<u>Chapter 5 - The Periodic Law</u>

5-1 History of the Periodic Table

I. <u>Mendeleev's Periodic Table</u>
   A. Organization
      1. Vertical columns in atomic weight order
         a. Mendeleev made some exceptions to place elements in rows with similar properties
            (1) Tellurium and iodine's places were switched
      2. Horizontal rows have similar chemical properties
   B. Missing Elements
      1. Gaps existed in Mendeleev's table
         a. Mendeleev predicted the properties of the "yet to be discovered" elements
            (1) Scandium, germanium and gallium agreed with predictions
   C. Unanswered Questions
      1. Why didn't some elements fit in order of increasing atomic mass?
      2. Why did elements exhibit periodic behavior?

II. <u>Moseley and the Periodic Table (1911)</u>
   A. Protons and Atomic Number
      1. Xray experiments revealed a way to determine the number of protons in the nucleus of an atom
      2. The periodic table was found to be in atomic number order, not atomic mass order
         a. The tellurium-iodine anomaly was explained
   B. The Periodic Law
      1. The physical and chemical properties of the elements are periodic functions of their atomic numbers

***Moseley was killed in battle in 1915, during WWI. He was 28 years old

III. <u>The Modern Periodic Table</u>
   A. Discovery of the Noble Gases
      1. 1868 - Helium discovered as a component of the sun, based on the emission spectrum of sunlight
      2. 1894 - William Ramsay discovers argon
      3. 1895 - Ramsay finds helium on Earth
      4. 1898 - Ramsay discovers krypton and xenon
      5. 1900 - Freidrich Dorn discovers radon
   B. The Lanthanides
      1. Early 1900's the elements from cerium (#58) to lutetium (#71) are separated and identified
   C. The Actinides
      1. Discovery (or synthesis) of elements 90 to 103
   D. Periodicity
      1. Elements with similar properties are found at regular intervals within the "periodic" table

5-2 Electron Configuration and the Periodic Table

I. <u>Periods and the Blocks of the Periodic Table</u>
   A. Periods
      1. Horizontal rows on the periodic table
      2. Period number corresponds to the highest principal quantum number of the elements in the period
   B. Sublevel Blocks
      1. Periodic table can be broken into blocks corresponding to *s, p, d, f* sublevels

II. <u>Blocks and Groups</u>
   A. *s*-Block, Groups 1 and 2
      1. Group 1 - The alkali metals
         a. One *s* electron in outer shell
         b. Soft, silvery metals of low density and low melting points
         c. Highly reactive, never found pure in nature
      2. Group 2 - The alkaline earth metals
         a. Two *s* electrons in outer shell
         b. Denser, harder, stronger, less reactive than Group 1
         c. Too reactive to be found pure in nature
   B. d-Block, Groups 3 - 12
      1. Metals with typical metallic properties
      2. Referred to as "transition" metals
      3. Group number = sum of outermost *s* and *d* electrons
   C. p-Block elements, Groups 13 - 18
      1. Properties vary greatly
         a. Metals
            (1) softer and less dense than d-block metals
            (2) harder and more dense than s-block metals
         b. Metalloids
            (1) Brittle solids with some metallic and some nonmetallic properties
            (2) Semiconductors
         c. Nonmetals
            (1) Halogens (Group 17) are most reactive of the nonmetals
   D. f-Block, Lanthanides and Actinides
      1. Lanthanides are shiny metals similar in reactivity to the Group 2 metals
      2. Actinides
         a. All are radioactive
         b. Plutonium (94) through Lawrencium (103) are man-made

5-3 Electron Configuration and Periodic Properties

I. <u>Atomic Radii</u>

A. Atomic Radius

1. One half the distance between nuclei of identical atoms that are bonded together

B. Trends

1. Atomic radius tends to decrease across a period due to increasing positive nuclear charge
2. Atomic radii tend to increase down a group due to increasing number energy levels (outer electrons are farther from the nucleus)

