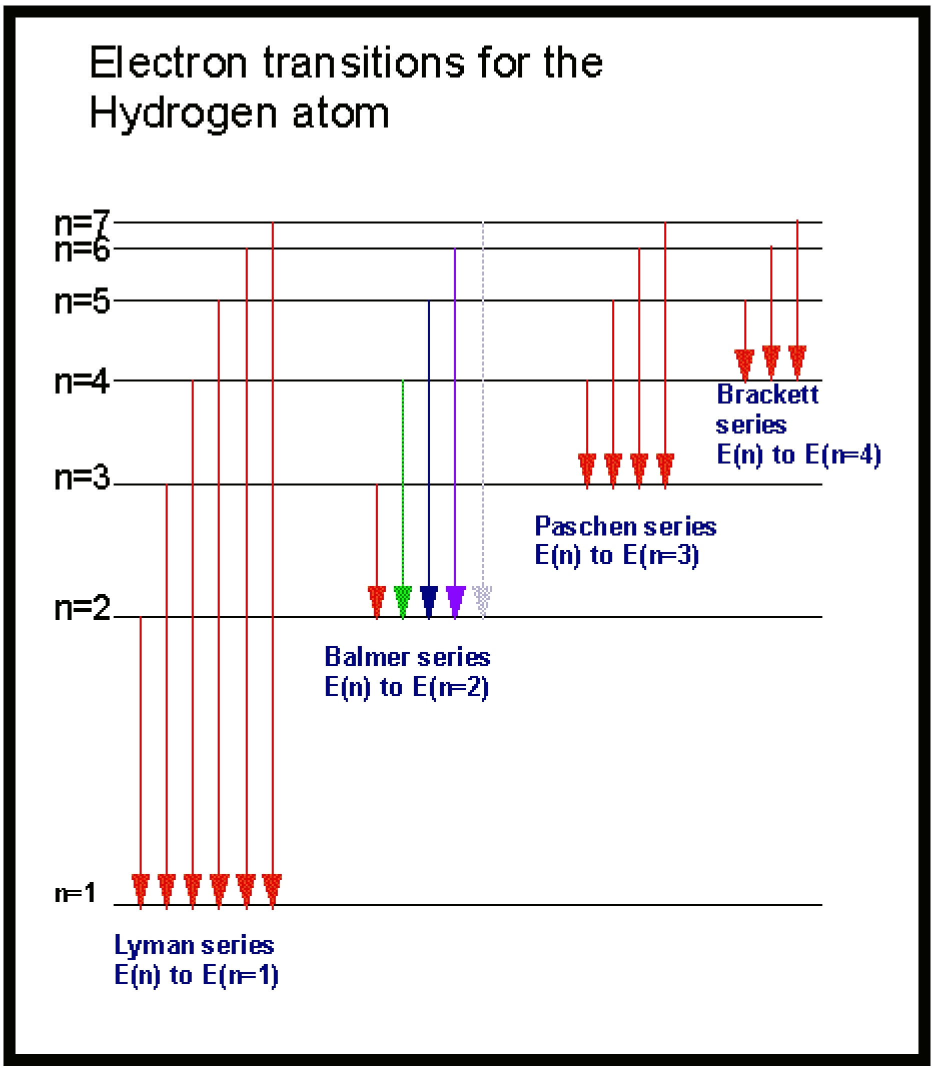
Unit 2 - Electrons and Periodic Behavior

I. The Bohr Model of the Atom

A. Electron Orbits, or Energy Levels
   1. Electrons can circle the nucleus only in allowed paths or orbits
   2. The energy of the electron is greater when it is in orbits farther from the nucleus
   3. The atom achieves the ground state when atoms occupy the closest possible positions around the nucleus
   4. Electromagnetic radiation is emitted when electrons move closer to the nucleus

B. Energy transitions
   1. Energies of atoms are fixed and definite quantities
   2. Energy transitions occur in jumps of discrete amounts of energy
   3. Electrons only lose energy when they move to a lower energy state



