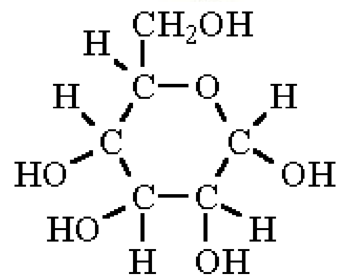
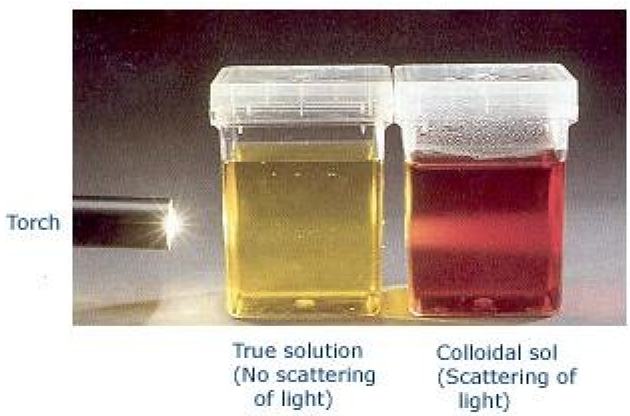
Unit 1 Notes - Matter & The Atom

1. Classification/Properties of Matter
   A. What is “matter?” Matter is “stuff.” Anything which has mass and takes up space.
   B. Classification of Matter
      a. Pure substances: One single substance, same properties throughout

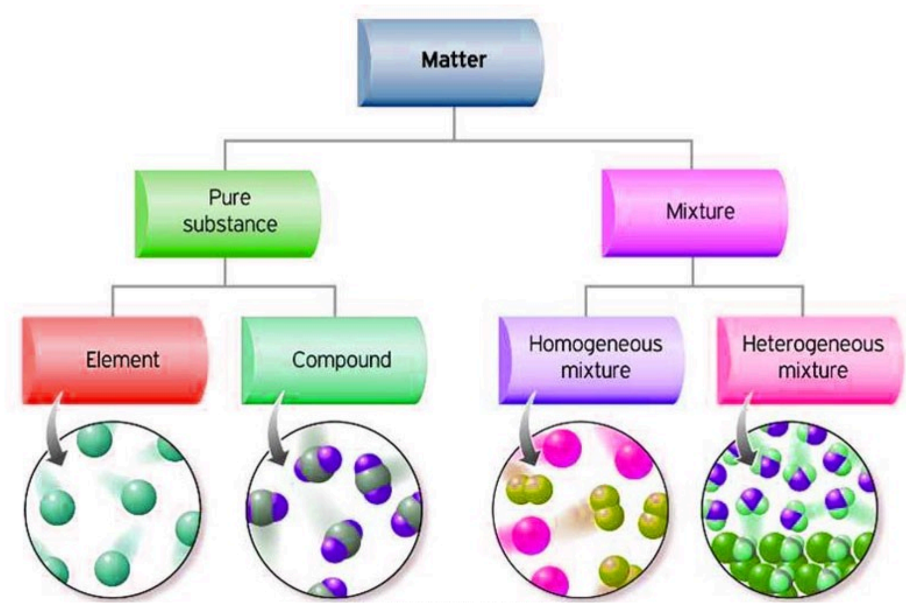
         i. Elements - simplest building blocks of matter, cannot be broken down further*. See periodic table for all 118.
            1. Examples: Carbon, Oxygen, Gold, Iron, Plutonium, Adamantium (joking…?), Krypton (not joking!)



