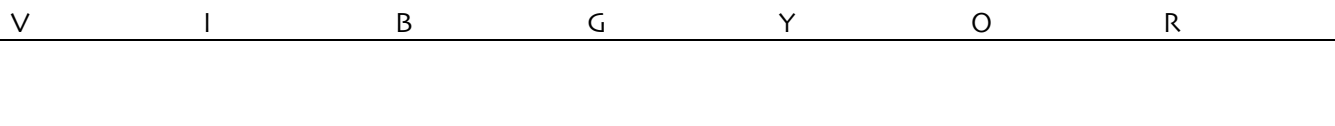


- I. Emission Spectra - _____. There are two types:
- Continuous spectra - when _____ wavelengths of visible light are represented.
Ex of continuous spectra:

- Line Spectra - when only specific wavelengths are emitted.
 - All elements have a unique emission spectrum (like a fingerprint)
 - Ex of line spectra for Hydrogen



- 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:

$$E_n = -R_h (1/n^2) \quad \begin{array}{l} ** R_h = \text{Ryberg's Constant} = 2.18 \times 10^{-18} \text{J} \\ ** n = \text{principle energy level} \end{array}$$

** 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 $n = \text{infinity}$... 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 its 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.