C. Shortcomings of the Bohr Model
   1. Doesn't work for atoms larger than hydrogen (more than one electron)
   2. Doesn't explain chemical behavior

4-2 The Quantum Model of the Atom

I. Atomic Orbitals and Quantum Numbers

Quantum Numbers specify the properties of atomic orbitals and the properties of the electrons in orbitals

A. Principal Quantum Number (*n*)
   1. Indicates the main energy levels occupied by the electron
   2. Values of n are positive integers
      a. n=1 is closest to the nucleus, and lowest in energy
   3. The number of orbitals possible per energy level (or "shell") is equal to n^2

B. Angular Momentum Quantum Number (*l*)
   1. Indicates the shape of the orbital
   2. Number of orbital shapes = *n*
      a. Shapes are designated *s, p, d, f*

C. Magnetic Quantum Number (*m*)
   1. The orientation of the orbital around the nucleus
      a. s orbitals have only one possible orientation

$$m = 0$$

      b. p orbitals have three, d have five and f have 7 possible orientations

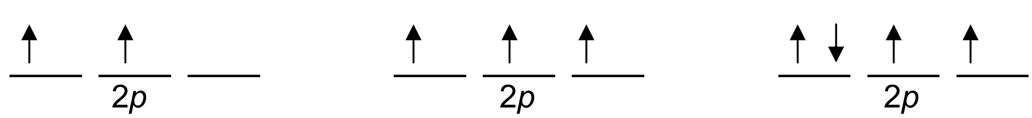
Principal Quantum Number (*n*)	Sublevels in main energy level (*n* sublevels)	Number of orbitals per sublevel	Number of electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

D. Spin Quantum Number
   1. Indicates the fundamental spin states of an electron in an orbital
   2. Two possible values for spin, +1/2, -1/2
   3. A single orbital can contain only two electrons, which must have opposite spins

4-3 Electron Configurations

I. Writing Electrons Configurations

A. Rules
   1. Aufbau Principle
      a. An electron occupies the lowest-energy orbital that can receive it
   2. Pauli Exclusion Principle
      a. No two electrons in the same atom can have the same set of four quantum numbers
   3. Hund's Rule
      a. Orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin



B. Orbital Notation
   1. Unoccupied orbitals are represented by a line, _____
      a. Lines are labeled with the principal quantum number and the sublevel letter
   2. Arrows are used to represent electrons
      a. Arrows pointing up and down indicate opposite spins

C. Configuration Notation
   1. The number of electrons in a sublevel is indicated by adding a superscript to the sublevel designation

      Hydrogen = $1s^1$
      Helium = $1s^2$
      Lithium = $1s^2 2s^1$

II. <u>Survey of the Periodic Table</u>

A. Elements of the Second and Third Periods
   1. Highest occupied energy level
      a. The electron containing energy level with the highest principal quantum number
   2. Inner shell electrons
      a. Electrons that are not in the highest energy level
   3. Octet
      a. Highest energy level *s* and *p* electrons are filled (8 electrons)
      b. Characteristic of noble gases, Group 18
   4. Noble gas configuration
      a. Outer main energy level fully occupied, usually (except for He) by eight electrons
      b. This configuration has extra stability

B. Elements of the Fourth Period
   1. Irregularity of Chromium
      a. Expected: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$
      b. Actual: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
   2. Several transition and rare-earth elements borrow from smaller sublevels in order to half fill larger sublevels

5-1 History of the Periodic Table

I. <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

II. <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. Periods and the Blocks of the Periodic Table

- 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. Blocks and Groups

- A. *s*-Block, Groups 1 and 2
  1. **Group 1 - The alkali meta**ls
     - 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. Atomic Radii

- 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. Ionization Energy

- 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

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

  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. <u>Electron Affinity</u>

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. <u>Ionic Radii</u>

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. Electronegativity

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