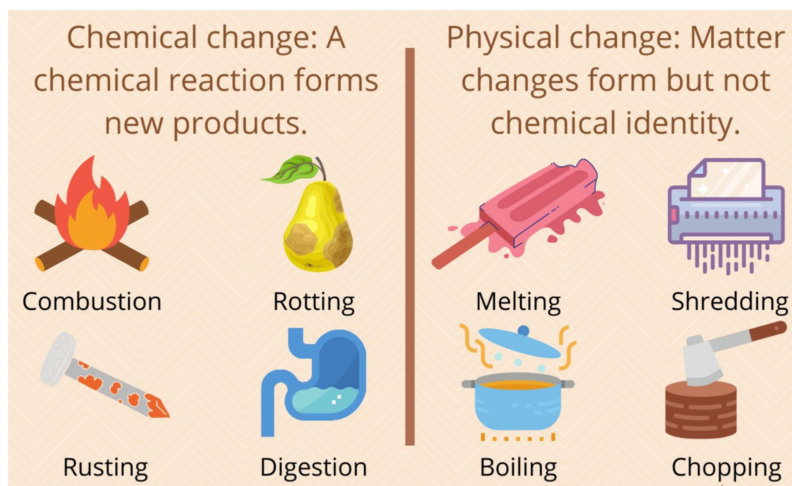
         ii. Compounds - two or more elements which have been chemically bonded together, can therefore be chemically separated
            1. Examples: Water (H_2O), Sodium Chloride ($NaCl$), Glucose ($C_6H_{12}O_6$)
      b. Mixture of substances - Two or more pure substances physically combined
         i. Heterogeneous mixtures - Can see the different (hetero) components, physically combined
            1. Suspensions: Large particle size, settles over time. Ex: river water
            2. Colloid: Medium particle size, does not settle. Can’t be seen with naked eye, but can using the Tyndall Effect



         ii. Homogeneous mixtures - Can not see different components, looks the same (homo) throughout. Again, physically combined, can therefore be physically separated.
            1. Solutions: Small particle size, does not settle. Synonym for homogeneous mixture, though (at least in chemistry) solvent is generally a liquid.

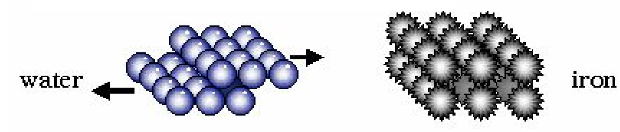


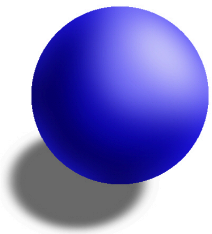
C. Properties of Matter
   a. Physical vs Chemical changes
      i. Physical changes - No new matter is formed. Properties/traits remain largely unchanged.
         1. Ex: Boiling, freezing, melting, dissolving, filtering, straining, tearing, bending
      ii. Chemical changes - New matter is formed. Properties/traits change
         1. Some indicators of chemical change -
            a. Temp change
            b. Light produced
            c. Precipitate forms
            d. Bubbles appear
            e. Color change
            f. Smell/taste change (not advised in chem lab)



   b. Physical vs chemical properties
      i. Physical properties - Characteristics of a substance observed without chemical change such as color, viscosity, luster, malleability. Often describe how a substance is, not how it acts.
         1. *Mass - How much matter is in a substance (grams)
         2. *Volume - Space occupied by a substance (cm^3, L)
         3. *Density - Mass per unit volume or "compactness" of a substance. D = m/v
         4. Temperature - Average KE of particles in a substance
         5. Thermal energy (Heat)
      ii. Chemical properties - Characteristics of a substance which describe its ability to undergo chemical changes. Often describe how they act/behave.
         1. Ex: Reactivity, flammability, acidic/basic, propensity to rust, decomposition, toxicity

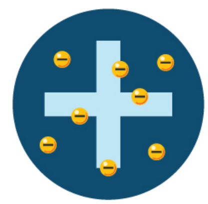
2. History of the Atomic Model
   - Model - A representation of a real object
   A. Democritus/Leucippus (400 B.C.E.) - Greek philosophers who thought there was a smallest piece of matter which they named “atomas” which means “uncuttable.”

