Topic 5: The Periodic Table

- **Atoms only lose/gain/share electrons in order to have the same number of electrons as a noble gas** (stable)
- When 2 atoms have the same number of electrons means its isoelectronic.

The Periodic Table (History):

- **Mendeleev organized the elements in the periodic table in order of increasing atomic mass.**
  a. Based on this organization there was a periodic repetition of physical and chemical properties.
  b. However, with this organization strategy there were pairs of elements that seemed out of order (I.g Iodine and Te)
- Henry Moseley created our current periodic table
  a. He arranged the elements based on ascending atomic number rather than atomic mass.
  b. He determined the atomic number of an element based on X-rays
- The Periodic Law states: **The properties of all elements are periodic functions of their atomic numbers.**

The Periods:

- **Horizontal rows of the table are called periods**
  a. Number at the beginning of the period indicates the principal energy level in which the valence electrons are located

(Ex: K and Br are in period 4, so their valence electrons are in the 4th principle energy level)

- The number of valence electrons increases from left to right
- The **properties of elements change systematically across a period**

The Groups:

- **Vertical columns of the periodic table are called groups** (families)
- The **number of valence electrons for each element is the last number of the electron configuration**

(E.g: Phosphorous has an electron configuration of 2-8-5, the number of valence electrons is 5)

- **All elements along the same group have the same number of valence electrons.**
  a. Exception to the above rule is Helium, in group 18 which has only 2 valence electrons while the rest in group 18 have 8.
- The **number of valence electrons determines much of the chemical reactivity of the element.** Members of the same group have similar chemical properties.
- **Down a group metallic character increases because e- are farther away from the nucleus, easier to remove them.**

Properties of groups:

- **Hydrogen does not have physical or chemical properties similar to the group 1 metals.**
  a. In a **chemical reaction hydrogen often loses its valence electron to form metallic hydrides**.
- **Group 1 (alkali metals) and group 2 (alkaline earth metals) easily lose their electrons and are never found in nature in their atomic state.**
- **Francium is the most reactive metal**

- Only when oxygen bonds to the only element more electronegative than itself (fluorine) does oxygen have a positive oxidation value
- Members of group 17 are called halogens when they are in the free state.
  a. When they form 1- ions the salts formed are called halides.
- **When combining a group 1 element with a group 17 element the resulting formula is normally MX** where M represents the group 1 element and X represents the group 17 element (E.g KI)
- **When combining a group 2 element with a group 17 element the general resulting formula is MX_2** where M is the group 2 element and X is the group 17 element (e.g. $BeCl_2$)

Metals, Metalloids and Nonmetals:
Metals:
- On the left are generally metals (comprising of 75% of the periodic table)
  a. All metals are solid with the exception of mercury
- The most active metals are located in groups 1 and 2
- Metallic properties of elements increase from the top to the bottom of the group
- **Metals are malleable, ductile and have luster**
- **Metals have relatively low ionization energy and electronegativity values**
- **They are good conductors of heat and electricity, this stems from the mobility of their valence electrons.**
- Elements of group 3 to 12 are called transition elements, or sometimes transition metals
  a. Transition elements in each period represents a series of elements in which the outermost d orbitals are being filled.

(E.g Along Period 4 the transition elements proceed from scandium to zinc and going from left to right across these elements the 3d orbital is being filled)

  b. Characterized by multiple oxidation states
  c. Transition elements are far less reactive than metals of groups 1 and 2
  d. Another unique property of such elements are that when they form ions they gain color
- Metals have lower ionization energy and electronegativity because they have fewer valence electrons
- **When metals become positive ions they decrease in size because they lose an energy level.**

Metalloids
- To the right of the metals are metalloids (only 6 metalloids), Bm Si, Ge, As, Sb,and Te which form a stair-step line
  a. Nonmetals include solids, a liquid (bromine) and gases.
- Metalloids have both the properties of metals and nonmetals
- Each period ends with a noble gas.
- Represent an “intermediate” type of element

Metalloids
- To the right of the metalloids are the nonmetals.
- Although the properties of nonmetals vary more than those of nonmetals, some standard properties can be observed

a. Many nonmetals are gases or molecular or network solids at room temperature. The exception is bromine at room temperature.

