

Chapter 8: Periodic Relationships Among the Elements

8.1: Development of the Periodic Table

Johann Dobereiner: - first to discover a pattern of a group of elements like Cl, Br, and I (called triads).

John Newland: - suggested elements should be arranged in “octaves” because they repeat their properties for every eighth elements.

Demitri Mendeleev: - conceived the first modern periodic table of elements (independently worked on by Julius Lothar Meyer).
- insisted certain spots of the table be left blank until the actual element is found that matched the predicted properties. This was done to preserve the elements with similar properties called groups or families.

8.2: Periodic Classification of the Elements

Representative Elements: - also called **Main Group Elements**.
- elements in Group IA to VIIIA (Groups 1, 2 and 13 to 18) that consists of partial or filled s and p orbitals of the same principal quantum number, n .
- other than helium, $1s^2$, all noble gases have ns^2np^6 as their electron configurations.

Transition Metals: - elements in Group 1B and 3B through 8B (Groups 3 to 11).
- all transition metals consists of d orbital electrons in their outer electron configurations. The group designation acknowledges the number of outer electrons. For examples, Mn has 7 valence electrons ($4s^23d^5$) and it is in Group 7B.
- Group 2B (Zn, Cd and Hg) are neither transition metals nor representative elements due to the fact they have filled $ns^2(n-1)d^{10}$ as their electron configurations.

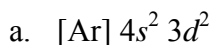
Lanthanides and Actinides: - are elements with incompletely filled f subshells.
- are sometimes called **Inner Transition Elements**.

Valence Electrons: - the outer electrons of an atom that are involved in chemical bonding.
- for representative elements, it takes 8 valence electrons (ns^2np^6) to achieve stability.
- in most cases, this means having the same electron arrangement of the **nearest** noble gas, except helium (only 2 electrons to fill the first energy level), and the transition metals.

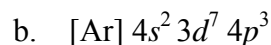
Ground State Electron Configuration: - the electron configuration of an atom at its lowest energy state.

Excited State Electron Configuration: - the electron configuration of an atom at a higher energy state.

Example 1: Identify the atom of the following electron configurations and indicate whether it is at ground state or excited state.



There are 4 valence electrons after the 3rd row of the Table of Elements (Ar). This mean the atom is at the 4th element in row 4. Hence, it is **Ti**. It is at its **ground state** because it did not indicate any electrons have gone into the $4p$ or higher orbitals than $4s$ and $3d$.



There are 12 valence electrons ($2 + 7 + 3$) after the 3rd row of the Table of Elements (Ar). This mean the atom is at the 12th element in row 4. Hence, it is **Zn**. It is at its **excited state** since there are three electrons in the $4p$ when the $3d$ subshell is not filled.

Electron Configurations of Ions from the Representative Elements:

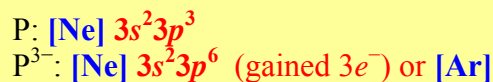
- **Group 1A and 2A metals, which like to lose electrons to form cations**, they have the same electron configurations as the **noble gas of the previous row**.
- for representative elements **non-metals, which like to gain electrons to form anions**, they have the same electron configurations as the **noble gas at the end of the same row**.

Example 2: Write the electron configurations of the following.

a. Mg and Mg^{2+}



b. P and P^{3-}

Electron Configurations of Ions from Transition Metals:

- because the proximity between the energy levels of ns and $(n-1)d$ orbitals, the interactions between the electrons and nucleus of transition metals ions are quite different than their parent atoms. Hence the **transition metal atom tends to lose the ns electrons before the electrons in the $(n-1)d$ orbitals**.

Example: Fe atom: $[\text{Ar}] 4s^2 3d^6$, Fe^{2+} ion: $[\text{Ar}] 3d^6$, Fe^{3+} ion: $[\text{Ar}] 3d^5$

Assignment

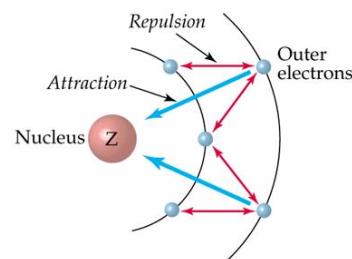
8.1 pg. 348 #1 to 4

8.2 pg. 348–349 #11, 13, 15 to 20, 22, 23, 24 to 32 (even)

8.3: Periodic Variation in Physical Properties

There are many different trends regarding the physical and chemical properties of the elements in the Periodic Table. However, we will limit to three atomic properties. They are atomic size, ionic radii, ionization energy, and electron affinity.

