*Unit 4 – The Periodic Table and Trends*





*Unit 4 Vocabulary*

Periodic Law

Atomic Radius

Ionic Radius

Electronegativity

Ionization Energy

Chemical Reactivity of Metals

Chemical Reactivity of Non-metals

Alkali Metals

Alkaline Earth Metals

Halogens

Noble Gases

Representative Elements

Transition Metals

Inner-transition metals

**Effective Nuclear Charge (Zeff)**

**Coulomb’s Law**

*Notes: The Periodic Table and Ion Formation*

I. The Periodic Table

**Dmitri Mendeleev** (Russian)

* Noticed that when the elements were placed in order of increasing atomic mass there was a repeating pattern or periodicity to the elements properties.
* He created a table that placed elements with similar chemical properties in vertical columns.
* Some elements were list in order of their chemical properties even though their masses did not agree. Published the 1st Periodic Table in 1869; it contained several "holes."
* Predicted the discovery of and the properties of 3 additional elements; all were discovered by 1886.

**Henry Moseley** (English)

* While working with Rutherford (discovered the nucleus) and discovered that the positive charge of the nucleus increased by one for element to the next in the periodic table; this work lead to the modern definition of the atomic number.
* His work justified Mendeleev's decision to arrange by properties when the masses did not agree.
* Credited with the modern Periodic Table being arranged by increasing atomic number.

### II. Definitions

**The Periodic Law:** The physical and chemical properties of the elements are periodic functions of their atomic #s.

**Main group/Representative Elements:** Elements in the s and p blocks (also known as the "A" group elements.)

**Transition Metals/Elements:** Elements in d block (also known as the "B" group elements.)

**Inner Transition Metals/Elements:** Elements in the f block (no group designation.)

**Valence Electrons:** The electrons available to be lost, gained, or shared in the formation of chemical bonds. These are the s and p electrons in the outermost energy level.

**Noble Gas Shorthand notation:** The Noble gases can be used as shorthand to writing out the electron configuration. Use the Noble gas that occurs before the element and then write only what comes after.

Example: Kr: 1s22s22p63s23p64s23d104p6 **Shorthand for Kr**: [ Ar ] 4s23d104p6

Ag: 1s22s22p63s23p64s23d104p65s24d9 **Shorthand for Ag**: [ Kr ] 5s24d9

### III. Properties of Specific Groups

**Alkali Metals:** Group 1, most reactive metals, combine vigorously with nonmetals, not found freely in nature, react vigorously with water to form hydrogen gas.

**Alkaline Earth Metals:** Group 2, slightly less reactive than the Alkali metals, not found freely in nature, harder, denser, and stronger than alkali metals.

**Halogens:** Group 17, most reactive nonmetals, "salt formers.”

**Noble Gases:** Group 18, also called inert or rare gases, very nonreactive due to outermost s and p sublevels being full, all discovered between 1894 and 1900.

**IV. Properties of Metals and Non-Metals**

The “staircase” line descending from group 13 down to group 16 separates metals on the left side from nonmetals on the right. About three-quarters of the elements are metals, including some main group elements and all of the transition and inner transition elements. Properties of metals include:

• solids at room temperature, except for mercury, which is a liquid

• generally shiny or lustrous when freshly cut or polished

• good conductors of heat and electricity

• generally malleable, which means they can be rolled or hammered into thin sheets

• generally ductile, which means they can be rolled or stretched into wires

***• during chemical changes, tending to give up electrons relatively easily to form cations***

The non-metals are located in the upper right portion of the table. Properties of non-metals include:

• usually gases or brittle solids at room temperature, except for liquid bromine

• poor conductors of heat and electricity

***• during chemical changes, tending to gain electrons from metals to form anions or share electrons with other non-metals.***

#### IV. Formation of Ions in the Main Group Elements

**Main Group Elements or Representative Elements**: The elements that are characterized by the filling of their outermost s and p sublevels; the s and p block elements; groups.1,2,13 - 18 (the "A" groups.)

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| Group  Number |  |  |  |  |  |  |  |  |
| Electron  Configuration |  |  |  |  |  |  |  |  |
| Valance  Electrons |  |  |  |  |  |  |  |  |
| Gain or  Lose |  |  |  |  |  |  |  |  |
| Ionic  Charge |  |  |  |  |  |  |  |  |

**Electron Configuration:** This is the electron configuration of the outermost energy level. The letter n stands for the energy level and corresponds to the period where the element is found in the periodic table.

**Valence electrons:** The electrons in the outermost energy level. These are also the electrons that are available to be gained, lost, or shared in chemical bonds. These are always found in the s and p sublevels of the outermost energy levels. They will never be found in a d sublevel.

