NOTES#20/Quantum Theory#2/ AP CHEMISTRY

I. Emission Spectra - ____. There are two types:

a. Continuous spectra - when ___________________ wavelengths of visible light are represented.

Ex of continuous spectra:

**b. Line Spectra - when only specific wavelengths are emitted.

- All elements have a unique emission spectrum (like a fingerprint)

- Ex of line spectra for Hydrogen

V	I	B	G	Y	O	R

- What's going on in an emission spectra anyway?

- Flame tests - the unique color observed results from the combination of the colors emitted in the visible region. (We don't have a way of separating the wavelengths when doing flame tests)

*** So, why is it that different elements produce different emission spectra?

II. Bohr's Model and the Hydrogen Emission Spectra:

- According to Bohr, a single electron in the hydrogen atom could be located only in certain energy levels. Energy levels are labeled n=1, n=2, n=3, n=4.

- As you go farther away from the nucleus, the energy of each level __________________________.

- n=1 represents the LOWEST and MOST STABLE energy level or the __________________________.

- n > 1 represent HIGHER ENERGY LEVELS or ____.

A. The Emission Process: *The Conceptual Approach* - Hydrogen's electron prefers to "hang out" in it's

ground state, however, if given enough energy, ____. Electrons don't like being in excited energy levels, so the hydrogen electron will drop back down. As the electron drops down to a *lower* energy, a photon of energy is _________________ in the form of radiation.

** Remember, to jump to higher energy levels, energy must be __________________________.

To fall down to lower energy levels, energy must be ________________________.

B. The Emission Process Revisited: *The Mathematical Approach* - We can calculate the amount of energy associated with different energy level transitions in the hydrogen atom.

- The energy that an e^- possesses in a certain energy level can be determined by the following eq:

$E_n = -Rh\ (1/n^2)$ ** Rh = Ryberg's Constant=2.18 x 10 -18J

** n = principle energy level

** As n increases, what happens to E_n? ____

** A larger negative energy value means ____

Close to nucleus = stability (the greater the - #) $E_n = 0$ when n = infinity . . . a free electron

- What if we want to calculate the energy absorbed/emitted during different energy level transitions? For example, what is the energy of a photon emitted when an electron drops from a **n = 4** energy level to the **n = 2** energy level? What would it's wavelength be? $\Delta E = E_f - E_i$......(since E is a state function....)

C. The Emission Spectra of Hydrogen: The emission spectra of hydrogen includes a wide range of wavelengths from the IR to the UV. Each different wavelength corresponds to a different allowable energy level transition.

* Label the spectrum region of each series.
* Series are named after scientists.

Problems:

1. Consider the following energy levels for the hydrogen atom:

   a. How many emission lines are possible?

   b. Which transition produces photons of the greatest energy?

   c. Which transition for the H atom produces the emission line with the longest wavelength?

2. What is the wavelength (in nm) of a photon emitted during a transition from $n_i = 6$ to $n_f = 4$ state in an H atom? What region of the electromagnetic spectrum corresponds to this wavelength?

D. Fluorescence - atoms absorb UV light in their electron structures and release it as energy (photons) with longer wavelengths (visible light spectrum)

E. Phosphorescence - same as fluorescence except the absorbed light is greater and this energy released (photons) will continue to be emitted for some time after the light source is turned off.