Shielding Effect: - the outer electrons are pushed away because of the repulsion between them and the core electrons. The net result is that the protons in the nucleus cannot hold on to these outer electrons as tightly as they would for the core electrons.



Effective Nuclear Charge (Z_{eff}): - the net nuclear charge actually experienced by an electron (the difference between the number of protons, Z , and the number of “shielded” core electrons).
- **the higher it is for Z_{eff} , the less shielding effect the outer electrons will experience.**

$$Z_{eff} = Z - \text{“Shield” Core Electrons}$$

Example 1: Calculate the effective nuclear charge of Na and Ar (first and last elements of period 3).

Sodium (Na): $Z = 11$ protons
 $Z_{eff} = 11 - 10$

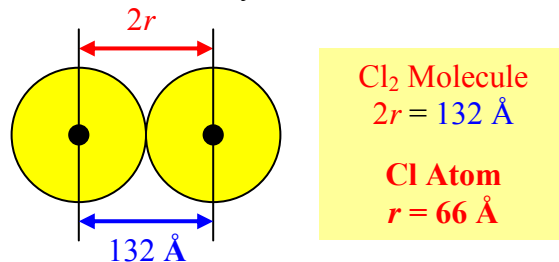
“Shield” Core $e^- = 10$ (e^- in the first two shells)
 $Z_{eff} = 1$ for Na

Argon (Ar): $Z = 18$ protons
 $Z_{eff} = 18 - 10$

“Shield” Core $e^- = 10$ (e^- in the first two shells)
 $Z_{eff} = 8$ for Ar

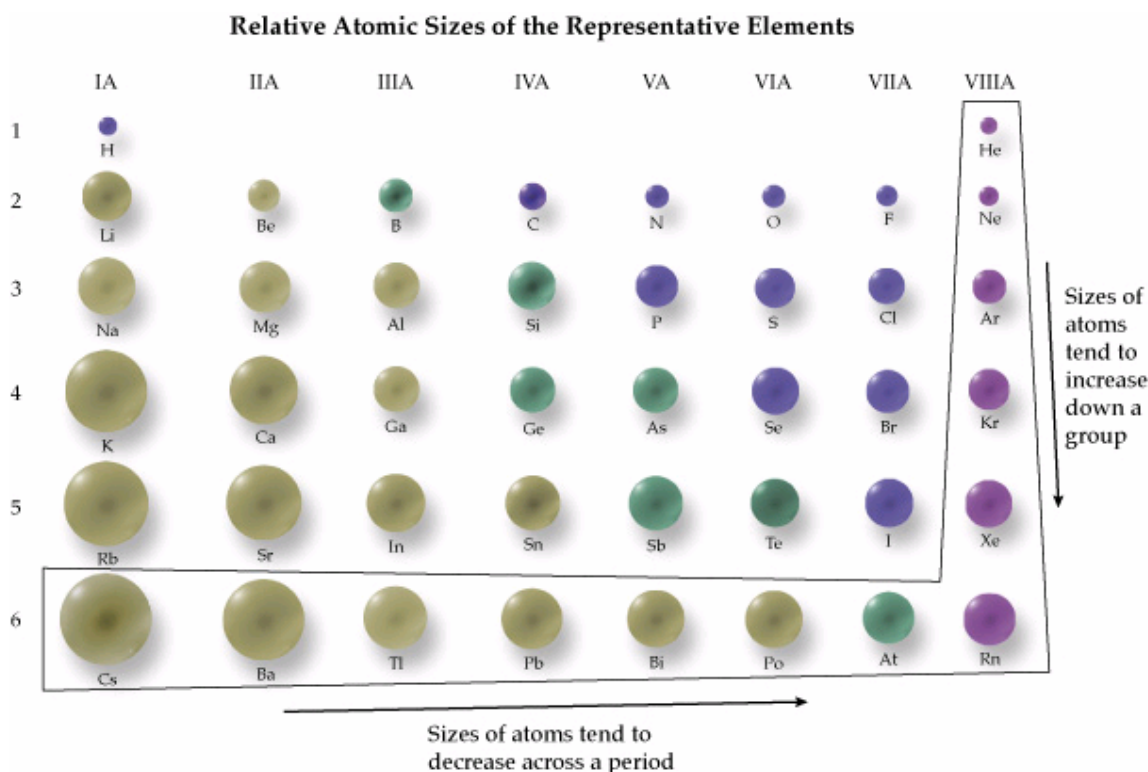
(Ar has experience LESS shielding effect than Na)

Atomic Radius: - the size of an atom as measured by the distances between atoms in chemical compound.