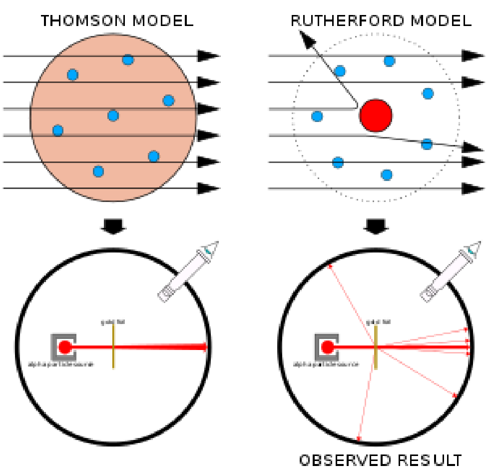




B. John Dalton (1766-1844) - English meteorologist/chemist had ideas similar to that of Democritus but made them more scientific & descriptive.
   a. Dalton’s Four Postulates
      1. All matter is made of atoms which are the smallest piece of matter.
      2. All atoms of an element are identical
      3. Compounds are whole number combinations of different atoms
         i. Law of definite proportions
         ii. Law of multiple proportions
      4. A chemical reaction is the rearrangement of atoms.
         Law of conservation of matter
   b. Dalton’s postulates today - which still hold up?

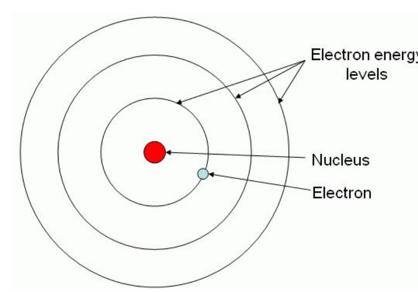
C. JJ Thomson (1856-1940) - Uses cathode ray tube to discover the electron (corpuscles). Suggested the “plum pudding” atomic model which had negatively charged electrons stuck in a positively charged spherical atom.

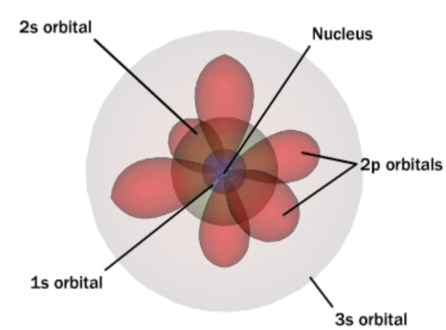




D. Ernest Rutherford (1871 - 1937) - Used [“gold foil experiment”]() to examine the composition of the atom. Determined that all an atom’s positive charge is in a dense region he called the “nucleus.”
   a. [Nucleus]() - Tiny, ultra dense center of the atom where an atom’s positive charge and nearly all the atom’s mass reside.
   b. Nearly all the rest of the atom is <u>empty</u>

E. Niels Bohr (1885-1962) - Used quantum physics and spectrum tubes to determine that electrons are contained in set rings or “orbits.”





F. Quantum Model - Uses wave functions to define 3D areas where electrons are likely/probable to exist.

3. Composition of the Atom

A. Parts of the atom

a. Proton

Mass: 1 amu*	Charge: +1	Location: Nucleus

*amu - atomic mass unit

i. Atomic number (Z) - The number of protons in an atom. Defines the identity of an atom and is used to arrange elements on the periodic table.

b. Neutron

Mass: 1 amu	Charge: neutral (0)	Location: Nucleus

i. Mass # (A) - The total mass of a given isotope. Equal to the # of protons + neutrons

ii. Isotopes - Versions of an element with different mass numbers due to different numbers of neutrons

Example: C-12 vs C-14

iii. Atomic mass - weighted average of all isotopes of an element, found on the periodic table

1. How to calculate: $\Sigma\ (a_i * m_i)_n$

c. Electron

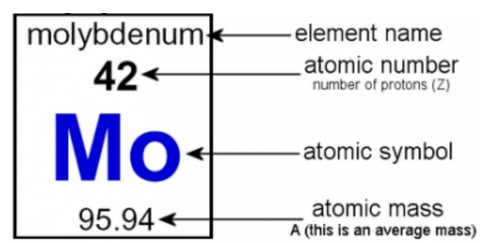
Mass: 1/2000 amu (no mass)	Charge: -1	Location: Orbits

i. Ion: An atom with a net overall charge due to an imbalance in protons & electrons

1. Cation - positive charge (+1, +2, etc), has *lost* electrons
2. Anion - negative charge (-1, -2, etc), has *gained* electrons