**Gain or Lose:** Atoms seek to have an electron configuration that is like a noble gas - one that ends in ns2np6. In other words, atoms will gain or lose electrons to achieve a noble gas configuration. In general, metals will lose electrons and nonmetals will gain electrons.

**Ionic Charge:** This is the charge an atom will have once it has gained or lost electrons to become an ion. If the atom loses electrons it becomes a cation and has a positive charge. If the atom gains electrons, it becomes an anion and has a negative charge.

**Transition Metals and Inner Transition Metals (Rare Earths):** For now, assume they lose 2 electrons resulting in an ionic charge of +2. They are actually much more complicated!

Notes: PERIODIC TABLE TRENDS

Atomic Radii

1. Increases with atomic number within a family; each element in family has one more energy level.
2. Decreases in size within a period from left to right
   1. More protons in nucleus.
   2. All electrons in same energy level.
   3. Electrons more attracted to nucleus.

Decreases

Increases

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*General trends in atomic radii of A Group elements with position in periodic table*

Ionic Radii

1. Follows same trend as atomic radii.
2. (+) ions are always smaller than neutral atoms.
3. (–) ions are always larger than neutral atoms.

Decreases

Increases

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*General trends in ionic radii of A Group elements with position in periodic table*

Ionization Energy (I.E.)



1. First I.E. required to remove an electron from an atom; it measures how tightly electrons are held.
2. Removal of electron forms (+) ion.
3. Metals have low I.E.; non-metals have high I.E.
4. Within a family, I.E. decreases with increasing atomic number.
5. Second I.E. is always greater than 1st I.E. because it’s more difficult to remove an e- from a (+) ion.

Increases

Decreases

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*General trends in first ionization energies of A Group elements with position in periodic table*

*(exceptions occur at Groups IIIA and VIA)*

Electronegativity

1. Measure of the tendency of an atom to attract electrons to itself when combined with another atom.
2. Elements with large differences in electronegativity tend to form ionic compounds.
3. Elements with low differences in electronegativity tend to form covalent compounds.

Increases

Decreases

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*General trends in electronegativity of A Group elements with position in periodic table*

Chemical Activity (Non-metals)

1. How easily if forms a compound.
2. Fluorine most active non-metal (large E.A.)

Increases

Decreases

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*General trends in chemical activity of nonmetals of A Group elements with position in periodic table*

Chemical Activity (Metals)

1. How easily if forms a compound.
2. Francium most active metal (smallest I.E.)

Decreases

Increases

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*Notes: Explaining Periodic Trends*

Tools to know to help explain periodic trends

* Effective Nuclear Charge: The attractive force a valence electron experiences from the nucleus.
  + # protons minus # core electrons = effective nuclear force

Z – core electrons = Zeff

* + This is the same as the number of valence electrons for most atoms.
* Coulombs law: The force between two charged particles is directly proportional to the product of their charges and inversely proportional to the distance separating the charges.
  + The higher the charges between two oppositely charged particles, the stronger the attractive force is between them.
  + The further apart two charges are, the smaller the magnitude of force between them.

Factors Affecting Atomic and Ionic Radius

Since the size of an atom is determined by the position of the electrons, we should identify what affects the size of an atom’s electron cloud. Two factors influence the size of the electron cloud

1. The number of energy levels present. As the number of energy levels (n) increases, the probability that outer electrons will spend more time further from the nucleus increases and so the atoms become larger.
2. The amount of nuclear charge experienced by the outer electrons. As the effective nuclear charge increases, the outer electron cloud is pulled closer to the nucleus and the atom becomes smaller. Chemists use the symbol “Zeff” to refer to the effective nuclear charge.

**Forces Affecting Ion Size**

The same forces that influence the sizes of atoms also influence the sizes of ions and in the same way. Consider these forces and rank the following species in order of size from largest to smallest:

Al3+ , F−, Mg2+, N3-, Na+, Ne, O2–.

Note that all of the ions have 10 electrons and are therefore isoelectronic with neon.



Since each species has the same number of electrons, Zeffis the predominant force that influences size. This means that the attractive force from each nucleus is the only factor influencing the size of each species. The greater the number of protons present, the stronger the attractive force on the electron cloud and therefore the smaller the atom or ion.

N3– > O2– > F− > Ne > Na+ > Mg2+ > Al3+

**Factors Affecting Ionization Energy**



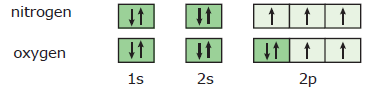
Ionization energy (IE) is the minimum energy required to remove an electron from a gaseous atom or ion. The term is often used to mean the “first” ionization energy (IE1) whereby a neutral atom becomes a 1+ cation according to the following equation:

atom (g) + IE1 ➝ ion + (g) + e–

Ionization energy tells us how strongly an atom holds onto its outermost electrons.

