Enloe Chemistry
Unit 2

Name:______________________________
Date:_______________Period:_________

Periodic Trends

Introduction

In Unit 1 we discussed the electronic structure of the atom. We learned that the elements in the periodic table can be categorized in to four distinct blocks, relating electronic structure directly to the periodic table. What is interesting to mention is that the periodic table was organized based off of the properties of the elements, it wasn't until later that the quantum mechanical model for the atom was presented and found to support the way the periodic table was organized.

As we have said before knowing the electronic structure of an atom is important in recognizing physical and chemical properties of elements. In this unit we will discuss these properties and trends that can all be related back to electron arrangement.

Chm.1.3.1 Classify the components of a periodic table (period, group, metal, metalloid, nonmetal, transition).
Chm.1.3.2 Infer the physical properties (atomic radius, metallic and nonmetallic characteristics) of an element based on its position on the Periodic Table.
Chm.1.3.3 Infer the atomic size, reactivity, electronegativity, and ionization energy of an element from its position on the Periodic Table.

Students should be able to:

<u>Periodic Table</u>

- Know that main group elements in the same family have similar properties, the same number of valence electrons, and the same oxidation number
- Understand that reactivity increases down in a group of metals and decrease down in a group of nonmetals

<u>Periodic Trends</u>

- Define atomic radius, ionic radius, ionization energy, and electonegativity.
- Know group and period general trends for atomic radius, ionic radius, ionization energy, and electonegativity.
- Apply trends to arrange elements in order of increasing or decreasing atomic radius, ionic radius, ionization energy, and electonegativity.
- Explain the reasoning for the trends
- Compare cation and anion radius to neutral ion
- Know group and period general trends for reactivity of metals and nonmetals and apply these trends.
- Know the characteristics of metals and nonmetals
- Use electron configuration and behavior to justify metallic character. (Metals tend to lose electrons in order to achieve the stability of a filled octet.)
- Relate metallic character to ionization energy, electron affinity, and electronegativity.

Assignment 1: Periodic Table

1. Elements in the same column of the periodic table contain the same number of valence shell electrons.

   a. Valence shell electrons are s and p electrons in the highest energy level.

Group #	Electron Configuration	Valence Electrons
IA(1)	ns^1	1
IIA(2)	ns^2	2
IIIA(13)	ns^2np^1	3
IVA(14)	ns^2np^2	4
VA(15)	ns^2np^3	5
VIA(16)	ns^2np^4	6
VIIA(17)	ns^2np^5	7
VIIIA(18)	ns^2np^6	8

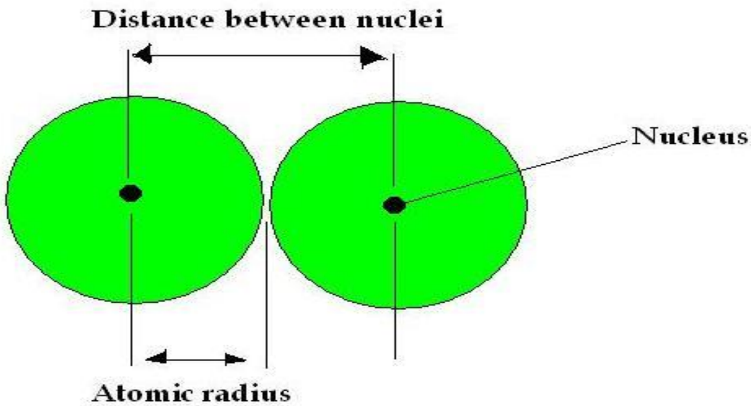
* Except for He which has 2

Columns are Groups or Families, Periods are Rows.
Group 1: Alkali Metals
Group 2: Alkaline Earth Metals
Group 3 – 12: Transition Metals

Group 1, 2, 13 – 18: Representative Elements
Group 16: Chalcogens
Group 17: Halogens
Group 18: Noble gases

Assignment 2: Periodic Trends

<u>Atomic Radius</u>: Half of the distance between the nuclei of two atoms of the same element that are bonded together



-Atomic Radius decreases as you move from left to right across the periodic table

-This is because the increasing number of protons (effective nuclear charge) draws the electrons closer to the nucleus.

-Atomic Radius increases as you move down a group in the periodic table

-This is because the valence electrons are in higher energy levels farther from the nucleus.