B. Determining #'s of protons, neutrons, electrons for an atom

a. Location of important information on the periodic table



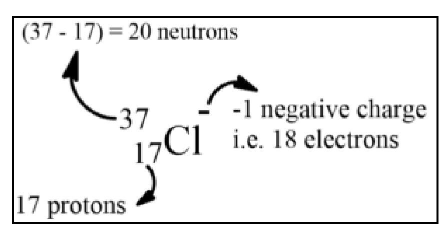
i. Most periodic tables minimally include the element's name, atomic number, atomic symbol, and atomic mass. Some include much more.

ii. Unfortunately, periodic tables do not standardize the locations of info in each element's box. Look for a key like the one shown here.

iii. Chemical names and symbols do not always match. Some elements' symbols match an older name (often Latin), not their modern name.

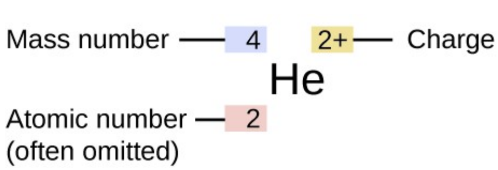
1. Example: Iron (originally named Ferrum), symbol Fe

b. Standard atomic symbol formats



i. The periodic table contains generic information about all versions of each element. For a specific isotope/ion/atom of an element, we write it one of two ways.

Standard (Nuclear) Notation vs. Hyphen notation



C-12 C-13 C-14

- The symbol for the element is written first.
- The number after the hyphen is the mass number.