This is an important property because an element with a low IE1 will be more likely to lose electrons and form cations during chemical changes. A high IE1 might signal an element’s tendency to gain electrons and form anions or perhaps not forms ions at all.

There are exceptions to the general trend for ionization energy which we can explain by analyzing electron configurations. One example of this occurs with nitrogen and oxygen. Even though oxygen is a smaller atom than nitrogen, oxygen has a lower first ionization energy. Nitrogen has a single electron in each of its three 2p orbitals. Oxygen, however, has a pair of electrons in one of its 2p orbitals. The increased electron-electron repulsion associated with that pairing makes it easier for oxygen to lose one of those electrons when first ionized.



This reminds us that the repulsive forces between electrons, as well as the attractive forces affecting them from the nucleus, have a role to play in determining properties such as size and ionization energy.

**Ionization Energies can be used to determine the number of valence electrons.**

Beryllium has two valence electrons and the electron configuration is 1s2 2s2.

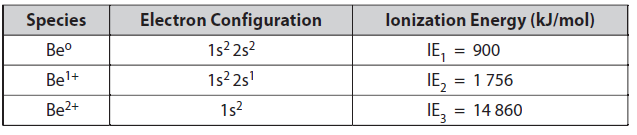
After the second (and last valence) electron is removed, the dramatic increase in

energy required to remove the third electron reflects the fact that it is an inner or “core”

electron. This shows us that core electrons are bound much more tightly to the nucleus,

and thus do not take part in chemical reactions. This holds true for all of the elements in

the periodic table.



*Homework #1 - Periodic Table and Ion Formation*

***Complete the following statements.***

1. If the outer electron configuration of an element is 2s1, the element directly below it in the periodic table will have an outer electron configuration of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

2. The chemical family in which the outermost s and p orbitals are completely filled is called the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

3. The element with an outer electron configuration of 1s22s22p5 is in the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_family.

4. Elements in the fourth period have the core electron configuration of the noble gas \_\_\_\_\_\_\_\_\_\_\_\_\_.

5. The number of valence electrons in all elements in Group 14 is \_\_\_\_\_\_.

6. The number of valence electrons in beryllium, an alkaline earth metal, is \_\_\_\_\_\_.

7. The number of valence electrons in chlorine is \_\_\_\_\_\_.

8. The electron configuration for sodium is 1s22s22p63s1. The electron configuration for the sodium ion is\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

9. The most likely charge on the ion formed by iodine is \_\_\_\_\_\_\_\_ and by aluminum is \_\_\_\_\_\_\_\_\_

10. The calcium ion is isoelectronic with the element \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

***Write the symbol for the most common ion formed by each of the following elements.***

11. oxygen O-2 14. argon \_\_\_\_\_\_ 17. aluminum \_\_\_\_

12. fluorine \_\_\_\_\_\_ 15. nitrogen\_\_\_\_\_\_ 18. hydrogen \_\_\_\_

13. cesium \_\_\_\_\_\_ 16. carbon \_\_\_\_\_\_ 19. barium \_\_\_\_\_

*Homework #2 – Trends*

***Define terms where needed and answer the questions:***

1. **Atomic Radius:**

2. Within a group, as the atomic number of an element increases, the atomic radii\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

3. Within a period, as the atomic number of an element increases, the atomic radii \_\_\_\_\_\_\_\_\_\_\_.

4. Circle the element that has the largest atomic Radius:

a. Mg / Na b. Al / B c. F / N d. K / Ca e. Br / Cl f. Ne /Ar

5. **Ionic Radius:**

1. When metallic atoms lose electrons, they form ( larger / smaller ) ions.
2. When nonmetallic atoms gain electrons, they form ( larger / smaller ) ions.
3. Circle the larger particle:

a. Al / Al+3 b. Mg / Mg+2 c. S-2 / S d. Br-1 / Se-2 e. As-3 / As f. F / F-1

9. **Ionization Energy:**

1. The 1st IE tends to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ as the atomic number increases in any period.
2. The 1st IE tends to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ as the atomic number increases in any group.
3. Metals have a \_\_\_\_\_\_\_\_\_\_\_\_\_\_ 1st IE, while nonmetals have a \_\_\_\_\_\_\_\_\_\_\_\_\_ 1st IE.
4. As the distance between nucleus and outer electrons of an atom increases, the IE will \_\_\_\_\_\_\_\_
5. As the charge of an ion increases, the IE will \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
6. Circle the element that has the lower 1st IE:

a. Li / Na b. Cs / Ba c. F / Ne d. Kr / Rb e. Cl / Br f. S / Cl

1. List and describe the factors that affect ionization energy.

17. **Electronegativity:**

1. Metals tend to have \_\_\_\_\_\_ electronegativities, while nonmetals tend to have \_\_\_\_\_\_ electronegativities.
2. \_\_\_\_\_\_\_\_\_\_\_\_\_ tend to accept the electrons and \_\_\_\_\_\_\_\_\_\_\_\_\_\_ tend to donate electrons.  
     
