Unit 2 - Electrons and Periodic Behavior

5.1 Models of the Atom

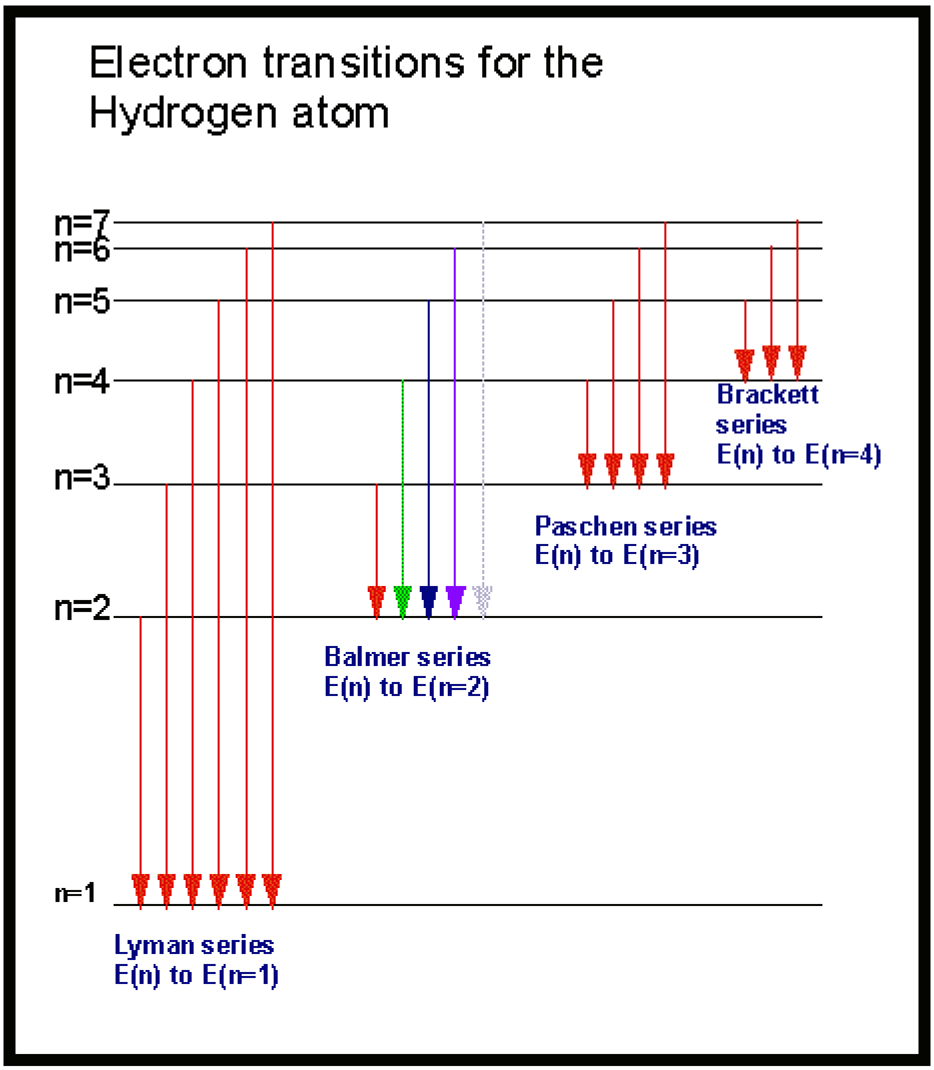
I. The Bohr Model of the Atom

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

II. The Quantum Mechanical Model

A. Probability and the Electron

1. The position and direction of motion of the electron cannot be simultaneously determined

   Translated: *"The more certain I am about where it is, the less certain I can be about where it is going. The more certain I am about where it is going, the less certain I can be about where it is."*

B. Regions of probability in which electrons may be found in an atom are determined by mathematical equations. These regions are called orbitals.

III. <u>Atomic Orbitals</u>

A. Atomic orbital

1. A region in space where there is a high probability of finding an electron

B. Energy Levels of electrons (n)

1. Indicates the distance of the energy level from the nucleus
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

C. Energy Sublevels

1. Indicates the shape of the orbital
2. Number of orbital shapes allowed in an energy level = *n*
   a. Shapes in the first four shells are designated *s, p, d, f*

Energy Level (*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. Electron Spin

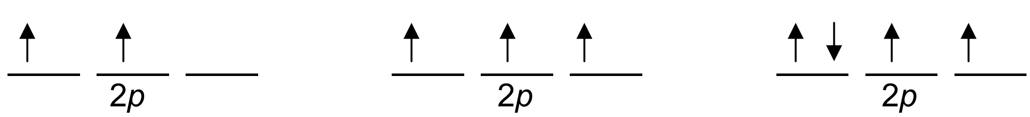
1. A single orbital can contain only two electrons, which must have opposite spins
2. Two possible values for spin, +1/2, -1/2

