

AP All Unit Miscellaneous Notes

Semester 1

Unit 1: Foundations

A = mass #, Z = atomic # = # of protons

e/m = charge to mass ratio of an electron, found by J.J. Thompson using a Cathode Ray tube.

Hot and Cold liquids can have different densities and volumes.

Since a hot liquid tends to expand, its volume will increase, thus decreasing its density.

Unit 2: Stoichiometry

For the Multiple Choice section on the AP test, you will not be allowed calculators. Often there are short stoichiometry problems on the multiple choice section. You need to look at the numbers in the stoichiometry and cancel/reduce the numbers in the problem. After canceling/reducing the numbers, it should be simple enough to work out the problem in your head or with simple multiplication. **Pay attention to short cuts, for example, doing radioactive decay with fractions and not the $\ln A = -kt + \ln A_0$ equation. See the Radiochemistry chapter below.

Unit 3: Reactions in Solutions

Disproportionation means something is both oxidized and reduced.

Solubility Graphs: The graph should be at its' steepest to recover the most solid while cooling. The steeper it is, the larger the difference in solubilities, so it will form more solid as it cools over this temperature range. (Less solid will form, if the graph is flatter, because the solubilities will be more similar.)

If you are trying to separate two compounds, you would want one line to be flat in a temperature range, while the other has a line with a steep slope in that same temperature range. The steep sloped one would form a lot of solid, if cooled, while the other would not form much, if any, solid.

Remember Intensive properties: If you have 1 L of a 0.500 M solution and then you pour out 50 ml from that 1 L, the molarity would still be 0.500 M.

Spectator Ions are the ions that cancel out in the complete ionic reaction, going to a net ionic reaction. They are watching (spectators) the other ions react to form solids or liquids. Spectator ions can also be canceled out of redox reactions, if they themselves are not oxidizing or reducing. In redox reactions, you usually see MnO_4^- , not KMnO_4 , because the K^+ ion has been cancelled out in the reaction.

Unit 4: Gases

****The larger the mass of the gas**, the larger the volume, so the **less ideal** (the less like a point mass).

****The higher the temperature, and the lower the pressure, the more ideal the gas.**

(High temperature and low pressure means the gas will spread out. The larger the volume of the container that the gas is in, the more the gas acts as a "point mass", the more ideal.) {High Temperature / Low

Pressure also had the most entropy in the energy chapter!)

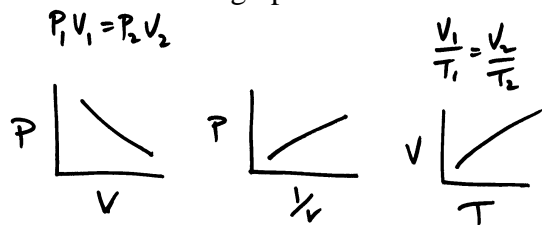
The hotter the gas, the higher the average velocity, so the gas will have a **larger range of velocities**.

Any gases at the same temperature, will have the same KE.
(KE = 3/2 RT, only temperature affects the Kinetic Energy.)

NH₃ is more soluble in water, than CO₂.
(NH₃ is polar, while CO₂ is non-polar (H₂O is polar.) Like dissolves like!)

Know the Greenhouse gases: CH₄, H₂O, CO₂, O₃, CFC's, N₂O.

** Know how the graphs look for the various equations derived from the Ideal Gas Law.



PV is an indirect relationship!
(inverse)

(one increases, while the other decreases)

V/T is a direct relationship!

(both increase or both decrease)

Dalton's Law of Partial Pressures: $P_1 + P_2 + P_3 \dots = P_{\text{total}}$

Since the partial pressures are related by how many mols of gas are present,

$P_1/P_{\text{total}} = \text{mol}_{\text{gas \#1}} / \text{total mols of gases} = X$, which is the mol fraction of gas #1.

So the mol fraction of the gas times the P_{total} will equal the Pressure of the gas: $X_{\text{gas}} \cdot P_{\text{total}} = P_{\text{gas}}$

{This is similar to the pressure of a solution: $X_{\text{solvent}} P_{\text{solvent}} + X_{\text{solute}} P_{\text{solute}} = P_{\text{solution}}$

The P_{solvent} here would be like P_{total} of the solvent, P_{solute} here would be like P_{total} of the solute.}

For a mixture of gases: $X_{\text{gas\#1}} P_{\text{total}} + X_{\text{gas\#2}} P_{\text{total}} + X_{\text{gas\#3}} P_{\text{total}} = P_{\text{total}}$, where $X_1 + X_2 + X_3 = 1$

Unit 5: Thermodynamics/Thermochemistry/Energy

Law of Dulong and Petit: (molar mass) (specific heat) $\approx 25 \text{ j}/(\text{mol} \cdot ^\circ\text{C})$

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

Remember, if it is at equilibrium, then $\Delta G = 0$, so

$$0 = \Delta H - T\Delta S$$

$$\frac{\Delta H}{T} = \Delta S$$

This equation works at equilibrium, which could be melting/freezing, boiling, etc.

