Delta H_f^\circ \text{ products} - \Sigma \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Sigma \Delta G_f^\circ \text{ products} - \Sigma \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= -RT \ln K$$

$$= -nFE^\circ$$

ΔH	ΔS	ΔG for the process under consideration
-	+	Spontaneous (Thermodynamically FAVORABLE) at any temperature
+	+	Spontaneous (Thermodynamically FAVORABLE) at HIGHER temperatures
-	-	Spontaneous (Thermodynamically FAVORABLE) at LOWER temperatures
+	-	NOT Spontaneous (Thermodynamically UN-FAVORABLE) at any temperature

- Heat moves from areas of HIGH heat to areas of LOW heat.
- Water can “store” more heat than most metals. In other words, water has a higher specific heat value.
- ΔS (entropy): Increases in entropy (where ΔS becomes more positive):
 - a) A move to a more disordered state
 - b) A substance dissolves (or dissociates) in water (solid to aqueous)
 - c) Increasing the number of electrons, atoms, molecules, or particles
 - d) Weaker bonds and increased “softness”
 - e) Increasing chemical complexity, example CH_4 has a lower ΔS than C_2H_6

$\Delta G = \Delta H - T\Delta S$ KNOW this equation! Be prepared to answer MANY questions about this equation.

ELECTRO / REDOX (1:23:20)

Nernst Equation: KNOW what happens when the galvanic cell is **NOT** at standard conditions (25°C, and one (1) molar concentration).

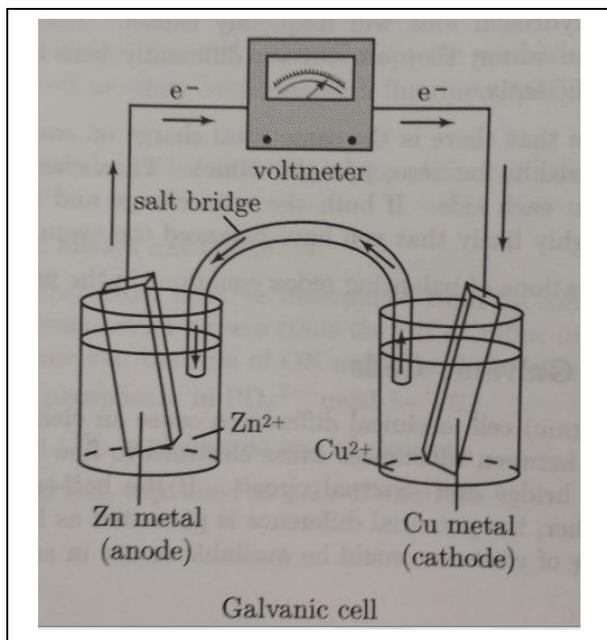
- Voltage **INCREASES** when the value for Q is **LESS** than one (1) – more reactants than products.
- Voltage **DECREASES** when the value for Q is **GREATER** than one (1) – more products than reactants.

• ONLY ions in the chambers are considered.
$$Q = \frac{[\text{anode ions}]}{[\text{cathode ions}]} = \frac{\text{site of oxidation}}{\text{site of reduction}}$$

Oxidation Numbers: Potential Charge may have when bonded with another element.

Galvanic Cell = Spontaneous, TWO chambers, Positive Voltage, Salt Bridge (flow of ions from one chamber to the other).

Electrolytic Cell, Electrolysis = Not Spontaneous, ONE chamber, Negative Voltage (yet looking for a voltage closest to zero is more optimal), Electrons must be supplied for reaction to occur.



$$\Delta G^\circ = -nFE^\circ$$

$$I = \frac{q}{t}$$

G° = standard free energy

n = number of moles

E° = standard reduction potential

I = current (amperes)

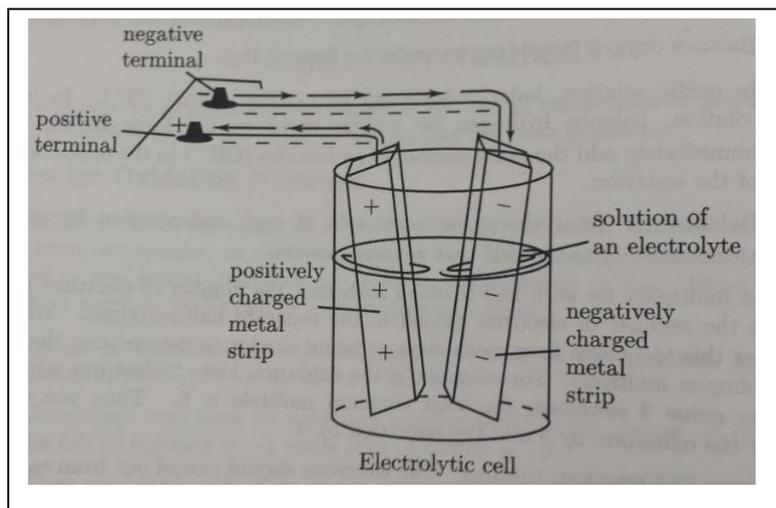
q = charge (coulombs)

t = time (seconds)

