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**
  - \( n = 1, 2, 3, \ldots \)
  - 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: \( 1s^22s^22p^63s^23p^1 \).
  - Therefore, the configuration can be written: \([\text{Ne}]3s^23p^1\).
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.

  o **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.