4. Electrons in atoms

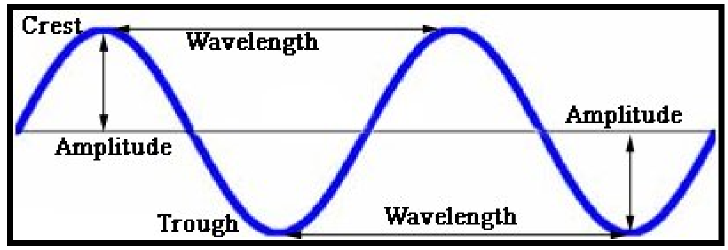
A. Electromagnetic radiation

a. $c = \lambda\nu$

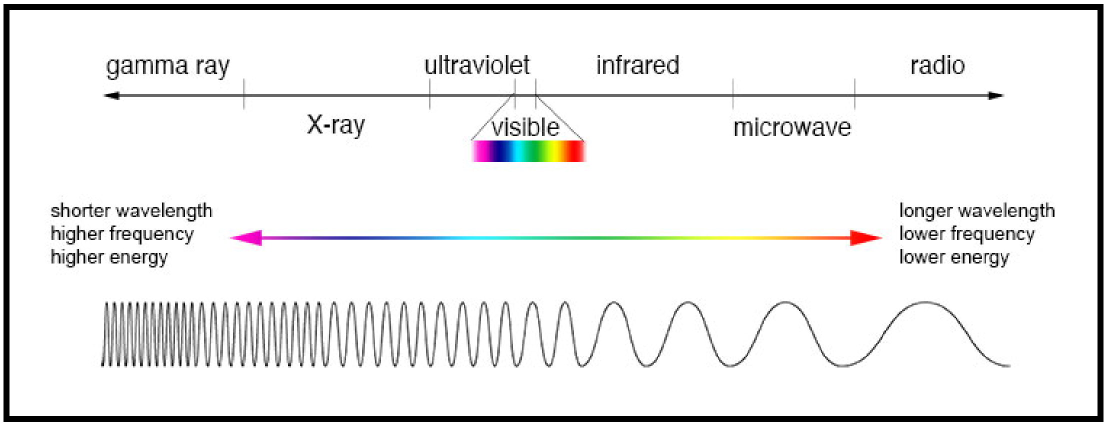
$c = 2.998 \times 10^8$ m/s

b. $E = h\nu$

h = Planck's constant = 6.626×10^{-34} J*s



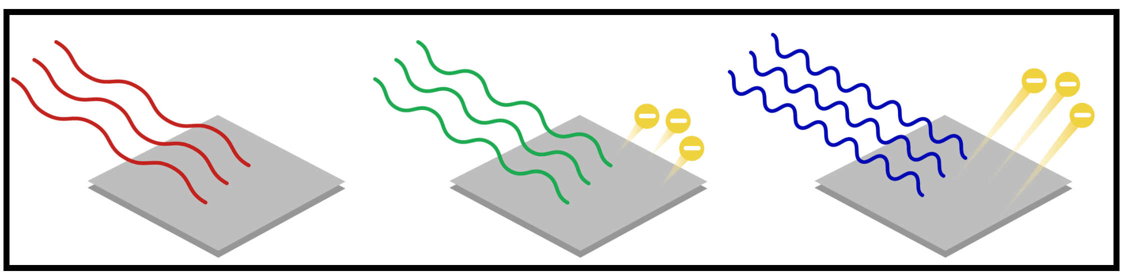
c. Electromagnetic spectrum



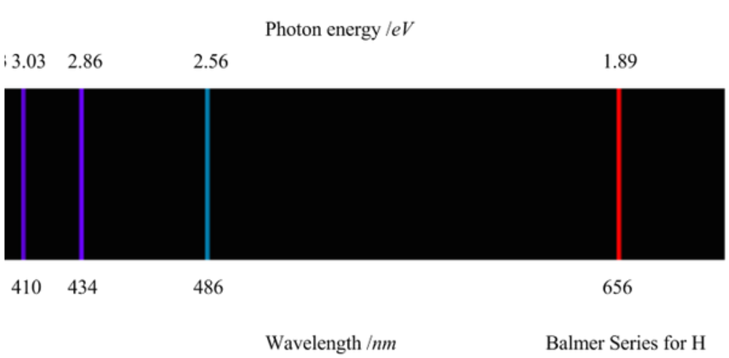
B. Bohr atomic model

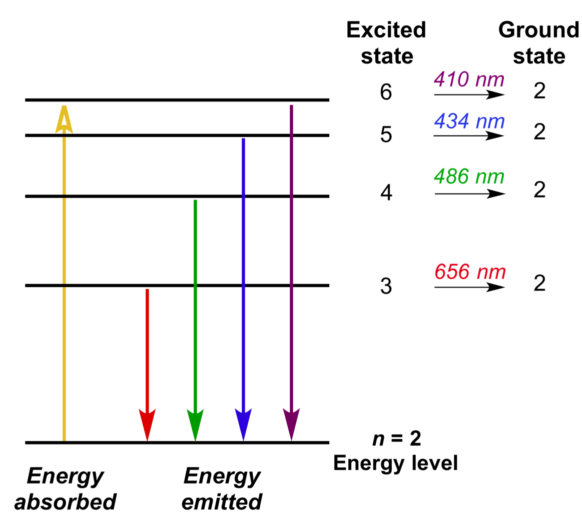
a. Photoelectric effect - electrons are emitted from substances when certain frequencies of light are shined upon them.

i. Photon- Smallest quantum of light.

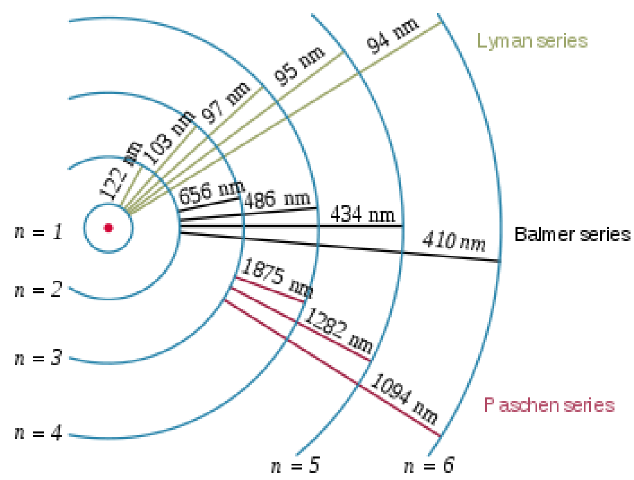


b. Emission spectra - When provided with energy (from electricity, fire, etc) elements emit light at unique wavelengths.