Faraday's constant, $F = 96,485$ coulombs per mole of electrons

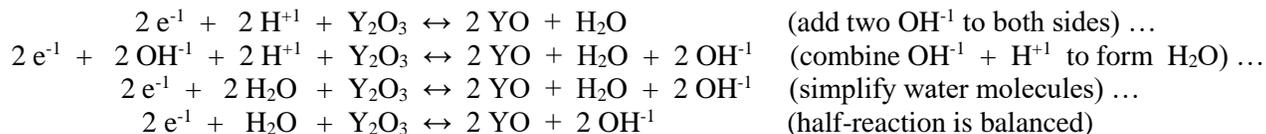
$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

$\Delta G = -nFE^0$ Expect a calculation with this equation.



ELECTRO / REDOX (CON'T)

- REDOX, ACID REDOX, BASE REDOX:
 - Oxidation – occurs at the Anode, Loses electron(s)
 - Reduction – occurs at the Cathode, Gains electron(s)
 - Metals “prefer” to lose electrons – oxidize
 - Non-Metals “prefer” to gain electrons – reduce
 - **Acidic Redox:**
 - Balance all the non-oxygen, non-hydrogen atoms first!
 - Balance the oxygen atoms by adding water molecules, H₂O
 - Balance the hydrogen atoms by adding hydrogen IONS, H⁺¹
 - Balance the charges by adding electrons
 - THE CHARGES NEED TO BE THE SAME... NOT ZERO!
 - **Basic Redox:**
 - Follow ALL the steps above (for Acidic Redox) **AND** add OH⁻¹ ions to the side that has H⁺¹ ions. Example: (see next page)



ATOMIC THEORY / PES (1:32:09)

Quantum Numbers: Address for an electron.

- First number = Energy Level, start @ one (1), "n"
- Second number = Orbital Type – THIS IS THE IMPORTANT quantum number (often referred to in questions) "l", (n - 1) → 0
 - 0 = s – orbital
 - 1 = p – orbital
 - 2 = d – orbital
 - 3 = f – orbital
- Third number = Orientation - depending on the orbital type (second quantum number), this number will indicate a possible orientation. $m_l = l \rightarrow 0 \rightarrow -l$
- Fourth number = spin.

Example: 3, 2, -1, + ½ could represent an element in the THIRD energy level, within a "d-orbital".

COULOMBIC ATTRACTIONS: $F = k \frac{q^+ q^-}{d^2}$; Translation: The force becomes greater (stronger) as the distance becomes less (smaller). This can help explain MANY periodic trends:

Example: Three ions have the same charge (+2), and the ions have ionic radius that are, 55 pm, 89 pm, & 122 pm. Which ion has a greater attraction for water molecules in a solution?

Answer: The interaction of ions with water is a Coulombic (specifically ion-dipole) interaction. Since all three ions have the same charge (+2), the strength of the interaction is related to distance, with shorter distances leading to the stronger interactions. The smaller ions therefore have stronger coulombic attraction to water.

- Ionization Energy: Energy needed to remove an electron – radius of atom becomes smaller = more energy needed to remove first electron.
- IntERmolecular: Force (bond) between MOLECULES. A stronger bond makes it difficult for molecules to change to the next state (liquid to a gas for example):
 - Higher Vapor pressures (gas pressure created by a liquid evaporating) equal **WEAKER** intermolecular forces.
 - Higher Boiling Points equal **STRONGER** intermolecular forces.

Types of Intermolecular forces:

- Hydrogen Bonding (Between MOLECULES), Polar compound, Hydrogen is bonded to N.O.F.
- Dipole – Dipole Bonding, Polar compound, and NOT Hydrogen Bonding ☺
- London Dispersion Forces, Forces between Non-polar molecules AND between ALIKE atoms when positioned next to another molecule.
- Ionic Bonding, Polar compound, Metal with a Non-Metal. Typically solids at Room Temperature.

$$E = h\nu$$

$$c = \lambda\nu$$

E = energy

ν = frequency

λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^8$ m s⁻¹

Avogadro's number = 6.022×10^{23} mol⁻¹

ATOMIC THEORY / PES (CON'T)

- **Intra**molecular: Force (bond) between ATOMS. These are the strongest types of bonds!!! These bonds are stronger than any type of intermolecular bond.

Types of Intramolecular forces:

- Covalent Bonds: Bonds where the electrons are shared UN-EQUALLY
 - Non-Polar Covalent Bonds: Bonds where the electrons are shared EQUALLY
 - Ionic Bonds, Bonds between a metal ion (Cation) and a non-metal ion (Anion)
- Electron Affinity: Attraction for other atom's electrons
 - PES Data/Graphs: Be sure that you can determine the atom from the data given in the graph. (look over review sheet given out not too long ago).