- Nonmetals are not malleable or ductile, they tend to be brittle.
- Solid nonmetals lack luster
- **They have high ionization energy and high electronegativity values.**
- Poor conductors of heat and electricity.
- **Nonmetals tend to gain electrons to become negative ions**

The Noble gases:

- Metals in group 18 are called noble gases
- **Noble gases don't have all the properties of nonmetals because they are generally unreactive**
- Only a few stable compounds containing noble gases have been formed.

each of these elements have completely filled out energy levels (8 valence electrons) with the exception of helium with only 2 valence electrons.

- Extremely stable
- Once called the inert gas group because it was thought these elements could not react to produce stable compounds.
- **However, fluorine with such a high attraction can sometimes even attract electrons from noble gases.**
- Only Argon and Xenon have formed stable stable compounds
- **Helium alone has not show any reactivity.**

Allotropes

- **Some nonmetals can exist in two or more forms in the same phase, known as allotropes.**

(Ex: Oxygen and Ozone)

- Allotropes have different physical properties and they also differ chemically.

Properties of elements:

Metallic character: How much an element is like a metal

Ionization energy:

- **The amount of energy needed to remove the most loosely bound electron from a neutral gaseous atom.** Atoms with more than one electron have more than one ionization energy, but the first ionization energy is the most significant.
- Trends in a period: **Values from left to right across a period generally increase, as the atomic number of an atom increases the nuclear charge increases.** Thus electrons become more strongly attacks thus more energy is needed to separate the electrons from the atoms.
- Trends in groups: **Down a group ionization energy decreases because valence electrons in every successive element are at a higher energy level, thus, they are further away from the nucleus.** As distance between the electron and the nucleus increases, the energy required to pull the electron away decreases.

Atomic Radii:

- The radius of an atom is a good measure of the size of the atom.
- **Defined as half the distance between two adjacent atoms in a crystal or half the distance between the nuclei of identical atoms that are bonded together.**
- **Down a period (from left to right) the repeating pattern is that the atomic radii decreases**. In each period the metals have larger radii than nonmetals. This is because although down a period the number of valence electrons are at the same general energy level, the number of protons attracting the electrons increases, causing the radius to decrease.
- Down a group (from top to bottom), each successive element have more inner-level electrons than the preceding member. These electrons shield the valence electrons from the attractive force of the nucleus. The valence electrons are more energetic than those of elements higher in the group. **Simply put, down a group with more energy levels and the same amount of valence electrons, the atomic radius increases because the amount of energy levels between the valence electrons and the nucleus increases, decreasing attractive force.**
- **As atomic radius' increases they favor losing electrons. lower ionization energy**
- Atomic radius decreases because the number of protons increases which means forces of attraction increase, p/e- ratio

Ionic Radii

- **Atomics gain or lose electrons to become charged particles called ions**
- **Nonmetals tend to gain valence electrons in chemical reactions to become negative ions.**
- As atoms gain or lose electrons they complete an octet of valence electrons
- **Ionic radius is the distance from the nucleus to the outer energy level of the ion.**
- When an atom loses its valence electrons and becomes positively charged it loses an energy level and its radius decreases.
- The radius of a metallic ion is smaller than the radius of the atom.
- When nonmetals for ions the radius increases

Electronegativity:

- **Value of an atom is a measure of its attraction for electrons when bonded to another atom.**
- electronegativity values can be used to predict the type of bond that will form between two atoms.
- Noble gases normally do not have electronegativity values because they are already stable and form very few stable chemical compounds.
- **Along a period electronegativity values normally increase from left to right. Metals tend to have low electronegativity values while nonmetals have higher**
- **In each group the highest electronegativity value is found at the top. Attraction for bonded electrons is less towards the bottom of the group.**
- **Fr has the lowest ionization energy and electronegativity** because it has 1 valence electron and 7 energy levels resulting in the lowest attraction.
- On the other hand, **Fluorine has the highest ionization energy and electronegativity** because it has 2 energy levels and 7 valence electrons, resulting in the strongest attraction.