Several Notes on Trends in Atomic Radii

- In general, **Atomic Radii decrease as one move to the right of a period**. This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus drawing these outer electrons closer to the nucleus, decreasing in sizes as the result.
- Atomic Radii INCREASES Down a Group**. This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus



Example 1: Order the following atoms from the smallest to the largest.

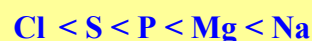
a. Te, S, Se, O

b. Na, S, Mg, Cl, P

These atoms are within the same Group (column). As we move down the column, atomic size increases. Therefore,



These atoms are within the same Period (row). As we move to the right, atomic size decreases. Therefore,

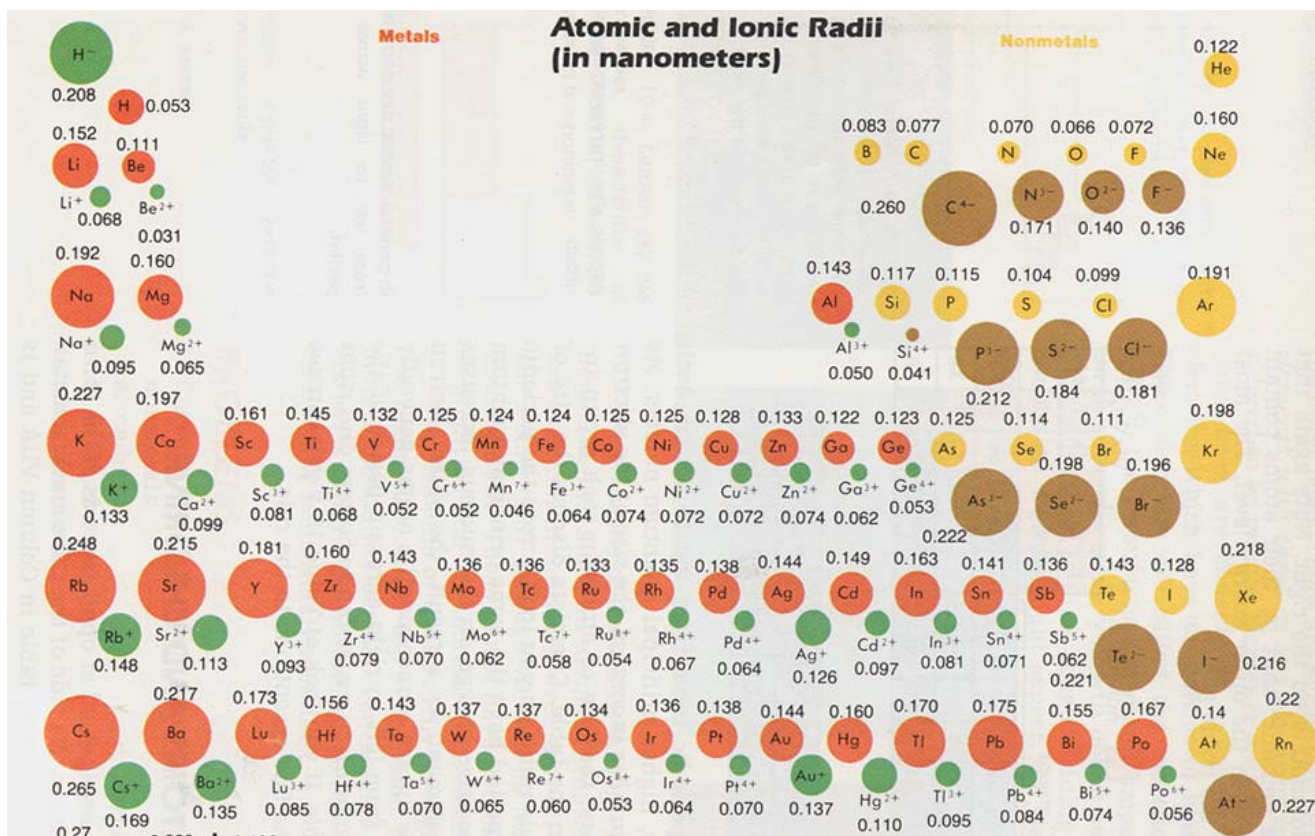


Ions Sizes: - the size of cations and anions.

