

### I. Emission Spectra

When a substance is exposed to a certain intensity of light or some other form of energy, the atoms absorb some of the energy (**excited state** atoms).

As the atom absorbs energy, one or more electrons change their orbit(s). Since this excited state is not stable, the electrons will return to their stable **ground state**, by releasing energy in the form of light (**emission spectra**).

The emission spectra of each element is unique, since certain bands of light (at certain energies) are produced. The energy is **quantized**, since only certain energies are allowed.

### II. Electromagnetic Spectrum

#### A. Light

$$c = \lambda \nu$$

$c$  = the speed of light ( $3.00 \times 10^8$  m/s)  
 $\lambda$  (Gr. letter lambda) = wavelength (meters)  
 $\nu$  (Gr. letter nu) = frequency ( $\frac{\text{waves}}{\text{sec}} = \text{sec}^{-1} = \text{Hertz} = \text{Hz}$ )

$$** 1 \text{ nm} = 1 \times 10^{-9} \text{ m}$$

#### B. Max Planck

In 1900 he measured the heat from glowing objects and realized the energy was **quantized**. (The energy was gained/lost in only certain amounts.)

$$\Delta E = n h \nu$$

$\Delta E$  = change in energy (J/quanta)  
 $n$  = whole number of quanta,  $n = 1$  for 1 quantum  
 $h$  = Planck's Constant ( $6.626 \times 10^{-34}$  J/Hz)  
 $\nu$  (nu) = frequency (Hz)

#### C. Einstein's Photoelectric Effect

Electrons can be ejected from metals, when light shines on the metal.

Einstein found that this light energy is quantized: A certain amount of energy must be absorbed to eject each electron.

\*\*Light, therefore, all electromagnetic radiation must be quantized and consists of particles, named photons.

$$\Delta E = n h \nu$$

applies to all electromagnetic radiation,  
where **n = the number of photons**.

Ex. 1) A certain red light has a wavelength of 690 nm. What is its frequency?

$$\frac{690 \text{ nm}}{1 \text{ nm}} \left| \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} \right. = 6.90 \times 10^{-7} \text{ m}$$

$$c = \lambda \nu \quad (3.00 \times 10^8 \text{ m/s}) = (6.90 \times 10^{-7} \text{ m}) \nu$$
$$\mathbf{4.35 \times 10^{14} \text{ Hz} = \nu}$$

Ex. 2) Find the energy of one quantum of light (**one photon**), having a frequency of  $2.76 \times 10^{15}$  Hz.

$$\Delta E = nh\nu$$

$$\Delta E = (1 \text{ photon}) (6.626 \times 10^{-34} \text{ J/Hz}) (2.76 \times 10^{15} \text{ Hz})$$

$$\Delta E = \mathbf{1.83 \times 10^{-18} \text{ J/photon}}$$

$$\rightarrow \text{J/mol} : \frac{1.83 \times 10^{-18} \text{ J}}{\text{photon}} \left| \frac{6.022 \times 10^{23} \text{ photons}}{1 \text{ mol}} \right. = \mathbf{1.10 \times 10^6 \text{ J/mol}}$$

Ex. 3) It takes  $4.42 \times 10^{-19}$  J to remove an electron from an unknown atom. Is light with a wavelength of 470 nm strong enough to remove an electron? (like Einstein's experiment)

$$\frac{470 \text{ nm}}{1 \text{ nm}} \left| \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} \right. = 4.70 \times 10^{-7} \text{ m}$$

$$c = \lambda \nu \quad (3.00 \times 10^8 \text{ m/s}) = (4.70 \times 10^{-7} \text{ m}) \nu$$
$$\mathbf{6.38 \times 10^{14} \text{ Hz} = \nu}$$

$$\Delta E = nh\nu = (1 \text{ photon}) (6.626 \times 10^{-34} \text{ J/Hz}) (6.38 \times 10^{14} \text{ Hz})$$

$$\Delta E = \mathbf{4.23 \times 10^{-19} \text{ J/photon}}$$

$$\mathbf{4.23 \times 10^{-19} \text{ J/photon}} \quad < \mathbf{4.42 \times 10^{-19} \text{ J}}$$

(smaller than the needed amount)

**This light is not strong enough.**

$$\mathbf{** \text{ if kJ/mol change to J/photon} \quad 1 \text{ kJ} = 1 \times 10^3 \text{ J}}$$
$$1 \text{ mol} = 6.022 \times 10^{23} \text{ photons}$$

#53 Notes D. De Broglie's Equation for Particles

He hypothesized that since light consisting of photons moves in waves, other small particles must also move in waves.

