Chem E-1a
Friday Review Notes
Chapter 7: The Electronic Structure of Atoms

Energy, Particles, and Waves:

•Electromagnetic Radiation: (e.g. light, microwaves, x-rays, infrared radiation, etc.)

$$\mathbf{c = \lambda\nu}$$

$$\mathbf{E = h\nu = \frac{hc}{\lambda}}$$

c = speed of light = 3.00 x 10^8 m/s
λ = "lambda" = wavelength of radiation in meters
ν = "nu" = frequency of radiation in sec^{-1} or Hertz (Hz)
E = Energy of one photon of radiation in Joules
h = Planck's constant = 6.63 x 10^{-34} J•s

•Particles: (anything that has mass, such as electrons, neutrons, baseballs, etc.)

$$\mathbf{KE = \frac{1}{2}mu^2}$$

$$\mathbf{\lambda = \frac{h}{mu}}$$

KE = kinetic energy of particle in Joules
m = mass of particle in kg
u = velocity of particle in m/s
λ = "lambda" = DeBroglie Wavelength of particle in meters
h = Planck's Constant = 6.63 x 10^{-34} J•s

•Photoelectric Effect:

$$\mathbf{E_{photon} = B.E. + KE_{electron}}$$

E_{photon} = Energy of one photon of incoming light or radiation in Joules = $h\nu$ = $(hc)/\lambda$
B.E. = Binding Energy of metal = work function of metal = $h\nu_0$
= the minimum amount of energy required to eject a single electron from a single atom
(Note: ν_0 = threshold frequency for photoelectric effect)
$KE_{electron}$ = Kinetic Energy of emitted electron = $\frac{1}{2}mu^2$

This equation can also be written as: $h\nu = h\nu_0 + \frac{1}{2}mu^2$

•Energy of an Electron in a Hydrogen Atom:

$$\mathbf{E}_n = -\mathbf{R}_H\left(\frac{1}{n^2}\right)$$

$$\Delta\mathbf{E} = \mathbf{R}_H\left(\frac{1}{n_i^{\,2}} - \frac{1}{n_f^{\,2}}\right)$$

E_n = Energy of an electron in a hydrogen atom with principal quantum number n
ΔE = Energy change when an electron in a hydrogen atom undergoes a transition from one principal energy level to another. This energy change will correspond to the energy of the photon that is absorbed or emitted during this process.
R_H = Rydberg's Constant = 2.18×10^{-18} J
n_i = initial principal quantum number of the electron in a hydrogen atom
n_f = final principal quantum number of the electron in a hydrogen atom

Note: These formulas can be adapted to apply to any ion with only one electron, such as the He^+ ion or the Li^{2+} ion, as follows:

$$E_n = -R_H\left(\frac{Z^2}{n^2}\right)$$

$$\Delta E = R_H\left(\frac{Z^2}{n_i^{\,2}} - \frac{Z^2}{n_f^{\,2}}\right)$$

where Z = charge of the nucleus = atomic number
For the hydrogen atom, Z = 1, and these formulas simplify to the formulas above.

Note: In all of this material, DO NOT confuse frequency (ν) of a light wave and velocity or speed (u) of an electron. Make sure there is a distinction in how you write these symbols.
Similarly, don't confuse wavelength (λ) with Planck's constant (h). Again, make sure you write these symbols in a distinct fashion.

Units: J = kg•m^2/s^2 (so in Planck's constant: J•s = kg•m^2/s)
Hz = sec^{-1}
nm = nanometers = 10^{-9} m
Angstrom = 10^{-10} m

Quantum Numbers:

n = Principal Quantum Number = 1, 2, 3, …

n indicates the general size of an orbital. It corresponds to the principal energy level of an orbital. For example, a 4*s* orbital has $n = 4$.

l = Angular Momentum Quantum Number = 0, 1, 2, … n–1

l indicates the shape of an orbital. It is indicated by the letter designation of an orbital:

$l = 0$ for an *s* orbital
$l = 1$ for a *p* orbital
$l = 2$ for a *d* orbital
$l = 3$ for an *f* orbital
etc…

m_l = Magnetic Quantum Number = integer values: $-l, \ldots, 0, \ldots, l$

m_l indicates the orientation of an orbital.

m_s = Spin Quantum Number = $+\frac{1}{2}$ or $-\frac{1}{2}$

m_s indicates the "spin" of the electron in an orbital.

Drawing Orbitals and Counting Nodes:

Total Number of Nodes = $n - 1$

Number of Angular Nodes = l

Number of Radial Nodes = $n - l - 1$

•Drawing the orbitals:

Know the shapes and orientations of all *s*, *p*, and *d* orbitals. See the textbook pp. 229-230. You should be able to draw these orbitals with any number of radial nodes.

Electron Configurations of Neutral Atoms:

•If you are using the noble gas abbreviations, find the previous noble gas. (As a general rule you should always use the noble gas abbreviations unless you are explicitly told not to.)
•After the previous noble gas, simply count electrons across the final period.
•Make sure you know each "block" of the periodic table:
  •The "*s*-block", where *s* orbitals are filled
  •The "*p*-block", where *p* orbitals are filled
  •The "*d*-block", where *d* orbitals are filled
  •The "*f*-block", where *f* orbitals are filled
•Know what principal energy level begins each "block" on the periodic table:
  1*s* 2*p* 3*d* 4*f*
•You should know 2 exceptions to the standard filling order. These exist because extra stability is gained by having a subshell completely filled (all orbitals with 2 e^-) or exactly half-filled (all orbitals with 1 e^-). In each of these cases, an "s" electron has been promoted into a d-orbital so that all subshells are either half-filled or completely filled:

Chromium (Cr): $[Ar]4s^1 3d^5$
Copper (Cu): $[Ar]4s^1 3d^{10}$

There are other atoms that have "abnormal" electron configurations, but you are not required to know them.