Chapter 7: Periodic Properties of the Elements

- Periodic table is arranged according to the repeating patterns of electron configuration
- Elements in the same column contain the same number of electrons in their valence orbital

7.1: Development of the Periodic Table

- Some elements can be found as elemental form in nature
- Early nineteenth century, advances in chemistry make it easier to isolate elements from their compound and therefore the number of known elements increase from 31 to 63 (1800-1865)
- In 1869, Dmitri Mendeleev and Lothar Meyer noted that the properties of elements recur periodically when arrange by increasing atomic mass
- Mendeleev is given credit because he also predict the unknown element that would fit in the hole
- Henry Moseley developed the concept of atomic number
  - He found that each element produces different X-ray frequency when shower with high-energy electrons

7.2: Effective Nuclear Charge

- Coulomb’s law tells that the strength of the interaction between two electrical charges depends on the magnitude of the charge and the distance between them
- Estimate the net attraction of each electron to the nucleus by considering how it interacts with the average environment created by the nucleus and other electrons
  - Allow the electron to be treated individually as if it were moving in a net electric field
    - The net electric field is treated as if it’s caused by the positively charged nucleus and is called effective nuclear charge ($Z_{\text{eff}}$)
      - Effective nuclear charge is less than the actual nuclear charge ($Z$) because it includes the electron repulsion
      - $Z_{\text{eff}} = Z - S$
    - For multi-electron atom, energy of the electron with the same n value increase in proportion to the increasing l value
      - The lower energy of lower l value is due to the ineffectiveness of electron shield or screen
• $Z_{\text{eff}}$ increase when move across any row of the periodic table
  o More positive in the nucleus with the same shielding
• The $Z_{\text{eff}}$ change far less when go down in column than going across the row
  o The effectiveness of the electron decreases as the size of the electron core increase

7.3: Sizes of Atoms and Ions

• Bonding atomic radius – The distance separating the nuclei of atoms when they are chemically bonded to each other
• Each element is assigned with their own bonding atomic radius
• Bonding atomic radius between the same element is half the distance between the nuclei of each atom
• Periodic Trends in Atomic Radii
  o In each column (group), the atomic radii increase from top to bottom
    ▪ Results from the increase in principle quantum number of the outer electrons
  o In each row (period), the atomic radii decrease from left to right
    ▪ Results of the increasing effective nuclear charge
• Periodic Trends in Ionic Radii
  o Cations are smaller than their parent atoms
  o Anions are larger than their parent atoms
  o For ions carrying the same charge, size increases as we move down a column in the periodic table
  o Isoelectronic series is a group of ions all containing the same number of electrons
    ▪ E.g. O$^{2-}$, F$^-$, Na$^+$, Mg$^{2+}$, Al$^{3+}$
    ▪ In the isoelectronic series, the element with smallest atomic number have the largest ionic radius
      • High atomic number = higher positive = more attraction towards the electrons

7.4: Ionization Energy

• Ionization energy - The minimum energy needed to remove an electron
  o First ionization energy, $I_1$, is the energy needed to remove the first electron thus $I_2$ is the energy to remove the second electron and so one
• Variation in Successive Ionization Energies
  o $I_1 < I_2 < I_3$
After every removal of electron, the positive charge is constant and so is more concentrated on the remaining electrons thus become harder and harder to remove

- Big increase in ionization energy from the valence electron to the core electron
  - When move to the inner shell, it’s closer to the nucleus thus the effective nuclear charge increase and so is the ionization energy
  - Support the idea that only valence electron are involve in chemical bonding and reactions because the inner electron is too tightly bound to be lost or shared

**Periodic Trends in First Ionization Energies**

- Within each row (period), I₁ generally increases with increasing atomic number
  - The effective nuclear charge increase while the atomic radius decrease
- Within each column (group), the ionization energy generally decreases with increasing atomic number
  - Atomic radius increase while effective nuclear charge increase gradually thus the attraction between the nucleus and electron decreases
- Decrease of ionization energy from group 2 to group 13 because the third valence electron must occupy the p subshell which was empty
  - The p orbital have higher energy than the s orbital
- Decrease in ionization energy from group 15 to group 16 because the repulsion of paired electrons in the suborbital
  - Each p orbital is occupy by a single electron first which minimize the electron repulsion

**Electron Configurations of Ions**

- When electron is removed, it’s remove from the occupied orbital with the largest principle quantum number
  - Goes from the greatest l to the least l value
- When electron is added, it’s added from the unoccupied or partially occupied orbital with the lowest principle quantum number