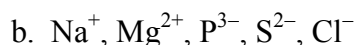
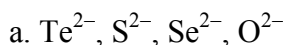
Isoelectronic Ions: - ions that have the same number of electrons. “iso” means same.

Several Notes on Trends in Ion Size

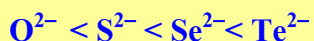
- 1. Metal Cations are generally Smaller than Non-Metal Anions WITHIN the Same Period.** Metal cations tend to lose electrons to achieve the electron configuration of a previous noble gas. Thereby, losing an energy level in the process. **Metal Cations are always Smaller than the Parent Neutral Atoms. Non-Metal Anions are always Larger than the Parent Neutral Atoms.**
- 2. In general, Ion Sizes Decrease as one move from LEFT to RIGHT of a period WITHIN the METAL GROUPS and WITHIN the NON-METAL GROUPS.** This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus drawing these outer electrons closer to the nucleus, decreasing in sizes as the result.
- 3. Ionic Radii INCREASES Down a Group.** This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus
- 4. When Comparing Sizes of Isoelectronic Ions, the rule is that the Ion with the MOST PROTONS is the SMALLEST Ions.** The increase amount of protons have pull the electrons inward, decreasing the general size.



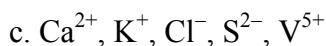
Example 2: Order the following ions from the smallest to the largest.



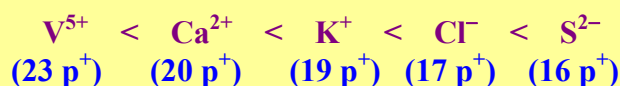
These anions are within the same Group (column). As we move down the column, ion size increases. Therefore,



These ions are within the same Period (row). As we move to the right, ion size decreases within each of the metal and non-metal groups. Therefore,



These ions are isoelectronic (all have $18e^-$). Since they have the same number of core and valence electrons in their various energy levels, the ions with the largest atomic number (number of protons) exerts more "pull" on these electrons. This in turn would make the size of the ion smaller.



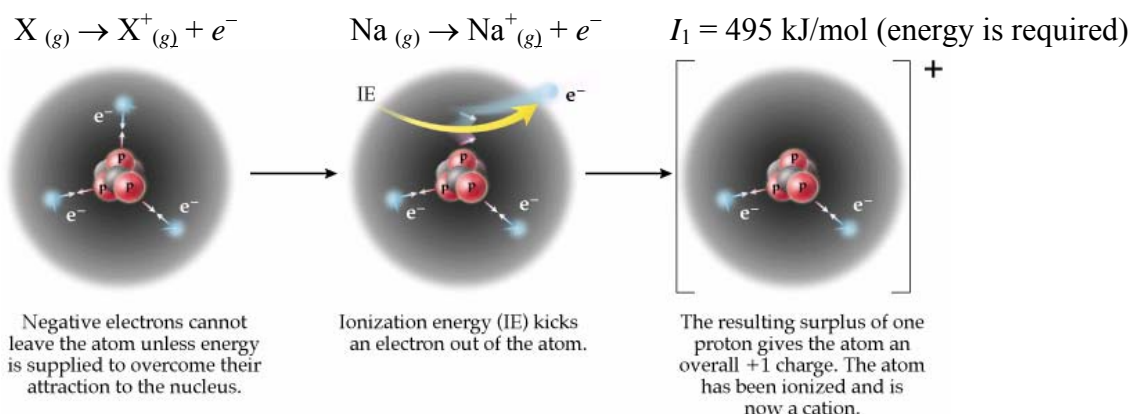
Assignment

8.3 pg. 348 #34 to 44 (even), 45 to 48

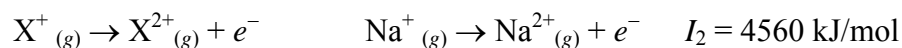
8.4 & 8.5: Ionization Energy & Electron Affinity

Ionization Energy: - the energy needed to completely remove an electron from a gaseous atom or gaseous ion (plasma).

First Ionization Energy: - the ionization energy required to remove the highest-energy electron from an atom.



Second Ionization Energy: - the ionization energy required to remove the second highest-energy electron from the ion.