Einstein:  $E = mc^2$

Planck:  $E = h\nu$

$c = \lambda\nu$

Therefore

$$\begin{array}{ccc}
 E = E & & \\
 mc^2 = h\nu & & c/\lambda = \nu \\
 \uparrow & \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow \leftarrow & \downarrow \\
 mc^2 = h(c/\lambda) & & 
 \end{array}$$

$$mc = h/\lambda$$

$$\lambda mc = h$$

$$\lambda = \frac{h}{mc} \quad \rightarrow \quad \lambda = \frac{h}{mv} \quad \begin{array}{l} m = \text{mass (kg)} \\ v = \text{velocity (m/s)} \end{array} \quad \textbf{**for particles}$$

- mass of electron =  $9.109 \times 10^{-31}$  kg
- mass of proton =  $1.673 \times 10^{-27}$  kg
- mass of neutron =  $1.675 \times 10^{-27}$  kg

Ex. 1) What is the wavelength of a neutron traveling at  $3.14 \times 10^6$  m/s?

$$\lambda = \frac{h}{mv} \quad \lambda = \frac{(6.626 \times 10^{-34} \text{ J /Hz})}{(1.675 \times 10^{-27} \text{ kg})(3.14 \times 10^6 \text{ m/s})}$$

$$\lambda = \mathbf{1.26 \times 10^{-13} \text{ m}}$$

\*\*If a problem states: 4% the speed of light = (0.04) ( $3.00 \times 10^8$  m/s)

Ex. 2) What is the speed of an electron traveling with a wavelength of 432 nm?

$$\frac{432 \text{ nm}}{1 \text{ nm}} \left| \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} \right. = 4.32 \times 10^{-7} \text{ m}$$

$$\lambda = \frac{h}{mv} \quad (4.32 \times 10^{-7} \text{ m}) = \frac{(6.626 \times 10^{-34} \text{ J /Hz})}{(9.109 \times 10^{-31} \text{ kg}) v}$$

$$(3.935 \times 10^{-37}) v = (6.626 \times 10^{-34} \text{ J /Hz})$$

$$\mathbf{v = 1.68 \times 10^3 \text{ m/s}}$$

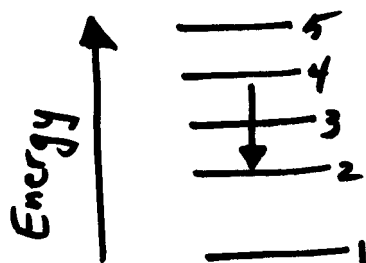
### III. Bohr Model

Bohr figured out an equation to calculate the energy given off as an electron in a Hydrogen atom moves between different energy levels. (Only certain orbits/layers are allowed, so it is quantized.)

$$\Delta E = -2.178 \times 10^{-18} \text{ J} \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right) \quad n = \text{layer number}$$

Ex. 1) Calculate the energy for  $n=4 \rightarrow n=2$ . Then calculate the wavelength.