II. <u>Ionization Energy</u>

A. Ion

1. An atom or a group of atoms that has a positive or negative charge

B. Ionization

1. Any process that results in the formation of an ion

C. Ionization Energy

1. The energy required to remove one electron from a neutral atom of an element, measured in kilojoules/mole (kJ/mol)

$$A + \text{energy} \rightarrow A^{+} + e^{-}$$

D. Trends

1. Ionization energy of main-group elements tends to increase across each period
   a. Atoms are getting smaller, electrons are closer to the nucleus
2. Ionization energy of main-group elements tends to decrease as atomic number increases in a group
   a. Atoms are getting larger, electrons are farther from the nucleus
   b. Outer electrons become increasingly more shielded from the nucleus by inner electrons
3. Metals have a characteristic low ionization energy
4. Nonmetals have a high ionization energy
5. Noble gases have a very high ionization energy

E. Removing Additional Electrons

$$Na + 496\ \text{kJ/mol} \rightarrow Na^{+} + e^{-}$$
$$Na^{+} + 4562\ \text{kJ/mol} \rightarrow Na^{++} + e^{-}$$
$$Na^{++} + 6912\ \text{kJ/mol} \rightarrow Na^{+++} + e^{-}$$

1. Ionization energy increases for each successive electron
2. Each electron removed experiences a stronger effective nuclear charge
3. The greatest increase in ionization energy comes when trying to remove an electron from a stable, noble gas configuration

III. Electron Affinity

A. Electron Affinity

1. The energy change that occurs when an electron is acquired by a neutral atom, measured in kJ/mol
    a. Most atoms release energy when they acquire an electron

$$A + e^- \rightarrow A^- + \text{energy (exothermic)}$$

    b. Some atoms must be forced to gain an electron

$$A + e^- + \text{energy} \rightarrow A^- \text{ (endothermic)}$$

B. Trends

1. Halogens have the highest electron affinities
2. Metals have characteristically low electron affinities
3. Affinity tends to increase across a period
    a. Irregularities are due to the extra stability of half-filled and filled sublevels
4. Electron affinity tends to decrease down a group
5. Second electron affinities are always positive (endothermic)

IV. Ionic Radii

A. Cations

1. Positive ions
2. Smaller than the corresponding atom
    a. Protons outnumber electrons
    b. Less shielding of electrons

B. Anions

1. Negative ions
2. Larger than the corresponding atoms
    a. Electrons outnumber protons
    b. Greater electron-electron repulsion

C. Trends

1. Ion size tends to increase downward within a group

V. Valence Electrons

A. Valence Electrons

1. The electrons available to be lost, gained, or shared in the formation of chemical compounds
2. Main group element valence electrons are outermost energy level *s* and *p* sublevels

Group #	1	2	13	14	15	16	17	18
Number of valence Electrons	1	2	3	4	5	6	7	8

VI. <u>Electronegativity</u>

A. Electronegativity
   1. A measure of the ability of an atom in a chemical compound to attract electrons
   2. Elements that do not form compounds are not assigned electronegativities

B. Trends
   1. Nonmetals have characteristically high electronegativity
      a. Highest in the upper right corner
   2. Metals have characteristically low electronegativity
      a. Lowest in the lower left corner of the table
   3. Electronegativity tends to increase across a period
   4. Electronegativity tends to decrease down a group of main-group elements

VII. <u>Periodic Properties of the *d*- and *f*- Block Elements</u>

A. Atomic Radii
   1. Smaller decrease in radius across a period within the *d*- Block than within the main-group elements
      a. Added electrons are partially shielded from the increasing positive nuclear charge
      b. Slight increase at the end of the *d*-Block is due to electron-electron repulsion
   2. Little change occurs in radius across an *f*-block of elements

B. Ionization Energy
   1. Tends to increase across *d*- and *f*-Blocks

C. Ion Formation and Ionic Radii
   1. Electrons are removed from the outermost energy level *s*-sublevel first
      a. Most *d*-block elements form 2+ ions (losing 2 *s* electrons)
   2. Ions of *d*- and *f*-Blocks are cations, smaller than the corresponding atoms

D. Electronegativity
   1. Characteristically low electronegativity of metals
   2. Electronegativity increases as atomic radius decreases