Chapter 3: Elements, Atoms, Ions, and The Periodic table

- 3.1 The Periodic Law and The Periodic Table
    - Dmitri Mendeleev and Lothar Meyer - two scientists working independently developed the precursor to our modern Periodic Table.
    - They noticed that as you list elements in order of atomic mass, there is a distinct repetition of their properties.
    - **Periodic Law** - the physical and chemical properties of the elements are periodic functions of their atomic numbers.
    - **Period** - horizontal row. Labeled 1 - 7.
    - **Groups** (or families) - columns of elements.
    - **Representative Elements** - Group A elements
    - **Transition elements** - Group B elements
    - **Alkali metals** - Group IA
    - **Alkaline earth metals** - group IIA
    - **Halogens** - group VIIA
    - **Noble gases** - group VIIIA
    - **Metals** - elements that tend to lose electrons during chemical change, forming positive ions.
    - **Nonmetals** - a substance whose atoms tend to gain electrons during chemical change, forming negative ions.
    - **Metalloids** - have properties intermediate between metals and nonmetals.
- 3.2 Electron Arrangement and The Periodic Table
    - **Electron configuration** - describes the arrangement of electrons in atoms.
    - The electron arrangement is the primary factor in understanding how atoms join together to form compounds.
    - **Valance electrons** - the outermost electrons.
    - These are the electrons involved in chemical bonding.
    - **Valence Electrons**
        - For the representative elements:
            - The number of valance electrons is the group number.
            - The period number gives the energy level (n) of the valance shell.
        - Let's look at an atom of fluorine as an example.
        - Fluorine has 7 electrons in the n=2 level
    - **The Quantum Mechanical Atom**
        - DeBroglie (French physicist) determined that electrons not only are particles, but they have a wave nature as well. Wave-particle duality.
        - Heisenburg Uncertainty Principle - cannot know the location and the momentum of an electron in an atom
        - Erwin Schrödenger - developed equations that took into account the particle nature and the wave nature of the electrons.
    - **Schrödenger equations:**
        - equations that determine the probability of finding an electron in a specific region in space.
        - give us Principle energy levels (n = 1,2,3…)

    - sublevels or subshells (s, p, d, f) and
    - Orbitals (odd number by subshell).
  - **Principle Energy Levels**
    - n = 1, 2, 3, …
    - The larger the n, the higher the energy level and the farther away from the nucleus the electrons are.
    - The number of subshells in the principle energy level is equal to n.
    - in n=1, there is one subshell
    - in n = 2, there are two subshells
    - The maximum number of electrons that can be in a principle energy level is equal to $2(n)^{2.}$
    - n = 1 can hold $2(1)^2 = 2$ electrons
    - Let's look at the periodic table and see how these numbers match up.
  - **Subshells**
    - Subshells increase in energy as follows: s<p<d<f (based on shape)
    - Therefore, electrons in 3d subshell have more energy than electrons in the 3p subshell.
    - Note: when giving a subshell, also give the principle energy level with it.
  - **Orbitals**
    - Orbital - a specific orbit path of a subshell containing a maximum of two electrons.
    - The two electrons in the orbital spin in opposite directions.
    - When the orbital contains two electrons, the electrons are said to be paired.
    - Let's look at these orbitals closely
- **Electron Configuration and the Aufbau Principle**
  - Electron Configuration - the arrangement of electrons in atomic orbitals.
  - Aufbau Principle - helps determine the electron configuration
    - Electrons fill the lowest-energy orbital that is available first
    - *Remember s<p<d<f in energy*
  - **Rules for Writing Electron Configurations**
    - Obtain the total number of electrons in the atom
    - Electrons in atoms occupy the lowest energy orbitals that are available.
    - Fill them in the order depicted in the following figure.
    - Remember:
      - How many subshells are in each principle energy level?
      - There are n subshells in the n principle energy level.
      - How many orbitals are in each subshell? s = 1, p = 3, d = 5, f = 7
      - s has 1, p has 3, d has 5, and f has 7
      - How many electrons fit in each orbital? 2
  - **Abbreviated Electron Configurations**
    - Uses noble gas symbols to represent the inner shell and the outer shell is written after.
    - For example: Let's look at Aluminum
    - The full electron configuration is: $1s^2 2s^2 2p^6 3s^2 3p^1$.
    - Therefore, the configuration can be written: $[Ne]3s^2 3p^1$

- 3.3 The Octet Rule
  - The noble gases are extremely stable.
  - The stability is due to:
  - the 1s being full in Helium
  - the outer s and p subshells being full in the other noble gases (eight electrons)
  - Octet Rule - elements usually react in such a way as to attain the electron configuration of the noble gas closest to them in the periodic table.
  - Ion Formation and the Octet Rule
    - Metallic elements tend to form positively charged ions called **cations**.
    - Metals tend to lose all their valance electrons to obtain a configuration of a noble gas.
    - Na^+ is "isoelectronic" with Ne
    - **Isoelectronic** - they have the same electron configuration (same number of electrons)
    - Nonmetallic elements tend to form negatively charged ions called **anions**.
    - Nonmetals tend to gain electrons so they become isoelectronic with its nearest noble gas neighbor.
    - The octet rule is very helpful in predicting the charges of ions in the representative elements.
    - Transition metals still tend to lose electrons to become cations but predicting the charge is not as easy. Transition metals often form more than one stable ion.
- 3.4 Trends in the Periodic Table
  - We will look at the following trends
    - in atomic size
    - in ionization energy
    - in electron affinity
  - **Atomic Size**
    - 1. The size of the atoms increases from top to bottom down a group.
      - This is due to the valance shell being higher in energy and farther from the nucleus.
    - 2. The size of the atoms decreases from left to right across a period.
      - This is due to the increase in magnitude of positive charge in the nucleus. The nuclear charge pulls the electrons closer to the nucleus.
  - **Ion Size**
    - Cations are always smaller than their parent atom.
      - This is due to more protons than electrons. The extra protons pulls the remaining electrons closer.
      - This size trend is also due to the fact that it is the outer shell that is lost.
    - Anions are always larger than their parent atom.
      - This is due to the fact that anions have more electrons than protons.
  - **Ionization Energy**

- **Ionization energy -** The energy required to remove an electron from an isolated atom.
- The magnitude of ionization energy correlates with the strength of the attractive force between the nucleus and the outermost electron.
- The lower the ionization energy, the easier to form a cation.
- Ionization increases across a period because the outermost electrons are more tightly held.
- Ionization decreases down a group because the outermost electrons are farther from the nucleus.

- **Electron Affinity**
  - **Electron Affinity -** The energy change when a single electron is added to an isolated atom.
  - Electron affinity gives information about the ease of anion formation.
  - Large electron affinity indicates an atom becomes more stable as it forms an anion.
  - E.A. generally decreases down a group.
  - E.A. generally increases across a period.