Show the transition:



$$\Delta E = -2.178 \times 10^{-18} \text{ J} \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)$$

$$\Delta E = -2.178 \times 10^{-18} \text{ J} \left( \frac{1}{2^2} - \frac{1}{4^2} \right)$$

$$\Delta E = -2.178 \times 10^{-18} \text{ J} \left( \frac{1}{4} - \frac{1}{16} \right)$$

$$\Delta E = -2.178 \times 10^{-18} \text{ J} (0.25 - 0.0625)$$

$$\Delta E = -2.178 \times 10^{-18} \text{ J} (0.1875)$$

$$\Delta E = -4.08 \times 10^{-19} \text{ J of the light} \quad **\text{now find the wavelength of the light}$$

$$\Delta E = nh\nu \quad 4.08 \times 10^{-19} \text{ J} = (1) (6.626 \times 10^{-34} \text{ J/Hz}) (\nu)$$
$$6.16 \times 10^{14} \text{ Hz} = \nu$$

$$c = \lambda\nu \quad 3.00 \times 10^8 \text{ m/s} = \lambda (6.16 \times 10^{14} \text{ Hz})$$
$$4.87 \times 10^{-7} \text{ m} = \lambda \quad \rightarrow 487 \text{ nm}$$

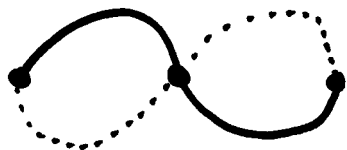
(blue-violet on spectrum chart)

## #54 Notes IV. Schrodinger's Quantum Mechanics

Unfortunately Bohr's Model only worked accurately on Hydrogen.

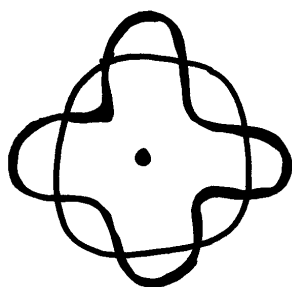
Schrodinger, like De Broglie, believed that the electron showed wave-like features.

The electron moves around the nucleus in a motion resembling a standing wave (waves that do not travel).



Ex. Guitar string, Tacoma Narrows Bridge Collapse,

Around the atom:



$n = 4$

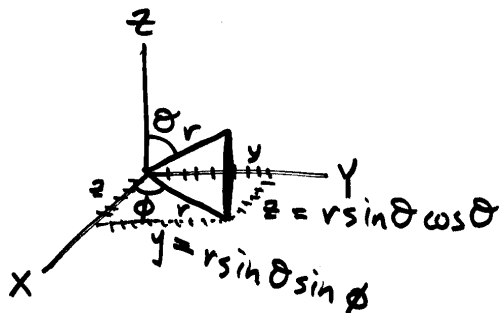
Schrodinger created an equation to describe this.

$$\hat{H} \Psi = E \Psi$$

$\hat{H}$  = mathematical operator,  $E$  = the energy operator  
 $\Psi$  = wave function describing the electrons orbit

For Hydrogen:

$$\frac{\partial^2 \Psi}{\partial x^2} + \frac{\partial^2 \Psi}{\partial y^2} + \frac{\partial^2 \Psi}{\partial z^2} + \frac{2\mu}{h^2} \left( E + \frac{ze^2}{r} \right) \Psi = 0$$



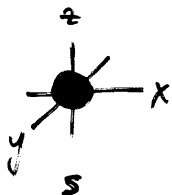
$$\frac{1}{r^2} \frac{\partial}{\partial r} \left( r^2 \frac{\partial \Psi}{\partial r} \right) + \frac{1}{r^2 \sin^2 \theta} \frac{\partial^2 \Psi}{\partial \phi^2} + \frac{1}{r^2 \sin \theta} \frac{\partial}{\partial \theta} \left( \sin \theta \frac{\partial \Psi}{\partial \theta} \right) + 2 \left( E + \frac{z}{r} \right) \Psi = 0$$

\*\*From this we know where the electron is, but not it's path.

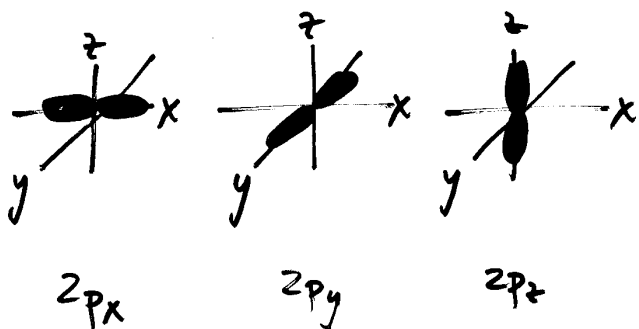
### A) Electron Orbitals

By using the square of the wave function, a probability map can be drawn to show where the electrons are located at specific times.