<u>Ionic Radius</u>

-Cations are smaller than their parent atoms.

-This is because the formation of a cation removes electrons from the outermost energy level.

-Anions are larger than their parent atoms.

-This is because the addition of electrons to the outer energy level results in the electrons repelling each other and spreading apart. Since the number of protons stays the same, but the number of electrons increases, the pull from the nucleus on each electron is less. (like a tug of war!)

1. Referring to a periodic table, arrange the following atoms in order of increasing size.

   a. O, S, F

   b. Na, Be, Mg

2. Select the larger species in each of the following pairs.

   a. Fr F

   b. Kr Xe

   c. Mg K

   d. Si P

   e. O P

   f. Na Cl

   g. Cl Cl^-

   h. Na Na^+

   i. Fe^{+2} Fe^{+3}

1st Ionization Energy: minimum energy required to remove the first electron from the ground state of an isolated gaseous atom or ion.

2nd Ionization Energy: minimum energy required to remove the second electron from the ground state of an isolated gaseous atom or ion.

3rd Ionization Energy: minimum energy required to remove the third electron from the ground state of an isolated gaseous atom or ion.

-Ionization energy generally increases moving from left to right across the periodic table.

-This is because the attraction from the nucleus increases and the electrons are closer to the nucleus.

-Ionization energy decreases moving down a group in the periodic table.

-This is because the attraction between the nucleus and the electrons decreases as the distance of the electrons from the nucleus increases.

TABLE 7.2 Successive Values of Ionization Energies, *I*, for the Elements Sodium Through Argon (kJ/mol)

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	496	4560			(Inner-shell electrons)		
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

Study the chart above.

1. Explain why there is a significant increase in the the 3rd and 4th Ionization energy for Aluminum.

2. Explain the drastic increase in the first ionization energy of Silicon and Phosphorus.

3. Which of the following has the greatest first ionization energy ?
   a. S, Cl, Se, Br
   b. B, Al, C, Si

4. Choose the atom in each pair that has the LOWER first ionization energy.
   a. Al, B
   b. O, C
   c. K, Ca
   d. C, N
   e. Na, K
   f. Ar, K
   g. Sn, Br
   h. B, C

Electronegativity: The ability of an atom in a molecule to attract electrons to itself.

-Electronegativity generally increases moving from left to right across the periodic table. *EXCEPTION: Electronegativity does not apply to noble gases.*

-This is because the pull from the nucleus increases, and the ease with which atoms will gain an electron increases.

-Electronegativity generally decreases moving down a group in the periodic table

-This is because the pull from the nucleus decreases.

1. What elements are most electronegative?

2. What elements are least electronegative?

Metallic Character: Metals are substances with low Ionization energies, because they willingly give up electrons to form cations. Nonmetals have high ionization energies because they DO NOT want to let go of their electrons. Nonmetals do have high electron affinities (ease at which atoms gain electrons) and high electronegativities, which is why they attract electrons from other atoms to form anions. Metals have low electron affinities and electronegativities because they do not want any electrons.

Remember the goal is to achieve an octet of electrons in the valence shell. It is easier for metals to lose electrons to achieve an octet, and for nonmetals to gain electrons to achieve an octet.

Metallic character decreases moving from left to right across the periodic table; nonmetallic character increases moving from left to right across the periodic table.

-This is because of the trends in Electronegativity, Ionization energy, and electron affinity.

Metallic character increases moving down a group in the periodic table; nonmetallic character decreases moving down a group in the periodic table.

-This is because of the trends in Electronegativity, Ionization energy and electron affinity.

Metals	Nonmetals
Have luster	Do not have luster
Solids are malleable and ductile	Solids are usually brittle
Good conductors of heat and electricity	Poor conductors of heat and electricity
Tend to form positive ions (cations)	Tend to form negative ions (anions)
Most metals are solids	Many nonmetals are gases
Most have 3 or fewer valence electrons	Most have 5 or more valence electrons
Low ionization energy	High ionization energy
Low electronegativity	High electronegativity
Low electron affinity	High electron affinity

1. Which group contains the most active:
    a. Metal
    b. Nonmetal
2. Which element is the most active
    a. metal
    b. nonmetal

Oxidation State

3. The oxidation state is numerically the same as a charge for an ion. The oxidation state depends on the number of valence electrons the atom has. What is the most common oxidation state of the elements below:
   a. Group 1: _______
   b. Group 2: _______
   c. Aluminum: _______
   d. Zinc: _______
   e. Silver: _______
   f. Nitrogen:_______
   g. Oxygen: _______
   h. Sulfur: _______
   i. Group 17: _______
   j. Group 18: _______

4. The oxidation state of the _______________ metals varies because of their partially filled ____ subshell
5. The oxidation state of metalloids and other elements near the staircase depends upon their electronegativity and the electronegativity of the atom in which they are bonded with.