Unit 8: Solids/Liquids

Carbon has 3 **allotropes** (or forms): diamond, graphite, charcoal

Fe₂O₃ (iron III) is rust.

For something to dissolve in water, the $\Delta H_{\text{solution}}$ should be (-) or a very small (+). It cannot be very endothermic, because it will need too much energy to dissolve.

$$\Delta H_{\text{solution}} = -(\text{Lattice Energy}) + \text{Hydration Energy}$$

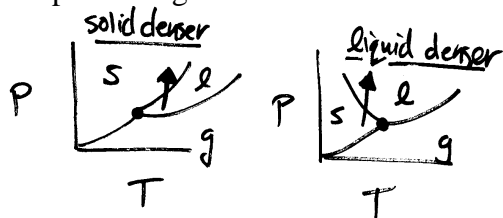
Since the lattice energy is usually (-), that term will become (+)

$$= -(-) + \text{Hydration Energy}$$

$$= (+) + \text{Hydration Energy}$$

For $\Delta H_{\text{solution}}$ to be (-), the Hydration Energy needs to be negative and larger than the Lattice Energy!

On phase diagrams:



If the line going between solid and liquid has a positive slope, the solid is denser than the liquid.

Increasing pressure, going up across the positively sloped line, goes from liquid to solid, so the solid is denser.

(This occurs in most substances, the solid sinks in the liquid.)

If the line going between solid and liquid has a negative slope, the liquid is denser than the solid.

Increasing pressure, going up across the negatively sloped line, goes from solid to liquid, so the liquid is denser.

(This occurs in water; the solid ice floats on the liquid.)

On a graph of freezing point, the solution's temperature may go below the freezing point and then come back up to the freezing point to form a flat line. It may go below, because it may need to get colder than the freezing point, before the solid crystals start sticking together to form solid. Remember during the freezing point lab, sometimes as you stirred the cold solution, it suddenly became slushy/froze, as you pushed the microscopic crystals together, so that larger crystals could form. When it suddenly became slushy/froze, the solution warmed back up to the normal freezing point.

Diagonal relationships for radius, ionization energy, electron affinity and electronegativity:

Radius increases going down and to the left.

(Diagonal would point to bottom left corner for the largest radius

(Fr-most active metal, most metallic).)

I.E., E.A. & Electronegativity increase going up and to the right.

(Diagonal would point to the top right corner for the largest energies

(F)-most active non-metal, most non-metallic.)

**see the notes!

Group Relationships:

Chemical Properties: reactivity

Metals are more reactive going down, easier to lose e⁻,

While nonmetals are more reactive going up, easier to gain e⁻.

Metallic Character: more metallic going down, more non-metallic going up {see above}

Electronic Structure: each group loses or gains the same number of electrons.

Physical Properties: trends in melting/boiling points, density, intermolecular forces, solubility etc.

Periodic Relationships:

Atomic radius: smaller to the right

Ionic radius (Ions): the more (+) charge, the smaller the radius & the more (-) the charge, the larger the radius.

Unit 9: Kinetics and Equilibrium

The **most common catalysts are Pt or Pd**, others include Au, enzymes, MnO₂. Catalysts are usually metals that can oxidize or reduce.

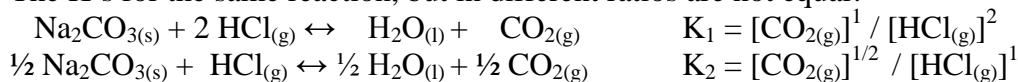
Increasing the temperature always increases the rate of reaction, whether it is endothermic or exothermic, since it adds more energy for E_{act}.

Increasing temperature will shift the equilibrium one way or the other, depending on whether the reaction is endothermic or exothermic.

Reaction Order can only be determined by experimentation (doing a lab).

The Activation Energy of a reverse reaction is not the same as for the forward reaction. Look at the Activation curve in the notes. The "hill" is two different heights going in the two different directions.