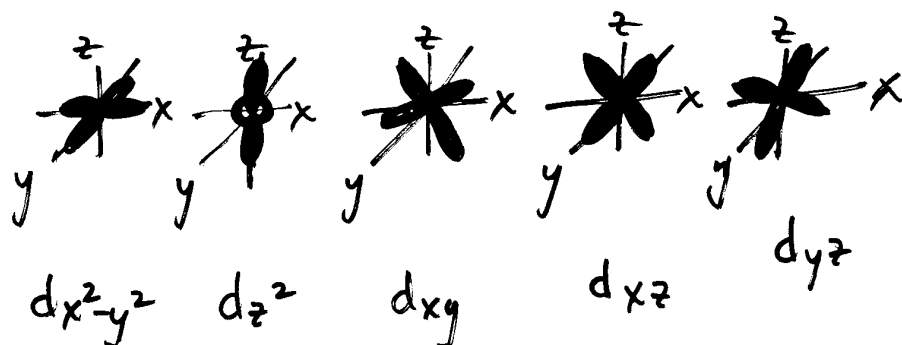
#### I. s – orbitals (1 type)



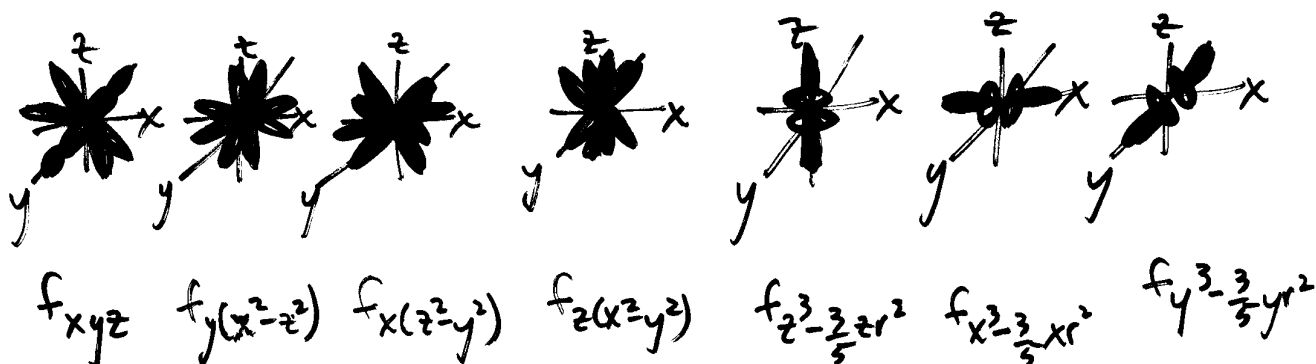
#### II. p – orbitals (3 types)



#### III. d – orbitals (5 types)



#### IV. f – orbitals (7 types)



B) Energy Levels: are the layers of electrons (1, 2, 3, 4, etc.)

C) Sublevels: are the energy levels and orbital (2s, 3p, etc.)

### V. Electron Configurations

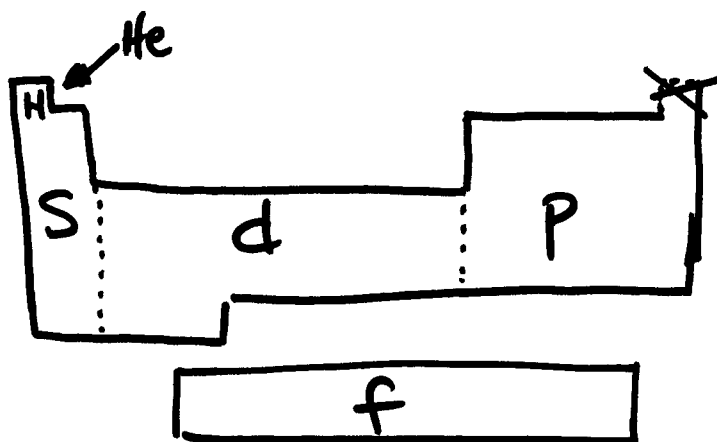
Ex. 1a) Mg:  $1s^2 2s^2 2p^6 3s^2$

↑↑↑

1 = 1<sup>st</sup> layer (row #), s = orbital type, power of 2 = the 2 electrons in the 1s orbital

