Section 2 Notes: The Hydrogen Atom and the Bohr Model (p. 305-309)

Before beginning our notes, I would like to give you the value of Planck's Constant (h) which I forgot to include in our first section of notes. **Planck's Constant (h) = 6.626 x 10^{-34} J s.**

(Watch Video #5) This video is also important for the next Section (3) notes (quantum theory)
Only watch first 15 min of this video *for Section 2 notes.*
HW: #58, 60, 62, 64, 68 (Due Tuesday, April 14, at 12:00 midnight)

When "white light" or visible light is passed through a prism, it breaks up into or is bent into its wavelengths. White light contains photons of all visible light. As a result, a **continuous spectrum** (rainbow) is produced in which the colors of the different wavelengths and frequencies "blend" into each other. (**See p. 305**)

When hydrogen gas is heated (by fire, electricity, hot metal, etc.), the hydrogen (H_2) atom break apart and the electron on the H-atom **absorbs** some of the energy (**becomes excited**). This causes the electron to "jump" to a higher **energy level (n). Energy level is represented by "n" and does not mean "moles" as it usually does in stoichiometry and gas laws.**
[Energy levels are regions of space around the nucleus of an atom. The energy level closest to the nucleus has the lowest energy and is energy level (n) 1. The further an electron is from the nucleus, the higher its energy.]
Remember that the second law of thermodynamics says that all things in nature go to lowest energy. (on the front board in my room)
Therefore, electrons in their **excited state** are unstable and try to go back to their original lowest energy level **(ground state)**. When an electron goes down to a lower energy level it emits energy **always in the form of "light".** In the case of Hydrogen, the ground state for its outermost electron is n=1.

When this happens in Hydrogen, the electron "jumps" back to its ground state in one or more energy level jumps. Each jump down to a lower energy level **emits** distinct or exact packets of energy **(quantized) or photons of light.** The electrons of different elements emit photons specific to the element. Hydrogen emits 4 distinct "bands" of visible light in a spectrum (**see p.305-306**). This spectrum is called a **Bright-line Spectrum** because of the bands of distinct colors. Eventually the electron will go back to its original energy level. In the case of hydrogen, it is energy level 1.

Neils Bohr was the scientist who discovered that electrons move up and down within an atom **absorbing** (moving up to higher energy levels) or **emitting** (moving down to lower energy levels) energy in distinct packets of energy called **photons** and coining the expression **"quantized" energy.** This is the beginning of the quantum theory.

Bohr said that no matter where and electron is in the atom, it possesses a certain **quanta** of energy. Bohr devised and equation for the hydrogen atom that could determine the Energy (E) and the energy levels available to an electron in hydrogen.

The equation is **$E = -2.178 \times 10^{-18}$ J (Z^2/n^2)**

-2.178×10^{-18} J (a constant) is the energy of an electron of hydrogen on the 1st Energy level (n=1). It is also the energy needed to remove the electron from the H-atom.

Z is the charge on the nucleus of the atom. For hydrogen Z = +1 or just 1.

n is the energy level where the electron of hydrogen is located in either its ground or an excited state.

So…

For an electron on the 1st energy level $E = 2.178 \times 10^{-18}$ J $(1^2/1^2) = -2.178 \times 10^{-18}$ J

For an electron on the 6th energy level $E = 2.178 \times 10^{-18}$ J $(1^2/6^2) = -6.050 \times 10^{-20}$ J **(See p.307)**

To find the energy releases (emitted) when an electron goes from an excited state (**6th for this example**) to lower state **(in this case, ground state)**,

$\Delta E = (E_{final} - E_{enitial})$ or $\Delta E = -2.178 \times 10^{-18}$ J – (-6.050×10^{-20})

$\Delta E = (E_1 - E_6)$ $= -2.117 \times 10^{-18}$ J

Bohr realized that the wavelength λ for the quantum of light emitted by an electron going down to a lower energy level in a hydrogen atom could be calculated with Planck's equation

$\Delta E = h(c/\lambda)$ Rearranged to **$\lambda = h(c/\Delta E)$**

So for the above equation, and electron going from the 6th energy level to the 1st energy level of hydrogen would have a wavelength λ of

$\lambda = h(c/\Delta E)$

$$\lambda = \frac{(6.626 \times 10^{-34}\ Js)(3.00 \times 10^{8} m/s)}{(2.117 \times 10^{-18}\ J)} = 9.39 \times 10^{-8}\ m$$

Instead of calculating the energy of each electron then subtracting them for ΔE, we have an equation to solve the ΔE knowing the two energy levels the electron travels.

Remember, if an electron goes from a lower n to a higher n, the change in E will be (+) because it's absorbing energy to go up. If the electron goes from higher n to lower n, the change in E will be (-) *because it releases or emits energy going down.*

This equation is

$$\Delta E = -2.178 \times 10^{-18}\ J\ (1/n_f^2 - 1/n_i^2)$$

Another equation for calculating the **λ wavelength** when we don't have E, is known as **Rydberg Equation:**

$\lambda = 1/[R_H(1/n_f^2)(1/n_i^2)]$ where R_H is **Rydberg's constant ($1.0968 \times 10^{-2} nm^{-1}$)** *(answer is in nm)*

Rydberg's Equation gives the wavelength of light emitted.

If an electron absorbs enough energy to "fly" off the atom (**ionization energy**—energy needed to remove an electron from an atom), the n_f would be (∞) and **$1/(\infty)^2 = 0$.** (See p.309)

So for Hydrogen atoms, ΔE = -2.178×10^{-18}J $(1/\infty^2 - 1/1^1) = -2.178 \times 10^{-18}$J (0-1) **$= 2.178 \times 10^{-18}$J**