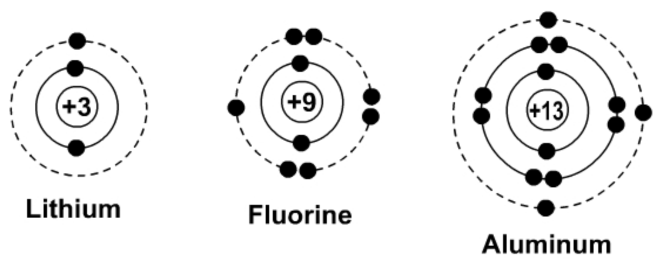


c. Bohr's model (see picture)



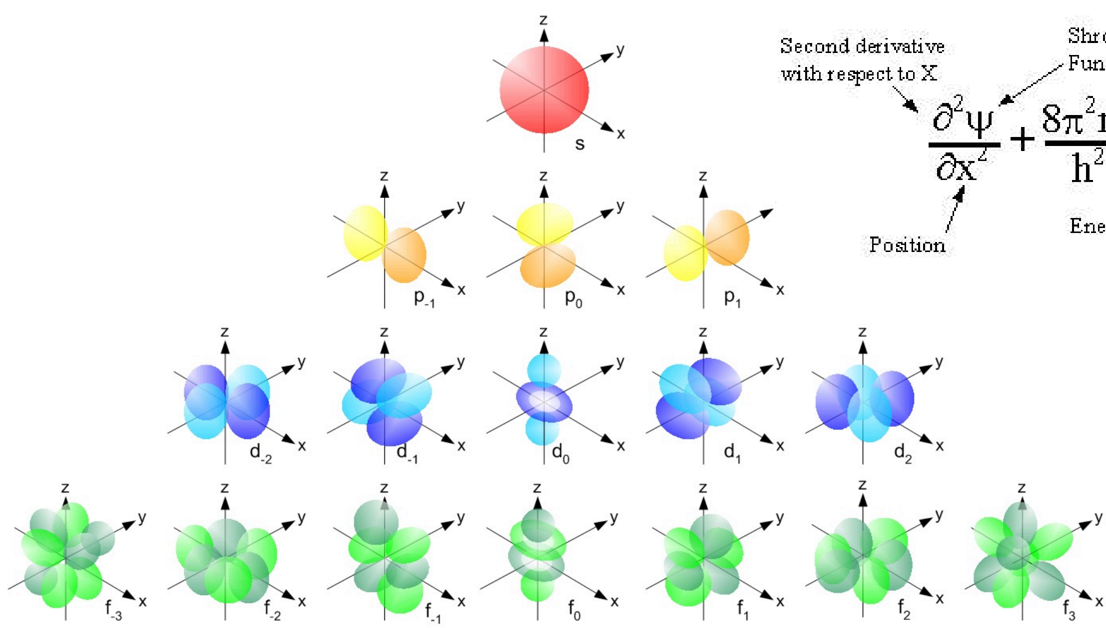
   i. Rules
      1. Electrons can only exist at distinct distances from the nucleus in energy levels (orbits).
      2. Max electrons per orbit = $2n^2$.
      3. Electrons fill from ground up, "preferring" to be closer to nucleus
      4. Electrons can absorb energy and "quantum leap" from their ground state to an excited state. When they drop back down to ground state, they release this energy as specific photons of EMR (colors of light, visible or otherwise).

   ii. Examples:



   iii. Issues
      1. Mathematical predictions only worked reliably for hydrogen
      2. Makes false predictions for atoms in the 4th row of periodic table and below
      3. Heisenberg uncertainty principle - cannot simultaneously know both an electron's position and it's velocity.

C. Quantum model - A mathematical model based upon the probabilities of where electrons are likely to be found around the nucleus (called orbitals).