1b) Br:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

1c) Ba:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$



s-orbitals: column #1-2

d-orbitals: column #3-12

p-orbitals: column #13-18

f-orbitals: elements #57-71 and #89-103

#55 Notes VI. Orbital Diagrams

A) Pauli Exclusion Principle

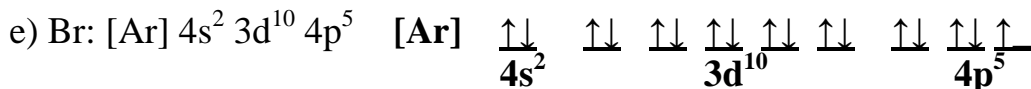
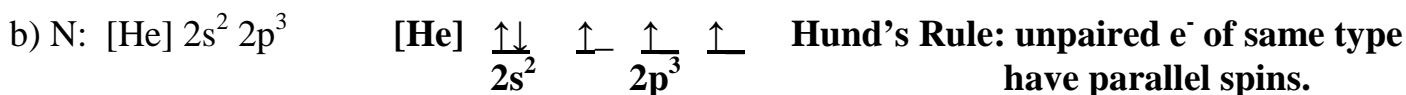
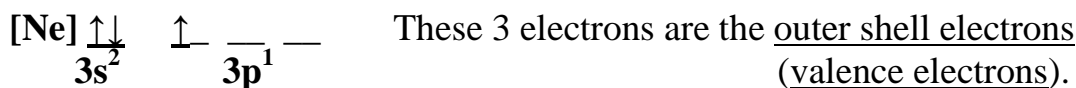
An orbital can only hold 2 electrons and they must have opposite spins.

B) Hund's Rule

The lowest energy configuration (most stable) will have the maximum number of unpaired electrons allowed by the Pauli Principle.

Ex. 1) Write the orbital diagram for:

a) Al:  $1s^2 2s^2 2p^6 3s^2 3p^1$  or  $[\text{Ne}] 3s^2 3p^1$  (short form: start at noble gas before Al)

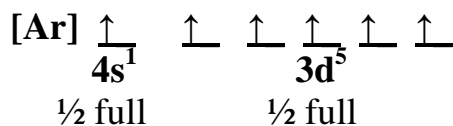
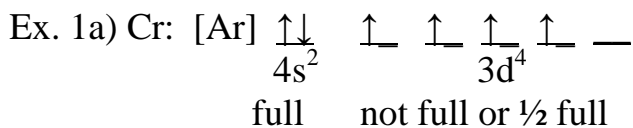


VII. Exceptions

Note in textbook, there are many exceptions in the d, f orbitals, since their energies are very similar.

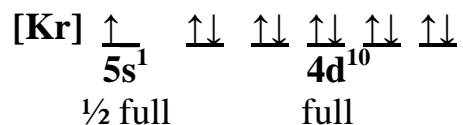
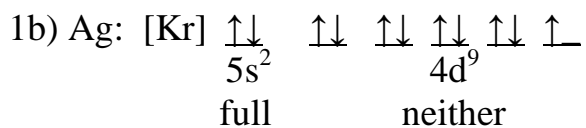
A) Some exceptions can be predicted.

\*\*There is special stability for full and 1/2 full orbitals.



\*\* "s" and "d" orbitals are similar in energy, so they can shift an electron from "s" to "d".  
(but not "s" to "p", stop)

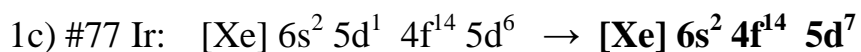
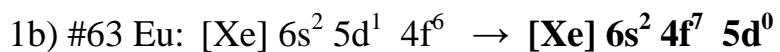
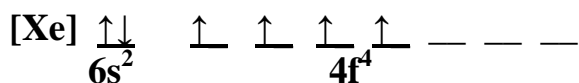
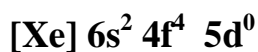




B) f-orbitals:

f-orbitals usually fill before d-orbitals.