Successive Ionization Energies (kJ/mol) for Elements in Row 3 of the Periodic Table

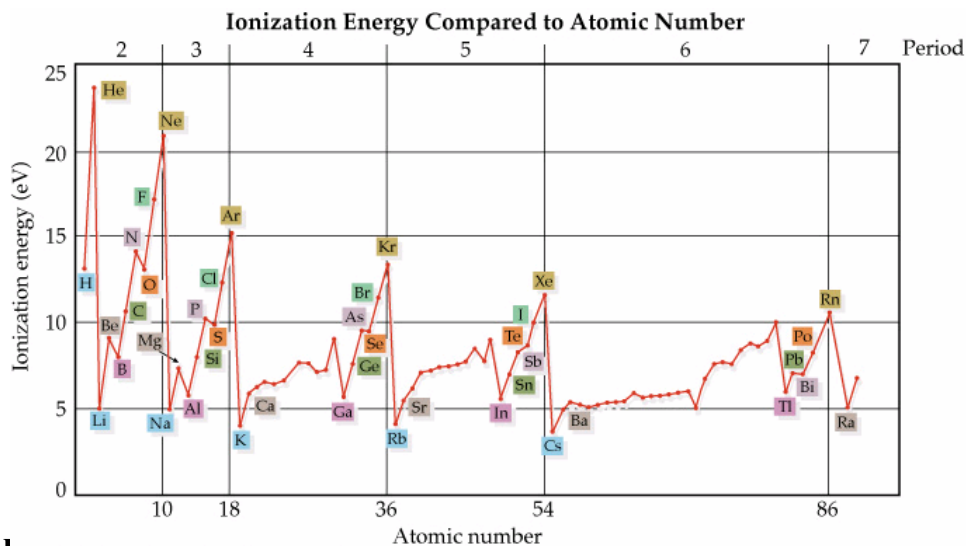
First Ionization Energies generally INCREASE within a Period.
 $(I_{\text{Na}} < I_{\text{Mg}} > I_{\text{Al}} < I_{\text{Si}} < I_{\text{P}} > I_{\text{S}} < I_{\text{Cl}} < I_{\text{Ar}})$

Elements	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4560					
Mg	735	1445	7730			Core Electrons	
Al	580	1815	2740	11600			
Si	780	1575	3220	4350	16100		
P	1060	1890	2905	4950	6270	21200	
S	1005	2260	3375	4565	6950	8490	27000
Cl	1255	2295	3850	5160	6560	9360	11000
Ar	1527	2665	3945	5770	7230	8780	12000

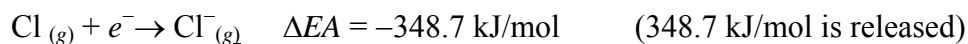
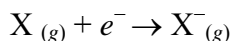
Successive Ionization Energies INCREASES within each element. ($I_1 < I_2 < I_3 < \dots$)

Several Notes on Trends in Ionization Energies

1. There is an **Increase in Successive Ionization Energies** because each successive electron has to jump from a lower level. Besides, these successive electrons are bind more tightly with the nucleus because they are closer to the protons.
2. **Ionization Energies Decrease Down a Group.** This is due to the fact as the atom has more orbitals, it is increasing in size. It is easier (takes less energy) to take away a valence electron because the protons are having a more difficult time to “hold on” to the electron.
3. In general, **Ionization Energies Increase as one move from Left to the Right of a Period.** This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus requiring more energy to ionize them.
4. The notable exceptions to the above statement is between Mg and Al along with P and S. In the case between magnesium and aluminum, the I_1 decreases because of the change in subshells ($[\text{Ne}] 3s^2$ and $[\text{Ne}] 3s^2 3p^1$). It is easier to remove a $3p^1$ electron than an electron in $3s^2$. Between P and S, the electron configuration changes between $[\text{Ne}] 3s^2 3p_x^1 3p_y^1 3p_z^1$ to $[\text{Ne}] 3s^2 3p_x^2 3p_y^1 3p_z^1$. Again, it is easier to remove an electron from a $3p_x^2$ orbital because of Hund's Rule (the lowest energy state is achieved when all electrons in the subshell is unpaired).

