<u>**Chapter 4 - Arrangement of Electrons in Atoms**</u>

The "Puzzle" of the nucleus:

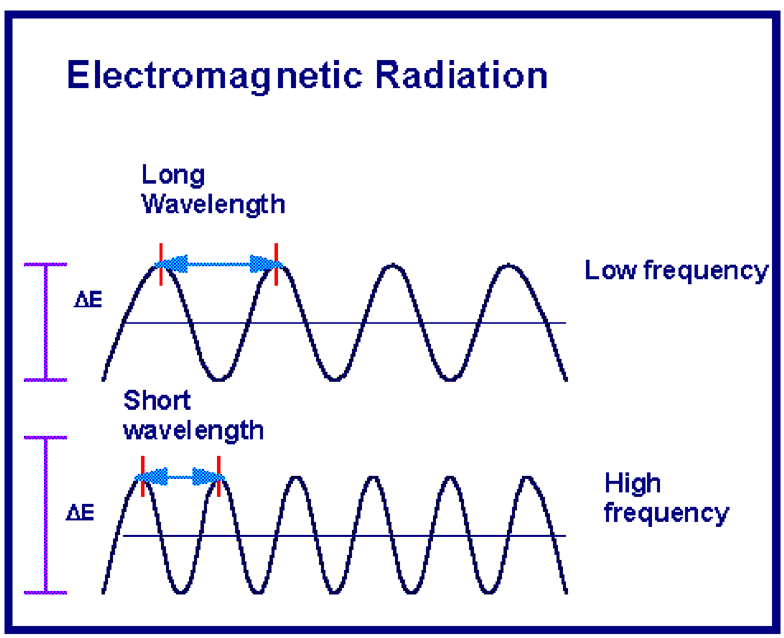
- Protons and electrons are attracted to each other because of opposite charges
- Electrically charged particles moving in a curved path give off energy
- Despite these facts, atoms don't collapse

4-1 The Development of a New Atomic Model

I. <u>Properties of Light</u>

A. Electromagnetic Radiation

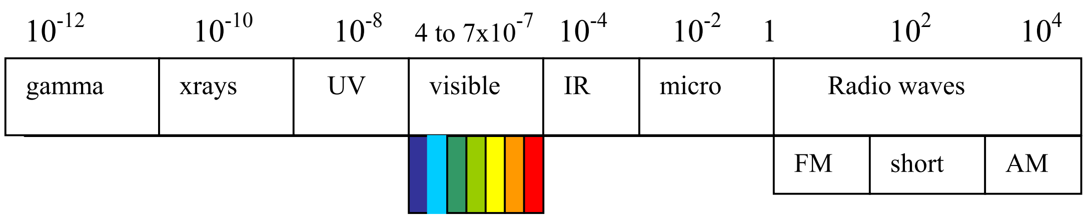
1. Many types of EM waves
   - a. visible light
   - b. x-rays
   - c. ultraviolet light
   - d. infrared light
   - e. radio waves
2. EM radiation are forms of energy which move through space as waves
   - a. Move at speed of light
     - (1). 3.00×10^8 m/s
   - b. Speed is equal to the frequency times the wavelength $c = \nu\lambda$
     - (1). <u>Freqency</u> (ν) is the number of waves passing a given point in one second
     - (2). <u>Wavelength</u> (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 - The Photoelectric Effect

1. The Photoelectric Effect
    a. Electrons are emitted from a metal when light shines on the metal
2. 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
3. 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 →

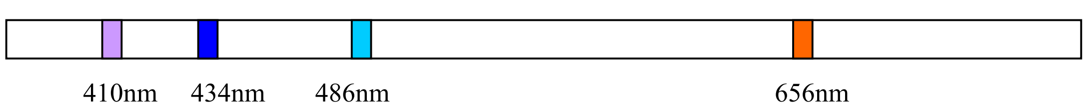
4. Wave-Particle Duality
    a. Energy travels through space as waves, but can be thought of as a stream of particles (Einstein)