Ex. 1a) #60 Nd: [Xe]  $6s^2 5d^1 4f^3$  \*\*The 1e- alone in the “d” is lonely and moves to the “f” with its friends.



## #56 Notes XI. Periodicity

Since an element's position on the periodic table describes its electron configuration, its position allows predictions about certain properties of the element.

### A) Atomic Radius (Size)

a) As the number of orbital layers/levels increase (going down a column), the radius will increase.

b) Going to the right across a row, each atom has an extra proton. This will attract the electrons more tightly, decreasing the radius.

c) Anions (-) have gained electrons. Since it will be harder to attract the extra electrons, anions will be larger.

d) Cations (+) have lost electrons. Since it has fewer electrons to attract, its (+) nuclear **charge** can pull the electrons in closer, reducing the radius.

Ex. 1) Which in each pair has a smaller radius?

a) S, Se      b) Ta, Ir      c) Mg, Mg<sup>2+</sup>      d) F, F<sup>-1</sup>      e) Te, Cr

**a) S has fewer layers of electrons**

**b) Ir has more protons to pull in the electrons**

**c) Mg<sup>2+</sup> has fewer electrons to attract**

**d) F has fewer electrons to attract**

**e) Cr has fewer layers of electrons \*\*\*(layers is a larger effect than within a row)**

### B. Ionization Energy

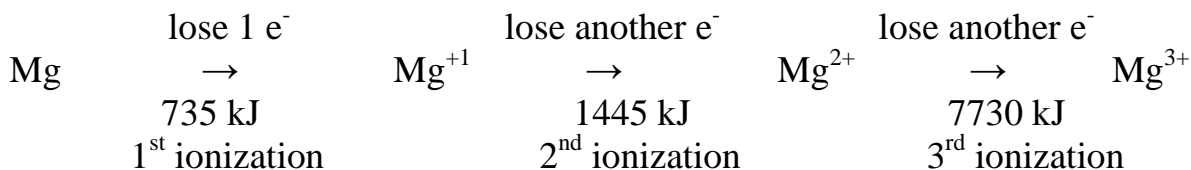
-is the energy to remove an electron.

a) Generally, the smaller the radius (the electrons are more highly attracted to the nucleus), the harder it is to pull away an electron, the larger the ionization energy.

The larger the radius, the more the inner electrons **shield** the outer electrons from the attraction of the nucleus, making it easier to remove them (small ionization energy).

**{Shielding Effect}**

b) The larger the (+) **nuclear charge**, the higher the ionization energy. It takes more energy to remove each successive electron.



3<sup>rd</sup> ionization is extremely high, since it removes an electron from a stable noble gas configuration.

c) The ionization energy is unusually high to remove an electron from a stable full or  $\frac{1}{2}$  full orbital.

## #57 Notes C. Electron Affinity

-is the attraction energy of an atom for another electron.

The smaller the radius, the stronger the attraction of the electrons, the more it wants to attract more electrons, so the higher the electron affinity.

The larger the radius, the weaker the attraction of electrons, so the smaller the electron affinity.

Atoms that have full or  $\frac{1}{2}$  full orbitals will have unusually small electron affinities, since they are already stable and do not need to attract more electrons.