REACTIONS / SOLUBILITY RULES (1:46:11)

- **NET EQUATION:** Only show the "particles" that change / involved in the reaction! No spectators!
 - **Spectators** are particles that do NOT change (state, charge, bonding orientation, etc)
 - **Changes include** but are not limited to:
 - State change: aqueous to solid (ppt formed), solid to aqueous, liquid to a gas, etc.
 - Charge change: gains or loses electrons
 - Switches "partners": bonded with a different atom(s)
 - Changes color – difficult to know without being told so
- Be sure you READ and LOOK at the **GIVEN** BALANCED EQUATION (for STATES and COEFFICIENTS)

Only write the MOLECULAR shape (shape the atoms form) when naming a shape!

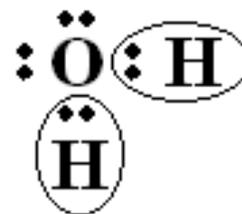
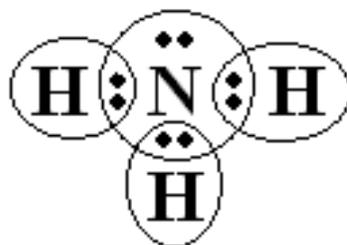
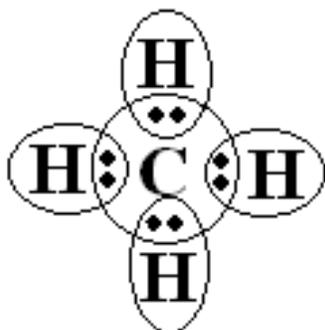
sp 180°	sp ² 120°	sp ³ 109°	dsp ³ 90° & 120°	d ² sp ³ 90°
	Bent		See - Saw	
		Bent		Square Planar
			Linear	

DRAWING LEWIS STRUCTURES:

- Count the TOTAL number of valence electrons in the molecule.
- Determine the skeletal structure (Carbon is ALWAYS a central atom, Hydrogen is ALWAYS an external atom).
- Be sure to give the external atoms eight electron (give hydrogen two)
- Whatever electrons remain, place on the central atom – be sure that the central atom has AT LEAST eight electrons.
- Of course, obey formal charges! As external atoms:
 - Group 5A = triple bonds or three bonds (with one lone pair of electrons)
 - Group 6A = double bonds or two bonds (with two lone pairs of electrons)
 - Group 7A = single bond or one bond (with three lone pairs of electrons)
- Single Bond has one SIGMA bond
- Double Bond has one SIGMA bond and a PI bond
- Triple Bond has one SIGMA bond and two PI bonds
- Resonance Structures:

- **NOT IN THE VIDEO... BUT IMPORTANT TO KNOW!**

Bond Angles when there are lone pairs of electrons bonded to the central atom:



Lone pairs of electrons will “push” the external atoms closer to each other – accounting for the change in the bond angles.

CHEMISTRY BASICS... SOMETHING EVERY CHEMISTRY STUDENT SHOULD KNOW! (1:53:35)

Gases: $PV = nRT$; $PV = \frac{m}{M}RT$; conditions at STP

- Average Kinetic Energy = Temperature
- ALL (any) gases at the same temperature, pressure, and volume will have the same number of particles!
- Partial Pressures. Total pressure of a sample of gases is equal the sum of the pressures of each gas. $P_T = P_1 + P_2 + P_3 + \dots$
 - To determine partial pressures on a reaction that starts with one gas to form two gases at equilibrium, use this simple trick:
 - Example Reaction: $A_{(g)} \leftrightarrow B_{(g)} + C_{(g)}$
 - Gas "A" has an initial pressure of 1.50 atm, at Equilibrium the total pressure is 2.20 atm – calculate the partial pressure of each gas at Equilibrium.
 - $1.50 + x = 2.20$; $x = 2.20 - 1.50 = 0.70$; "x" equals the partial pressure for **EACH** "B" & "C" gases.
 - So to find the partial pressure of gas "A", $2.20 - (2 (0.70)) = 0.50$
- Know how to solve for Gas Stoichiometry Problems!
- Diatomic molecules, H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 (liquid at room temp), I_2 (solid at room temp)

Stoichiometry:

- 1) Balance chemical equation
- 2) Convert what is "given" into moles
- 3) Use mole ratio
- 4) Convert moles to "desired" unit

LIMITING REACTANTS: Limiting reactants deal with a reactant being completely consumed (reacted). Keep solubility rules in mind when thinking about L.R. problems!!!

- Soluble ions will be in solution – whether from the L.R. or Excess reactant !!!
- Precipitates will form solids (at the bottom of a container), BUT if ions that make up the ppt are from a reactant in excess, YOU WILL HAVE SOME OF THOSE IONS still in solution!

Empirical Formulas:

- Smallest Whole Number Ratio (obtained from number of moles of EACH ELEMENT)
- Molecular Mass Units: Grams / Moles , example; Oxygen gas = 32 g / mol

Reading Graphs / Charts, Drawing Graphs

