CHAPTER 7: HW Pp.303-310 **Set 1**: 8, 10, 16; **Set 2**:20, 28, 32; **Set 3**: 34, 40, 58; **Set 4**: 70, 78, 90; **Set 5:** 102, 104, 110; **Set 6:** 124, 130, 132.

I. Arrangement of the electrons in the atom
   A. Properties of Light
      1. Light is a type of electromagnetic radiation
      2. EM radiation is energy traveling in waves at the speed of light ($3.0 \times 10^8 m/s^2$)
      3. Examples are: radio waves, visible light, UV light, infrared light, x-rays, microwaves, gamma rays
   B. Terminology of waves :
      1. crests- tops of waves
      2. troughs - bottoms of waves
      3. wavelength - distance from crest to crest , $\lambda$, measured in cm, nm
      4. frequency - how often a wave travels past a certain point, $\nu$, measured in cycles per second, $seconds^{-1}$, Hertz
      5. amplitude - the height of the wave from the origin
      6. If the frequency of a wave is multiplied by its wavelength we get the speed of light (**c**). $\lambda\nu = c$ Frequency and wavelength are inversely proportional.
      7. Visible light is a continuous spectrum of colors, ROY G BIV. In order of decreasing wavelength and increasing frequency.
   C. Particle Description
      1. Planck's Quantum Theory
         a. Metals heated emit electromagnetic radiation over wide range of wavelengths.
         b. Could not explain the longer wavelengths of light.
         c. Max Planck suggested that the energy was emitted in small, definite amounts called a quantum.
         d. A quantum is the minimum quantity of energy that is lost or gained by an atom. A small packet of energy.
         e. The relationship between energy and frequency was developed by Max Planck $E = h\,\nu$ , where $h$ is Planck's constant $6.6262 \times 10^{-34}$ Js.
      2. Einstein proposed duality of light, exhibits both wave and particle properties. Each particle of light carries a quantum of energy. These quanta are called photons.
         a. Einstein explained the photoelectric effect by stating that the EM radiation is absorbed in whole numbers of photons, like climbing a ladder. In order for electrons to be ejected from a metal, the photons striking the electrons must possess the minimum amount of energy to kick out electron. $E_{photon} = hv$, this corresponds to min. frequency of the light. **<u>The Photoelectric Effect</u>** – electrons are ejected from the surface of certain metals exposed to a minimum frequency of light called the threshold frequency.
         b. If frequency is greater than threshold frequency then the electrons gain the excess energy. $\mathbf{E = h\,\nu = KE + BE}$ where KE is the kinetic energy of the ejected electron and BE is the binding energy of the electron ion the metal. Rearrange the equation and $\mathbf{KE = h\,\nu - BE}$ thus the more energy the photon has the greater the kinetic energy of the ejected electron.

i. Use – solar cells on calculators

D. Bohr's Theory
1. Emission spectrums
   a. When elements are sufficiently charged they also emit light but not all the wavelengths of light.
   b. When the electrons have absorbed enough energy they jump into a higher E-level and are said to excited. In order for the electrons to fall back to their ground state they emit the energy as photons of light. (photons are light quanta)
   c. The amount of energy absorbed directly corresponds to a specific wavelength of light which corresponds to definite frequency. By splitting the light emitted into its components we can determine the wavelengths / frequencies and thereby determine how much energy was given off and thus know where the electron came from and its arrangement.
2. This is the basis for the Bohr model of the atom. Good but only works for hydrogen. Electrons exist in orbits of **specific** energy. This energy is calculated using

$$E_n = -R_H(1/n^2)$$

$R_H$ (Rydberg consant) = $2.18 \times 10^{-18}$ J n = principal quantum number,1,2,3 …
3. This same idea can be used to explain how things glow in the dark. Matter (ground state) absorbs photons from light , get excited, jump to higher energy level (excited state), go back to ground state by giving off photons of light.
4. By looking at the bright line spectrum we can determine the energy level jumps by corresponding them with the Balmer series(pg 278) and use Bohr's theory to determine the energy. (See derivation page 277).

$$\Delta E = \boldsymbol{h\nu} = R_H(1/n_i^2 - 1/n_f^2)$$

E. De Broglie's Argument
1. De Broglie related the wave and particle properties using

$$\lambda = h / mu$$

where $\lambda$ = wavelength of particle (m); $h$ = Planck's Constant ($6.68 \times 10^{-34}$ Js); $m$ = mass of particle (kg) and $u$ = velocity of particle (m/s).

