Chemistry – Chapter 4 Partial Study Guide & Review Assignment

<u>REVIEW ASSIGNMENT: Pages 118-120 in the book, complete #18-20, 23, 30-32, 35, 39, 40, 41-42, 43a-f, 44-46, 48</u>

DUE MONDAY, 12/16 (THE DAY OF THE CHAPTER 4 TEST)

What is on this study guide is a GUIDE—it is not exhaustive. <u>You should use your Ch. 4 Learning Reflection Packet as your official study guide.</u> Study all your assignments from this section and any notes taken.

Section **Objectives** **Key Terms**

4.1 *Explain the mathematical relationship between the speed, wavelength, and frequency of electromagnetic radiation.

- ***$c = \lambda v$ (speed of light = wavelength x frequency); c =* 3.0 x 10^8 m/s**
  - **When given a problem and asked to find either wavelength or frequency, you will typically use this equation.**
- **$E_{photon} = hv = hc/\lambda$ *(Energy of a photon = Planck's constant x frequency = Planck's constant x speed of light / wavelength)***
  - **If asked to find the energy of a photon, use one of these equations (the equation you use will depend on whether you are given frequency or wavelength in the problem. If you can, always choose E=hv because it is simpler).**
  - **If asked to find the energy of a mole of photons, you will multiply the energy of a single photon by # of photons in a mole (6.022 x 10^{23})**

*Discuss the dual wave-particle nature of light.

- **Einstein and Planck's research led to the concept that light can behave as both a wave (as it travels through space with wavelength and frequency) and a particle (photoelectric effect – see notes handout)**

*Discuss the significance of the photoelectric effect and the line-emission spectrum of hydrogen to the development of the atomic model.

- **The photoelectric effect proved that light of different colors had different wavelengths and, therefore, different energy which explained why only colors with higher frequencies (and therefore energies) are able to knock electrons off of metal.**
- **Each element has an identifying color it gives off when its electrons are excited by absorbing energy. The electrons absorb energy in the form of heat or light, causing them to jump to higher energy levels. They are unstable here, so they fall back to ground state and give off a color of light that matches the energy different between the two energy levels**

jumped. Taken together, all the colors given off from a specific element make up its spectra, which is used to identify elements in distant space.

__Key Terms:__

- *Electromagnetic radiation, Electromagnetic spectrum, Wavelength (λ), Frequency (v), Photoelectric effect, Quantum, Photon, Ground state, Excited state*

4.2

*Explain how the Heisenberg uncertainty principle and the Schrödinger wave equation led to the idea of atomic orbitals. *See notes taken from the board on this topic (in notebook).*

- **Heisenberg said you can never know exact location and speed of an electron at the same time (you can know speed OR location, but never both at once).**
- **Schrodinger used wave functions to calculate regions of space around the nucleus where you are most likely to find an electron at any given time (the idea of orbitals).**

*List the four quantum numbers, and describe their significance.

- **1st quantum number shows which energy level or shell each electron is in**
- **2nd quantum number indicates there are 4 different types of "sublevels" including S P D and F, which each have different shapes and can fit different numbers of electrons.**
- **3rd quantum number shows how many different orientations each of the different sublevels (s, p, d, and f) can have in 3-D space.**
- **4th quantum number shows that electrons can have a spin or "+1/2 or -1/2"**

*Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

- **See table in notes.**
- **Understand that in 1st energy level, you can only have an s sublevel; in 2nd you can have s and p sublevels, and so on**
- **Understand there is only ever 1 s orbital (or orientation) in the sublevel, 3 for p, 5 for d, and 7 for f**
- **You can fit 2 electrons in each of those orbitals so that you can fit 2 total electrons in s sublevel, 6 in p, 10 in d, and 14 in f.**

4.3

*List the total number of electrons needed to fully occupy each main energy level.

- **You can fit 2 electrons in each orbital so that you can fit 2 total electrons in s sublevel, 6 in p, 10 in d, and 14 in f.**

*State the Aufbau principle, the Pauli exclusion principle, and Hund's rule.

- **See your Electron Configuration Packet for these (Aufbau – electrons are lazy—they will fill up orbitals in order of the lowest energy orbital to highest; Hund – electrons are unfriendly. The prefer to be unpaired if possible, so each orbital in a sublevel will fill in with up spin electrons first, then down; Pauli Exclusion Principle says that no two electrons will ever have the same quantum numbers (or the same location in the atom) so two electrons in the same orbital must have <u>opposite spins—one up, one down</u>.)**

*Describe the electron configurations for the atoms of any element using orbital notation, electron-configuration notation, and, when appropriate, noble-gas notation.

- **Practice electron configurations**

<u>Key Terms:</u>

- *Aufbau Principle, Paul Exclusion Principle, Hund's Rule, electrons configuration, highest-occupied level, inner-shell electrons, noble gases, noble-gas configuration*

Be able to do calculations using the following formulas:

$c = \lambda v$ $\qquad$ $c = 3.0 \times 10^8$ m/s (speed of light)

$E_{photon} = hv$ or $E_{photon} = (hc) / \lambda$ $\qquad$ $h = 6.626 \times 10^{-34}$ J·s

$E_{mole\ of\ photons} = E_{photon} \times 6.022 \times 10^{23}$ photons/mol

1 nm = 10^{-9} m
1 μm = 10^{-6} m