II. <u>The Hydrogen Line Spectrum</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

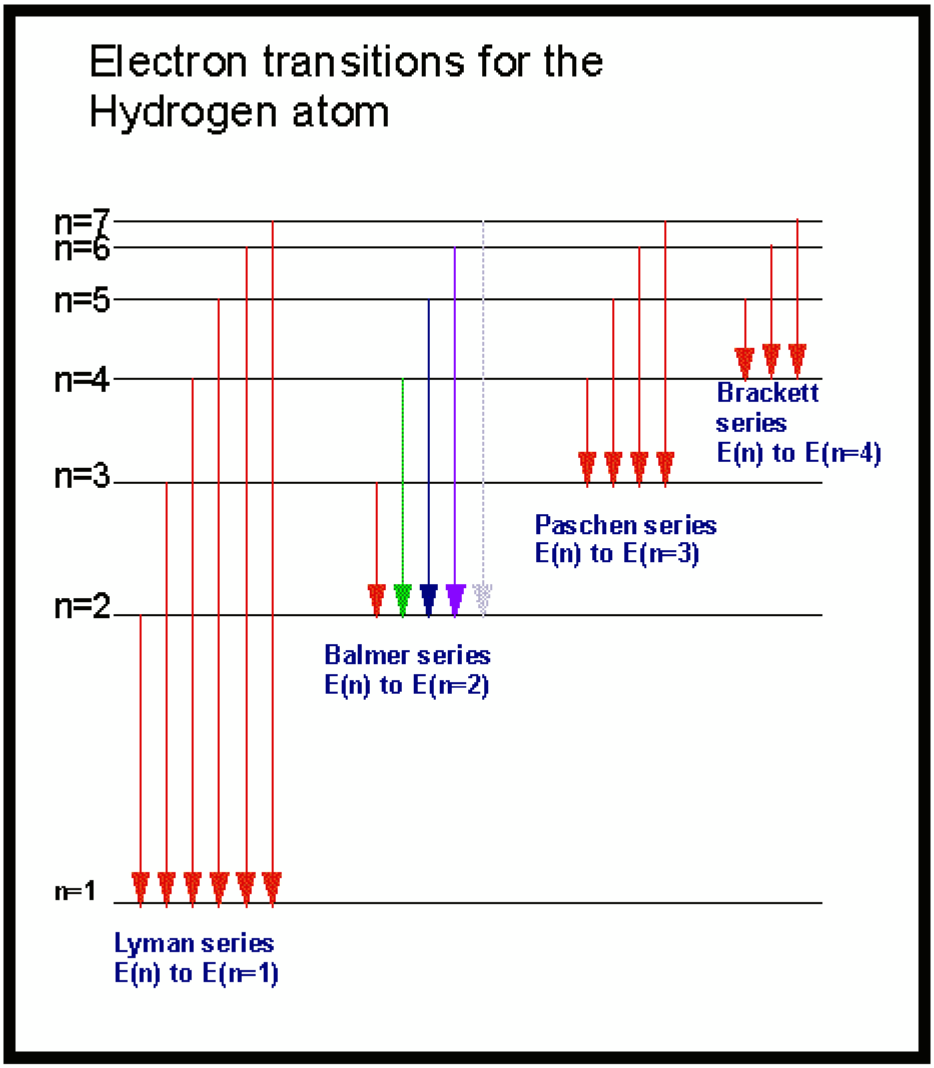


a. Definite frequency
b. Definite wavelength

$$\Delta E = h\nu = \frac{hc}{\lambda}$$

III. <u>The Bohr Model of the Atom</u>

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. Electrons as Waves and Particles

A. Louis deBroglie (1924)
   1. Electrons have wavelike properties
   2. Consider the electron as a wave confined to a space that can have only certain frequencies

B. The Heisenbery Uncertainty Principle (Werner Heisenberg - 1927)
   1. "It is impossible to determine simultaneously both the position and velocity of an electron or any other particle
      a. Electrons are located by their interactions with photons
      b. Electrons and photons have similar energies
      c. Interaction between a photon and an electron knocks the electron off of its course

C. The Schroedinger Wave Equation
   1. Proved quantization of electron energies and is the basis for Quantum Theory
      a. Quantum theory describes mathematically the wave properties of electrons and other very small particles
   2. Electrons do not move around the nucleus in "planetary orbits"
   3. Electrons exist in regions called orbitals
      a. An orbital is a three-dimensional region around the nucleus that indicates the probable location of an electron

$$-\frac{h^2}{8\pi^2 m}\frac{d^2\psi}{dx^2}+V\psi=E\psi$$

Schroedinger equation for probability of a single electron being found along a single axis (x-axis)

II. 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

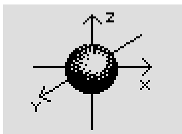
B. Angular Momentum Quantum Number (l)
   1. Indicates the shape of the orbital
   2. Number of orbital shapes = *n*
      a. Shapes are designated *s, p, d, f*

C. Magnetic Quantum Number (*m*)

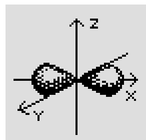
1. The orientation of the orbital around the nucleus
   a. s orbitals have only one possible orientation

      $m = 0$

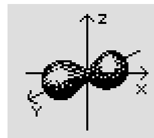
   b. p orbitals have three, d have five and f have 7 possible orientations



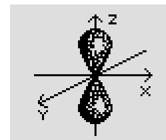
s orbital



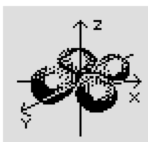
p_x orbital



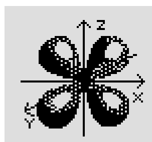
p_y orbital



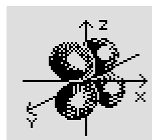
p_z orbital



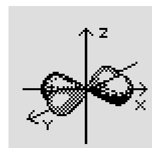
d_{xy} orbital



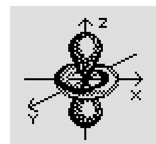
d_{xz} orbital



d_{yz} orbital



$d_{x^2-y^2}$ orbital



d_{z^2} orbital

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

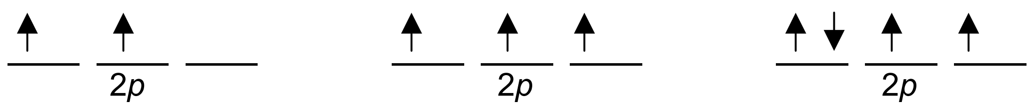
D. Spin Quantum Number

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

<u>4-3 Electron Configurations</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 = $1s^2 2s^1$

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