<u>5.2 Electron Arrangement in Atoms</u>

I. <u>Writing Electrons Configurations</u>

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



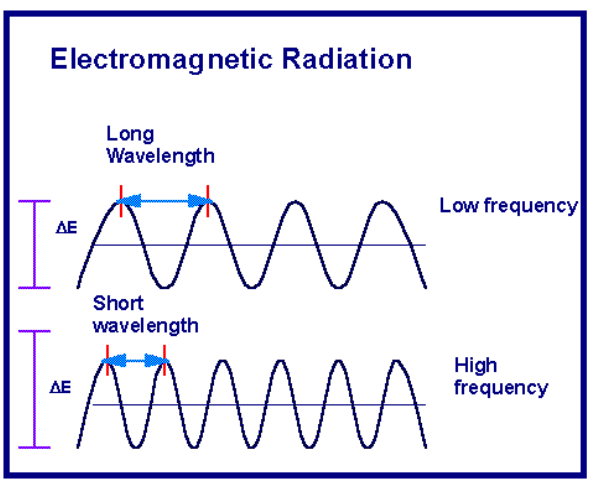
B. Orbital Notation
1. Unoccupied orbitals are represented by a line, _____
a. Lines are labeled with the principal quantum number and the sublevel letter
2. Arrows are used to represent electrons
a. Arrows pointing up and down indicate opposite spins
C. Configuration Notation
1. The number of electrons in a sublevel is indicated by adding a superscript to the sublevel designation

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

D. Exceptional Electron Configurations
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. Irregularity of Copper
c. Expected: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
d. Actual: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
3. Numerous transition and rare-earth elements transfer electrons from smaller sublevels in order to half-fill, or fill, larger sublevels

5.3 Physics and the Quantum Mechanical Model

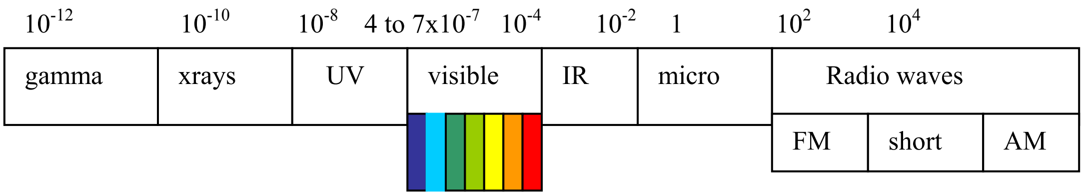
I. Properties of Light
A. Electromagnetic Radiation
1. Many types of EM waves
a. visible light
b. x-rays
c. ultraviolet light
d. infrared light
e. radio waves
2. EM radiation are forms of energy which move through space as waves
a. Move at speed of light
(1). 3.00×10^8 m/s
b. Speed is equal to the frequency times the wavelength $c = \nu\lambda$
(1). Freqency (ν) is the number of waves passing a given point in one second
(2). Wavelength (λ) is the distance between peaks of adjacent waves



c. Speed of light is a constant, so $\nu\lambda$ is also a constant
(1) ν and λ must be inversely proportional

B. Light and Energy

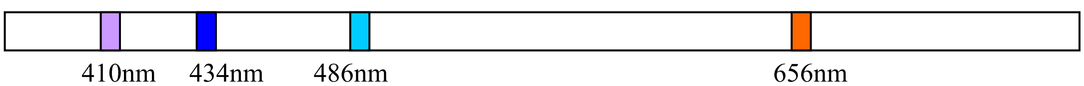
1. Radiant energy is transferred in units (or quanta) of energy called <u>photons</u>
    a. A photon is a particle of energy having a rest mass of zero and carrying a quantum of energy
    b. A quantum is the minimum amount of energy that can be lost or gained by an atom
2. Energy of a photon is directly proportional to the frequency of radiation
    a. $E = h\nu$ (h is Planck's constant, 6.62554×10^{-27} erg sec)



Wavelength increases →
Frequency decreases →
Energy decreases →

II. <u>Atomic Spectra</u>

A. Ground State
    1. The lowest energy state of an atom
B. Excited State
    1. A state in which an atom has a higher potential energy than in its ground state
C. Bright line spectrum
    1. Light is given off by excited atoms as they return to lower energy states
    2. Light is given off in very definite wavelengths
    3. A <u>spectroscope</u> reveals lines of particular colors



        a. Definite frequency
        b. Definite wavelength

6.1 Organizing the Elements

I. <u>The Periodic Law</u>

A. The physical and chemical properties of the elements are periodic functions of their atomic numbers
B. Elements on the table are arranged in order of increasing atomic number (number of protons)

II. <u>Metals, Nonmetals, and Metalloids</u>

A. Metals
    1. Good conductors of heat and electricity
    2. Lustrous (shiny)
    3. Solids (except mercury)
    4. Ductile (can be drawn into wire)
    5. Malleable (can be hammered into thin sheets)
B. Nonmetals
    1. Poor conductors of heat and electricity
    2. Most are gaseous
    3. Solids tend to be brittle
C. Metalloids
    1. Some properties of metals, some of nonmetals

6.2 Classifying the Elements

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

6.3 Periodic Trends

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. Trends in 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

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

   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. Trends in Ionic Size

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

IV. Trends in 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