Reactivity of elements:

- Some elements can be found uncombined in nature, the atomic state.
- Noble gases are always found as free elements, they rarely react.
- Elements in Groups 1, 2, and 17 are rarely found to be uncombined.

Properties of Groups

- All members of a group have the same number of valence electrons. Although each member of a group have the same number of valence electrons, there can still be a change in the type of element from top to bottom.
- Hydrogen: Hydrogen is separated from the rest of group 1. Hydrogen does not have any chemical or physical properties similar to Group 1 metals. In a chemical reaction hydrogen often loses its valence electron.
- Groups 1 and 2: The elements of Group 1 are alkali metals and group 2 elements are alkaline earth metals, show typical metallic characteristics. Easily lose valence electrons and are never found in nature in their atomic state, always found in compounds. Can be reduced to their free states by electrolysis of their compounds. From top to bottom generally reactivity increases. Group 1 elements are more reactive in than Group 2 elements.
- Group 15: Nitrogen and Phosphorous are typical nonmetals and can acept 3 electrons. Nitrogen is a stable gas ar room temperature as it contains a triple bond between the two nitrogen atoms in N2.
- Group 16: Elements show progression from nonmetal to metal with increasing atomic number. Selenium and tellurium are metalloids. Oxygen is diatomic at room temperature. Oxygen is a reactive elements, easily forming compounds with other elements.
- Group 17: Are halogens when they are in the free state. When atoms of elements in this group gain an electron, the salts formed are halides. All elements are nonmetals,
- Group 18: Was once called the inert gas group as it was thought the elements could not react to produce stable compounds. Fluorine has such a high electronegativity, it can attract electrons from some noble gases. Like the halogen group, noble gases are nonpolar and therefore held in the liquid and solid phase by van der Waals forces.

Topic 2:

- Diatomic molecule: Elements that exist in nature as two identical atoms. (they are all nonmetals)
- Molecule: smallest unit of a substance
- Formulas: use of chemical symbols and numbers to show both qualitative and quantitative information about a substance.
- Qualitative: Things that cannot be counted or measured, e.g what elements are in a compound
- Quantitative: Things that can be either counted or measured, e.g. number of atoms in each element present in a unit of the compound

Endothermic Process: Process that require energy in order to occur. Physical change of ice melting. Energy required is absorbed from the surroundings, thus lowing the surrounding temperature. In endothermic processes, the reactions absorb energy as they become products Products have more potential energy.

Exothermic Processes: Process that release thermal energy. Burning of carbon in oxygen. Freezing of water. Energy released is given to the surroundings, raising the surrounding temperature. Products have less energy than the reactants.

Types of Reactions

Synthesis: A synthesis reaction is when two or more simple compounds combine to form a more complicated one. These reactions come in the general form of:

A + B ---> AB

One example of a synthesis reaction is the combination of iron and sulfur to form iron (II) sulfide:

$8\ Fe + S_8$ ---> $8\ FeS$

Decomposition: A decomposition reaction is the opposite of a synthesis reaction - a complex molecule breaks down to make simpler ones. These reactions come in the general form:

AB ---> A + B

One example of a decomposition reaction is the electrolysis of water to make oxygen and hydrogen gas:

$2\ H_2O$ ---> $2\ H_2 + O_2$

Single replacement : This is when one element trades places with another element in a compound. These reactions come in the general form of:

A + BC ---> AC + B

One example of a single displacement reaction is when magnesium replaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

$Mg + 2\ H_2O$ ---> $Mg(OH)_2 + H_2$

Double displacement: This is when the anions and cations of two different molecules switch places, forming two entirely different compounds. These reactions are in the general form:

AB + CD ---> AD + CB

One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

$Pb(NO_3)_2 + 2\ KI$ ---> $PbI_2 + 2\ KNO_3$

- Ionic formula:

a. Must be made up of a metal (s) and nonmetals) NaCl
b. Must be metals (s) and polyatomic ion(s) $CuSO_4$
c. Must be 2 polyatomic ions $(NH_4)_3PO_4$