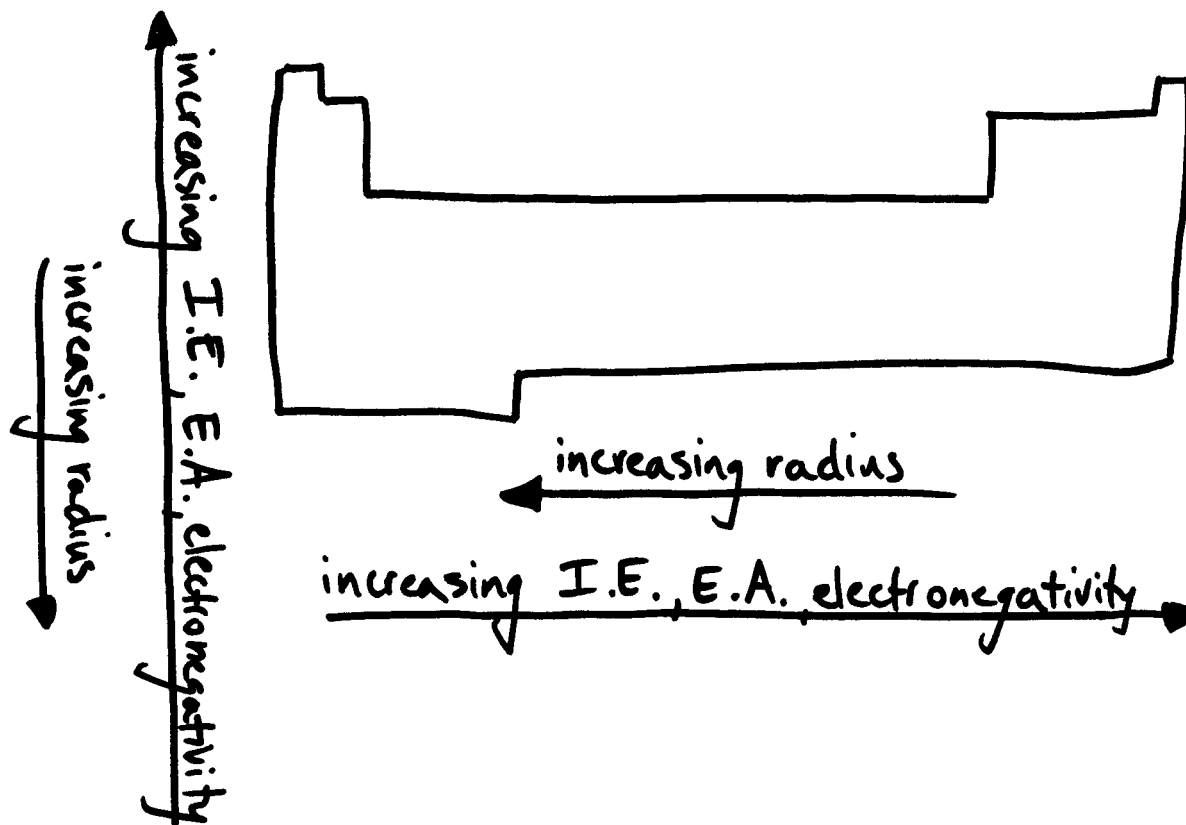
\*\*Electron affinity is (-) {exothermic}, since energy is released as electrons are gained and the atom becomes more stable.

Big (-) # = large electron affinity = highly exothermic

## D. Electronegativity

- is the ability of an atom, **in a molecule**, to attract shared electrons to it.

Small radius, strong electron attraction, large electronegativity.



Most Active Metal (Most Metallic)

(Metals lose electrons)

Loses electrons the best!

Lowest Ionization Energy

Largest Radius: Bottom/Left on Table

**Fr**

Most Active Nonmetal (Least Metallic)

(Nonmetals gain electrons)

Gains electrons the best!

Highest Electron Affinity

Smallest Radius: Top/Right (excluding noble gases)

**F**

**Ionization Energy** (should be increasing as radius decreases, to the right)

C	N	O	F
1086	1402	1313	1681

**N is  $\frac{1}{2}$  full, stable and  
doesn't want to lose any e<sup>-</sup> (very high ionization energy).**

**Electron Affinity** (should be increasing as radius decreases, to the right)

C	N	O	F
-122	0	-141	-328

**N is  $\frac{1}{2}$  full, stable and  
doesn't want to gain any e<sup>-</sup> (very low electron affinity).**

**\*End of Notes\*** (Assignments #58-59 are Review Assignments. There are no notes for these assignments.)