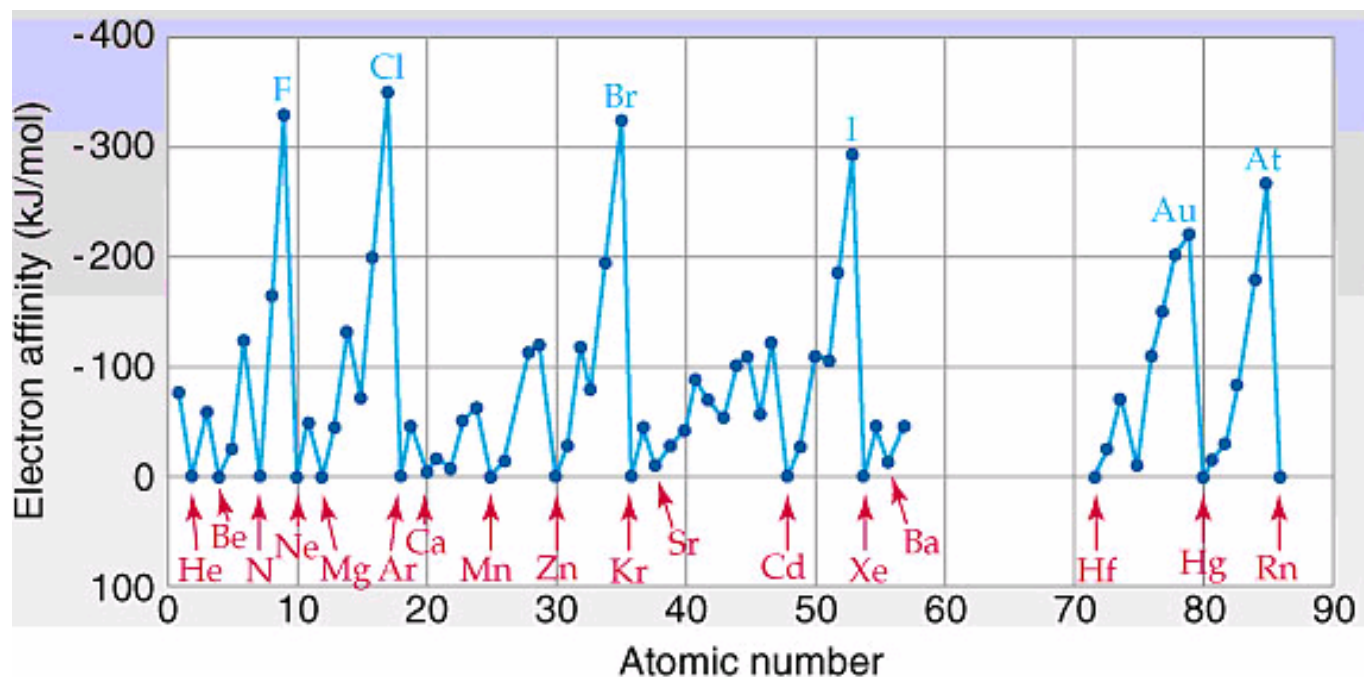


Electron Affinity: - the change in energy associates with an addition of an electron to a gaseous atom.
 - the larger negative electron affinity, the more stable the anion formed.



Several Notes on Trends in Electron Affinities

- In general, **Electron Affinity DECREASES Down a Group (less energy is released)**. This is due to the fact there are more orbitals as the number of row increases. Since the protons in the nucleus cannot attract another electron as effectively due to the increase distance involved, less energy is released. **The trend holds except for row 2. The reason is because of the small size of the 2p orbitals. Electron repulsion cause smaller values of electron affinities than expected for row 2.**
- In most cases, **Electron affinity INCREASES (becomes more negative) across the Period from Left to Right up to the Halogen group (more energy is released)**. Metals have $EA > 0$ because they like to form cations (low ionization energy). Non-metals have $EA < 0$ because they like to form anions in order to form a stable octet.
- Exceptions can be found in the 2A and 5A Groups**. With the 2A Alkaline Earth Group, it is more difficult to add an electron into the s orbital such that it will be filled to $ns^2 np^1$. This is because the 2A Group already has an electron configuration of ns^2 (subshell is already filled). Therefore, they have a more positive (less negative) electron affinity. In the case of the 5A Group, the atoms find themselves difficult to add another electron because this will mean a np^4 configuration, which contains a paired electron from three unpaired electrons in np^3 . Again, Hund's Rule frowns on such arrangement because it is not of the lowest energy state. Thus, Group 5A tends to have a more positive (less negative) electron affinity.



Assignment

8.4 pg. 349–350 #49 to 52, 54, 56

8.5 pg. 350 #59 to 64