The K's for the same reaction, but in different ratios are not equal!



$$\begin{aligned} ([\text{CO}_{2(g)}]^{1/2} / [\text{HCl}_{(g)}]^1)^2 &= [\text{CO}_{2(g)}]^1 / [\text{HCl}_{(g)}]^2 \\ \text{so } (K_2)^2 &= K_1 \end{aligned}$$

*If $K < 1$, then there are more reactants, than products. { $K = \text{products} / \text{reactants}$ }

*If $K > 1$, then there are more products, than reactants.

Don't forget the Beer-Lambert Law:

$$A = \epsilon \vartheta c$$

A = absorbance ϵ = molar absorptivity ϑ = path length (1 cm) c = concentration of the sample

$$A_1 = \epsilon \vartheta c_1 \quad A_2 = \epsilon \vartheta c_2$$

$$\frac{A_1}{c_1} = \epsilon \vartheta \quad \frac{A_2}{c_2} = \epsilon \vartheta$$

$$\frac{A_1}{c_1} = \frac{A_2}{c_2} \quad \text{Absorbance is directly proportional to concentration.}$$

For the spectrophotometer,
the optimum wavelength would have the highest absorbance and the least transmittance.

Unit 10: Acids & Bases

The higher/larger the ratio of Oxygens to Hydrogens, the stronger the acid.

Remember, if the #O - #H \geq 2, it is a strong acid.

{The more oxygens present, the easier to stabilize the negative charge, once the H⁺ leaves. So the more oxygens present the more stable the (-) Conjugate Base, so more H⁺ can leave, making a stronger acid. Also the element attached to the oxygens will usually have a larger oxidation number, attracting the oxygens more tightly. i.e. ClO₄⁻¹ is more stable than ClO₃⁻¹ and the Cl has a larger oxidation number in Cl⁺⁷O₄⁻¹ than in Cl⁺⁵O₃⁻¹.} HClO₄ is a stronger acid than HClO₃.

So, the fewer oxygens, the stronger the conjugate base (the less stable), the more the CB wants to gain H⁺ to become stable, the H's do not want to leave the acid, the weaker the acid.

(The larger the K_a, the stronger the acid. The larger the K_b, the stronger the base.)

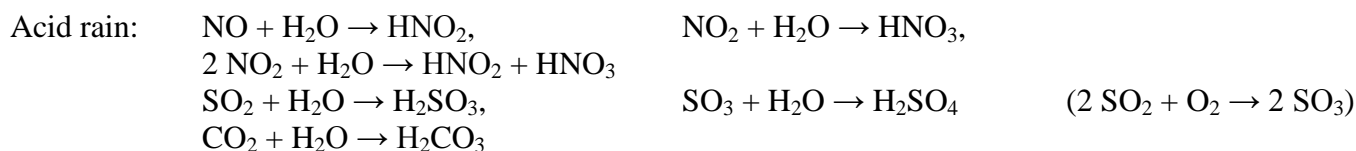
The larger the K_{sp}, the more soluble the substance!

Decreasing the temperature, decreases the solubility, so it decreases the K_{sp}.

Non-aqueous Acid Solutions:

The bond strength determines how much each acid will fall apart in non-aqueous solutions. HCl has a stronger bond, than HI. (Cl⁻ has a higher charge density (smaller radius) and will be more strongly attracted to the H⁺.) The HI will fall apart more, since the I⁻ will have a lower charge density (larger radius) and will be less attracted to the H⁺. Thus, **the HI will be a stronger acid, than HCl, in non-aqueous solutions!**

In water solutions, both are strong acids and should completely fall apart, essentially making them equal in strength. {But the HI should fall apart more easily, since the I⁻ (larger radius) will stabilize the negative charge better, than the Cl⁻ (smaller radius), technically making HI stronger.} Larger ions tend to be more stable, since the charge can be distributed over more space!



Common Lewis Acids: BH₃

Common Lewis Bases (have extra electron pairs): :NH₃

(Neutral nitrogen compounds have a lone pair of electrons on the nitrogen.)

(NH₄⁺ is the conjugate acid (Bronsted-Lowry Acid) of NH₃, but NH₄⁺ cannot be considered a Lewis Acid, because it cannot accept anymore electrons. N cannot have more than 4 bonds, since it doesn't have any "d" orbitals for hybridization! Only the 3rd row down can have more than 4 bonds, since they can use the 3d orbitals for hybridization.)