7.5: Electron Affinities

- Electron affinity – energy change that occurs when electron is added to a gaseous atom
  - Show how much the atom wanted the electron
  - Usually negative (energy release)
- There are some exceptions in which the electron affinity is positive
  - Noble gas have positive electron affinity because it requires the electron to be in a higher-energy subshell which is highly unfavorable
  - Beryllium and Magnesium have a positive electron affinity because to add the electron, the electron must be add to the currently empty p subshell which is higher in energy level
- Nitrogen group have a lower electron affinity than the elements beside them because have each orbital filled singly with an electron

7.6: Metals, Nonmetals, and Metalloids

- H is a nonmetal even though it’s at the top left of the periodic table
- The more an element exhibits the physical and chemical properties of metals, the greater its metallic character
- Metals
  - Shiny luster, conduct heat and electricity, generally malleable and ductile
  - All metal except mercury are solid at room temperature (25°C)
  - Melting point can be low or high
  - Tends to have low ionization energy thus tend to form cations relatively easy
  - Metal in group 13 to 17 formed ion by either losing just electron from p orbital or from both p and s orbitals
  - Compound of metals and nonmetals tend to form ionic substance
  - Most metal oxides are basic
    - The oxide ion in the metal oxide reacts with water and form base
- Nonmetals
  - Vary greatly in appearance
  - Not lustrous, poor conductor of heat and electricity, generally have lower melting point than metal
  - Under normal condition, seven nonmetal exist as diatomic molecules
    - H₂, N₂, O₂, F₂, and Cl₂ as gas
    - Br₂ as liquid
    - I₂ as a volatile solid
  - Tend to form anions because of their electron affinity
  - Compound with only nonmetals are usually molecular substances
  - Most nonmetal are acidic
    - Reacts with water and form acid
- Metalloids
  - Can have properties of both metals and nonmetals
    - Have some properties of metals while lacking others and likewise
7.7: Group Trends for the Active Metals

- **Group 1: The Alkali Metals**
  - Soft metallic solid
  - Have low densities and melting points
  - Form 1+ ion
  - Only exist as compound in nature
    - Highly reactive
  - React with hydrogen in its hydride ion form, $\text{H}^-$
  - React violently with water
    - Highly exothermic
    - Possibly cause flame or explosion
  - The heavier members of the groups is more reactive because of their weaker hold of the electron
  - React with oxygen
    - Can form oxide, peroxide, or even superoxide
  - Emit characteristic color when placed in flame

- **Group 2: The Alkaline Earth Metals**
  - Harder, more dense, and higher melting point that the alkali metals
  - Reactivity increase from top to bottom
  - Form 2+ ion
  - Emit characteristic color when strongly heated in flame
  - 99% of calcium in human is found in the skeletal system

7.8: Group Trends for Selected Nonmetals

- **Hydrogen**
  - Nonmetal that occurs as a colorless diatomic gas under most conditions
  - Can be metallic under tremendous pressures
  - Can form both cation and anion
  - Have high ionization energy

- **Group 16: The Oxygen Group**
  - There's change from nonmetal to metal
  - Oxygen, sulfur, and selenium are typical nonmetals while tellurium is metalloid and polonium is metal
  - Oxygen is a colorless gas at room temperature while the others are solid
    - $\text{O}_3$ is ozone
    - Can be formed from $\text{O}_2$ in electrical discharges, such as lightning storm
- Less stable than $O_2$
- $O_3$ absorb certain wavelength of UV light
- Is a powerful oxidizing agent
  - Oxygen has great tendency to attract electrons so usually present as $O^{2-}$
  - Oxygen could also form peroxide ($O_2^{2-}$) and superoxide ($O_2^-$)
    - Often react with themselves to produce and oxide and $O_2$
  - Sulfur usually exist as $S_8$, a yellow solid
    - $S_8$ is sometimes just written as S(s)
    - Sulfur also have tendency to attract electron, although not as much as oxygen thus form sulfides, which contain the $S^{2-}$ ion
      - Most sulfur in nature is present as metal sulfide

- **Group 17: The Halogens**
  - Comes from Greek *halos* and *gennao* which means “salt formers”
  - All are typical nonmetals
  - Melting and boiling points increase with increasing atomic number
  - Have highly negative electron affinities
  - Lighter members of the groups is more reactive

- **Group 18: The Noble Gases**
  - All monatomic
  - Very nonreactive
  - Have large first ionization energy
    - The ionization energy decrease from top to bottom
  - Was called inert gases because they were thought to completely be unable to form chemical compounds