   20. **Chemical Activity:**

21. Circle the most active element:

a. Rb / Sr b. S / Cl c. O / Te

*Homework #3 - More Practice With Periodic Table Trends*

***I. Define each of the following:***

Atomic radius:

Ionic radius:

1st ionization energy:

Electronegativity:

Chemical Activity:

***II. Answer the following:***

1. Circle the larger particle:

a. I or I-1 b. Sr or Sr+2 c. Mg or Be

d. N-3 or F-1 e. Cs+1 or Ba+2 f. K or Ca

g. Si or S h. O or Se i. Ga+3 or As-3

2. Circle the particle with the lowest first ionization energy

a. F or Cl b. Li or Cs c. Rb or Sr

d. C or O e. F or Fr f. Sr or I

3. Circle the particle with the highest electronegativity

a. P or S b. Na or Mg c. C or Si

d. Be or Ba e. Al or Cl f. Li or F

4. Circle the particle with the lowest electronegativity

a. S or Cl b. K or Cs c. B or N

d. Na or Al e. O or Te f. Mg or K

5. Circle the particle with the greatest chemical reactivity

a. N or O b. Br or I c. Ba or Sr

d. Sn or Pb e. Ga or Ge f. Li or Be

6. Metals will \_\_\_\_\_\_\_\_\_\_ electrons to form \_\_\_\_\_\_\_\_\_\_ which have a \_\_\_\_\_\_\_\_\_\_\_\_ charge. (gain/lose) (cation/anion) (positive/negative)

7. Nonmetals will \_\_\_\_\_\_\_\_\_\_ electrons to form \_\_\_\_\_\_\_\_\_\_ which have a \_\_\_\_\_\_\_\_\_\_ charge. (gain/lose) (cation/anion) (positive/negative)

8. Cations are always \_\_\_\_\_\_\_\_\_\_ than the neutral atom because the effective nuclear

(larger/smaller)

charge on cations is \_\_\_\_\_\_\_\_\_\_.

(larger/smaller)

9. Anions are always \_\_\_\_\_\_\_\_\_\_ than the neutral atom because the effective nuclear

(larger/smaller)

charge on anions is \_\_\_\_\_\_\_\_\_\_.

(larger/smaller)

10. Determine the charge of the most common ion formed by each of the following:

a. Na b. S c. Al d. Sn

e. Ba f. Cl g. P h. C

i. Cd

*Homework #4 Explaining Trends*

1. Which of the two opposing factors that influence atomic size predominates as we move across a chemical period? What is the general result?
2. Which of the two opposing factors that influence atomic size predominates as we move down a chemical family? What is the general result?
3. Why is it difficult to measure the sizes of individual atoms?
4. Briefly explain why fluorine is a smaller atom than lithium. Consider which factor is predominating across a period.
5. The attraction of electrons to the nucleus and repulsion of the electrons between each other both influence the size of an atom or ion. Use this to complete the following statements.
   1. A cation will always be \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_(smaller or larger) than its parent neutral atom because of \_\_\_\_\_\_\_\_\_\_ (increased or decreased) attraction of the outer electrons for the nucleus and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (increased or decreased) repulsion of the electrons for each other.
   2. An anion will always be \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_(smaller or larger) than its parent neutral atom because of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (increased or decreased) attraction of the outer electrons for the nucleus and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (increased or decreased) repulsion of the electrons for each other.
6. What role do inner or core electrons play in determining atomic size and ionization energy?
7. Consider the first two ionization energies for lithium: IE1 = 519 kJ/mol IE2 = 7 285 kJ/mol. Explain why lithium’s second ionization energy is more than 10 times its first.
8. The nature of the 2s sublevel is such that 2s electrons have a higher probability of being found closer to the nucleus than electrons in the 2p sublevel. Consider this and the following electron configurations:

beryllium: 1s2 2s2 boron: 1s2 2s2 2p1. Suggest a reason why boron’s first ionization energy is less than beryllium’s, even though boron is a smaller atom.

1. Write the electron configuration for nickel and zinc. Use these to explain why an atom of zinc is larger than an atom of nickel.