Assignment 4: Problems

A. Atomic Size: Circle the larger atom in each pair.

1. Ar or Ne 2. N or P 3. Ca or Sc

4. B or C 5. Cl or F 6. K or Br

Arrange in order of increasing atomic radius.

7. O, S, F 8. Na, Be, Mg

9. Li, O, C, F 10. Br, At, F, I, Cl

11. Mg, Cl, Na , P 12. Ca, Be, Ba, Sr

B. Ionic Size: Circle the larger species in each pair.

1. Li or Li^+ 2. Cl or Cl^- 3. In^{3+} or Rb^{1+} 4. P^{3-} or Cl^-

5. N^{3-} or As^{3-} 6. Mg or Mg^{2+} 7. S or S^{2-} 8. K^+ or Li^+

9. Put the following in order of decreasing atomic radius:

As, As^{3+}, As^{3-} , As^{5+} _________________________

C. Ionization Energy

1. What feature of the electron configuration will cause a large increase in the first ionization energy?

2. Which elements tend to have higher first ionization energies?
   Metals or Nonmetals

3. Circle the atom with the larger first ionization energy

a. B or F	b. N or P	c. Rb or Cs
d. O or S	e. S or Cl	f. Ga or Ge

D. Electronegativity: Write the element that has the highest electronegativity on the line

1. Na, Al, P, S ______
2. O, S, Se, Te ______
3. Li, Be, Mg, Na ______
4. P, S, As, Se ______

Circle the element with the higher electronegativity.

5. Na or Cl   6. C or O   7. Cl or H   8. Na or H   9. Fe or O

E. Oxidation Number: Write the most common oxidation number for the following elements.

1. Al	_______	2. S	_______
3. Mg	_______	4. Ag	_______
5. N	_______	6. K	_______
7. I	_______	8. Ne	_______
9. H	_______	10. Zn	_______
11. Bi	_______	12. Cl	_______

13. Which of the following cations is not likely to form:

Sr^{2+}, Al^{3+}, K^{2+}, Cl^{2-}? ________

14. Which of the following anions is least likely to form:

I^{-} , Cl^{-} , Na^{2+} , O^{3-} ________

F. Chemical Activity: Circle the more active element

1. K or Cs	2. Ba or Mg	3. F or Cl	4. Se or O
5. K or Zn	6. Mg or Au	7. Ag or Au	8. Cl or I

Assignment 5: Electron Configuration of Ions

- We have written electron configurations for neutral atoms in their ground state. We can apply the same principles we have already learned and use them to help us write electron configurations for ions.
- Remember an ion is an atom that has either lost or gained an electron.
- If an atom loses an electron, then it will become positively charged, forming a cation.
- If an atom gains an electron, then it will become negatively charged, forming an anion.

EX)

Na: $1s^2 2s^2 2p^6 3s^1$

Na^{1+}: $1s^2 2s^2 2p^6$

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

Cl^{1-}: $1s^2 2s^2 2p^6 3s^2 3p^6$

- When writing the electron configuration of an ion of a representative element we will always add or remove electrons from the highest energy **s** and **p** orbital**.**
- With that in mind if we wanted to write an electron configuration for an ion of any element on the periodic table, remove electrons in the following order:
    - Highest energy p electron then highest energy s electron, and lastly the highest energy d electron.

EX)

Fe:	Cu:	Mn:	Ga:
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$
Fe^{2+}:	Cu^{1+}:	Mn^{2+}:	Ga^{1+}:
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^4$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
Fe^{3+}:	Cu^{2+}:	Mn^{4+}:	Ga^{3+}:
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^9$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^2$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

Practice: Write the electron configuration for the following ions:

1. Cr^{6+}
2. As^{3-}
3. Sn^{4+}
4. Ca^{2+}
5. N^{3-}
6. S^{2-}