Indicators:

For a titration, the pH of the equivalence point should be in the pH range of the indicator. (or at least ± 1 of the range)

The pH of the equivalence point can be found by writing a reaction and solving for the pH, using equal amounts of the Acid and Base used in the titration. (Remember the equivalence pH is 7 only for a strong acid mixed with a strong base! See the titration curve notes!)

i.e. $\text{HNO}_2 + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{NO}_2^-$ (weak acid and strong base) Solve this to find the equivalence point pH.
0.1 M 0.1M

For the **greatest buffer action** a titration graph should be at its' **flattest**, when the acid and base are mixing **before the equivalence point!** {It is before the equivalence point, because usually a strong base would be added to a weak acid, or a strong acid would be added to a weak base. If the weak acid or base was added to the strong base or acid, then the flattest part would be after the equivalence point.} (This point will resist pH changes the most. That is why the graph (pH) is flat for a period of time, even though acid or base is added in the titration.)

A buffer has to be a weak acid with its' CB or a weak base with its' CA.

****Strong acids with CB or strong bases with CA do not form buffers**, since strong acids and strong bases completely fall apart! There is no equilibrium.

$\Delta G^\circ = -RT \ln K_p$ If K_p is >1 (many products), then $\ln K_p$ will be (+),
making ΔG (-), which is spontaneous.

$\Delta G^\circ = -RT \ln K_p$ If K_p is <1 (many reactants), then $\ln K_p$ will be (-),
making ΔG (+), which is nonspontaneous.

$$\Delta G^\circ = -RT \ln K_p \qquad \Delta G = \Delta H - T\Delta S$$

$$-RT \ln K_p = \Delta H - T\Delta S$$

$$\ln K_p = \frac{-\Delta H}{RT} + \frac{\Delta S}{R}$$

Unit 11: Electrochemistry and Radiochemistry

SCE, saturated calomel electrode is a common reference electrode in electrochemistry.

Bauxite is a source of aluminum.

A Faraday is 96485 C/ mol e^- . But a problem can give a Faraday of something, like a Faraday of electrons, which would be 96485 e^- .

Nernst Equation

$$\mathcal{E} = \mathcal{E}^\circ - 0.0592/n \log ([\text{products}]/[\text{reactants}])$$

To increase voltage the fraction needs to be less than one, so that the log of it will produce a negative.
(Decrease products or Increase reactants)

To decrease voltage the fraction needs to be greater than one, so that the log of it will produce a positive.
(Increase products or Decrease reactants)

****Nernst Equation when at equilibrium:**

$$\varepsilon = \varepsilon^{\circ} - 0.0592/n \log Q$$

Remember $\Delta G = \Delta G^{\circ} + RT \ln Q$,

$$\Delta G^{\circ} = -RT \ln K_p \text{ therefore,}$$

$$- \varepsilon^{\circ} = - 0.0592/n \log K_{eq}$$

but at equilibrium $\Delta G = 0$, so the equation would be

** (see notes for how Nernst equation was derived)

$$\frac{n\varepsilon^{\circ}}{0.0592} = \log K_{eq}$$

A large K_{eq} , would mean that there are a lot of products, so the reaction would go to completion.

**If the temperature is raised, the reaction will go faster to achieve equilibrium, but the cell potential is intensive and the voltage will not change!

When the equilibrium is reached, there will be no flow of electrons, the voltage is 0 V. (This is when the battery is used up.)

Remember half-life short cuts:

After one half-life, there is $\frac{1}{2}$ of the original left (50%).

After two half-lives, there is $\frac{1}{2} \times \frac{1}{2} = \frac{1}{4}$ of the original left (25%).

After three half-lives, there is $\frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} = \frac{1}{8}$ of the original left (12.5%).

(4 half lives: $1/16 = 6.25\%$, 5 half lives: $1/32 = 3.125\%$ etc.)

For dating objects using $\frac{1}{2}$ lives of radioisotopes:

There should be at least 10% decayed and 10% remaining of the substance.

So the $\frac{1}{2}$ live should be fairly similar to the age of the object that is to be dated.

α -particles will go towards a negative plate, since the α -particle is a helium nucleus ((+) charge on nuclei).

β -particles will go towards a positive plate, since the β -particle is an electron (-).

γ -particles (gamma) will not be affected by charged plates, since electromagnetic energy (photon-light energy) is not charged.

Unit 12: Introduction to Organic and Biochemistry

**Make sure you look at the Organic notes and can figure out the formula of the functional group compounds from their names. Often the functional group compounds come up in the reaction problems, often in the form of combustion reactions.

End of Notes