F. Quantum Mechanics
1. Heisenberg Uncertainty Principle
   a. Electrons are detected by their interaction with photons and since photons have about the same energy as electrons, the photons tend to move the electrons from its original location.
   b. His principle states that it is impossible to determine simultaneously both the position and velocity of an electron or any other particle.
   c. Schrodinger treated electrons as wave-particles and by combining with Heisenberg's uncertainty principle, quantum theory (mathematical description of the wave properties of electrons and other very small particles) was born.
2. Brief review of atomic structure

a. Electrons exist in energy levels (regions where electrons of the same energy exist)
b. As distance from nucleus increases the energy in the energy level increases
c. Designated by letters K,L,M,N. . .then the word shell
d. Within these levels are orbitals (highly probable location where electrons might exist

3. The energy state of an electron in an atom may be described by a set of four "numbers" called quantum numbers
4. Quantum numbers describe the space orbital the electron occupies in terms of distance from the nucleus, shape, orientation about the three axes in space, and the direction of the electron spin
5. Principal quantum number, symbolized by *n*
    a. Indicated the most probable distance of the electron from the nucleus of the atom
    b. Positive whole number (1,2,3. . .)
    c. Corresponds directly with energy level designation, or identifying number, of an orbital
    d. The principal quantum number of 1 corresponds to the first energy level, or the K shell, which is the closest to the nucleus and has the lowest energy. All others are at increasing distance and energy
6. Angular momentum quantum number, symbolized by $\ell$ ($\ell = n-1$)
    a. Indicated the shape of the orbital the electron occupies
    b. Shapes are designated by s, p, d, f, g, h
    c. s has the lowest energy, h has the highest
    d. The number of different possible shapes in an energy level is equal to the value of the principal quantum number
    e. n = 1 then 1 shape, s; n = 2 then 2 shapes, s and p; n = 3 then 3 shapes, s, p, d and so on
    f. Website for orbital pictures [http://micro.magnet.fsu.edu/electromag/java/atomicorbitals/index.html](http://micro.magnet.fsu.edu/electromag/java/atomicorbitals/index.html)
7. Magnetic quantum number, symbolized by $m_\ell$
    a. There are $(2\ell + 1)$ values of $m_\ell$ which equal $-\ell, (-\ell + 1), 0, (+\ell - 1), +\ell$.
    b. Thus if $\ell = 2$ then $m_\ell = -2, -1, 0, 1, 2$
    c. Indicated the orientation of the orbitals about the three axes in space
    d. s orbitals have one orientation $\ell = 0$
    e. p orbitals have three orientations $\ell = 1$
    f. d orbitals have five orientations $\ell = 2$
    g. f orbitals have seven orientations $\ell = 3$
    h. The first (K) energy level has one s orbital
    i. The second (L) energy level has one s and three p's for a total of four
    j. The third (M) energy level has one s, three p's and five d's for a total of 9
8. Spin quantum number, symbolized by $m_s$
    a. Indicates the electron's spin; electrons like to be in pairs
    b. Referred to as clockwise or counterclockwise ($+\frac{1}{2}$ or $-\frac{1}{2}$)

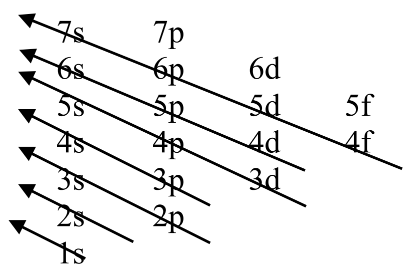
c. No two electrons in the same energy level, the same orbital and the same orientation can have the same spin

9. Relationships
   a. The total number of orbitals in an energy level is equal to $n^2$
   b. The maximum number of electrons that may occupy an energy level can be determined by $2n^2$
   c. To determine the total number of electrons an orbital can hold equals $2(2\ell + 1)$

G. Putting it all together in one notation called orbital notation
1. This notation applies all four quantum numbers and describes energy and distance
2. Must apply the Aufbau principle - electrons will enter the lowest energy orbital first
3. Must apply Hund's rule - electrons fill orbital evenly before pairing
4. Must apply Pauli exclusion principle - paired electrons must have opposite spins; no two electrons may have the exact same set of four quantum numbers
5. Diamagnetic vs Paramagnetic –
   a. Paramagnetic have single electrons unpaired are attracted by magnet
   b. Diamagnetic have all electrons paired are repelled by magnet (dissed)
6. Practice diagram - to remember placement of orbitals and the order in which the electron fill use the "magic D"

II.
III.
IV.
V.
VI.

a.




a. Fill diagonally from bottom right to upper left

Other notations

2. **Electron configuration notation** - shorthand method from orbital notation
   a. Uses the principal, orbital and number of electrons
   b. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 4f^{14}$
   c. Keep energy levels together
   d. Noble gas notation – shorthand of electron configuration notation.
      i. use the closest noble gas as shorthand
      ii. Example – Na – $1s^2, 2s^2, 2p^6, 3s^1$ would become [Ne] $3s^1$
3. **Electron dot notation** - only represents the outer shell, valence, electrons
   a. Each atom can only have eight valence electrons, these are the s and p electrons
   b. Electrons are placed around the element symbol with the two s electrons to the left, 3rd electron top, 4th electron left,

5th electron below, 6th electron top, 7th left, 8th bottom.
See diagram below

c. All electron notations represent the lowest possible energy state, ground state of the atoms
d. Ionic notations are done the same except they have a charge

| | | |
|---|---|---|
| | 3 6 | |
| 4<br>7 | X | 1<br>2 |
| | 5 8 | |

VII.

## EXCEPTIONS TO THE RULES