Second derivative with respect to X

Shrodinger Wave Function

$$\frac{\partial^2 \psi}{\partial x^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

Position

Energy

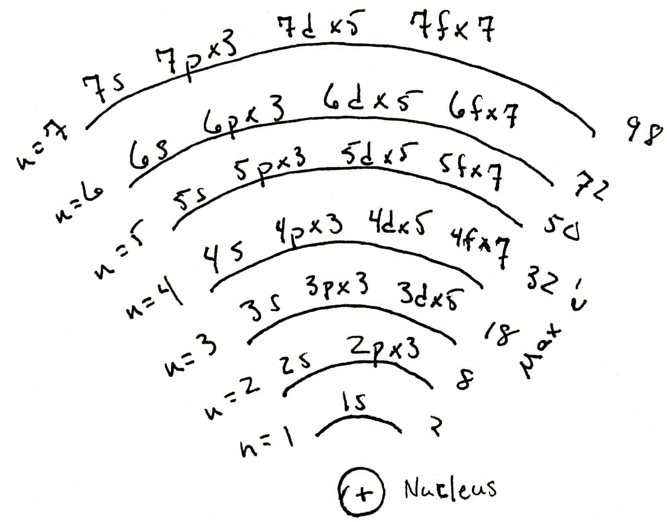
Potential Energy

a. Quantum Numbers: This is the math from which the 3D model (below) comes.

Name	Symbol	Allowed Values	Property
Principal	n	positive integers 1,2,3…	Orbital size and energy level
Secondary (Angular momentum)	l	Integers from 0 to (n-1)	Orbital shape (sublevels/subshells)
Magnetic	m_l	Integers $-l$ to $+l$	Orbital orientation
Spin	m_s	$+½$ or $–½$	Electron spin Direction

b. Structure of the quantum model atom:

   i. Max # of electrons per energy level = $2n^2$ (as in Bohr model)
   ii. Two electrons per orbital, one “up” spin, one “down” spin
   iii. One s orbital per energy level, starting with n=1
   iv. Three p orbitals per energy level, starting with n=2
   v. Five d orbitals per energy level, starting with n=3
   vi. Seven f orbitals per energy level, starting with n=4
   vii. Beyond this we get to g, h, and i orbitals, but these are never filled

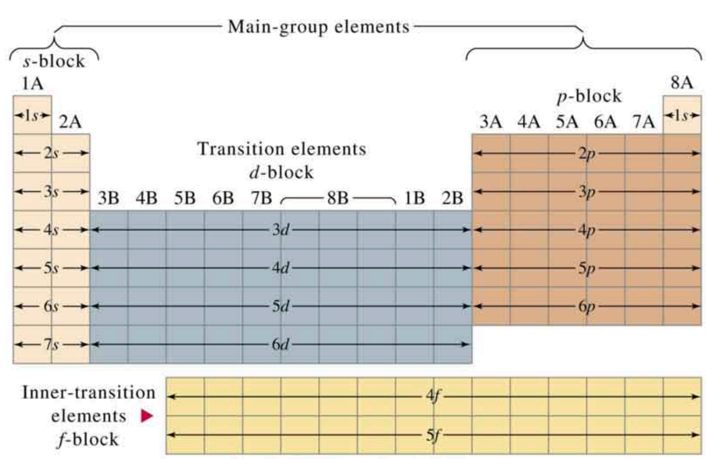


c. Rules for how electrons fill the quantum model:

   i. Aufbau Principle - “building up.” Electrons fill from low energy to high energy. This energy is determined by the interplay between pairing energy and promotion energy.
      1. Promotion energy - Energy determined by energy level an electron occupies.
      2. Pairing energy - Energy determined by # of electrons on the same energy level.
   ii. Pauli Exclusion Principle - No two electrons can have the same 4 quantum #’s
   iii. Hund’s Rule - Electrons fill degenerate orbitals singly first, then pair.

d. Electron configurations: A “ledger” showing where all the electrons in a quantum atom are. Written in the order <u>*in which electrons fill.*</u>

   i. Examples:
      Sodium: $1s^22s^22p^63s^1$
      Iron: $1s^22s^22p^63s^23p^64s^23d^6$
   ii. Periodic table is organized into s, p, d, and f blocks, mirroring the quantum atom.
   iii. ECs can most easily be written by “walking” the periodic table, starting at top with hydrogen, down to element in question



e. Orbital diagrams: A “map” showing the locations of all the electrons in an atom. Shows actual location, not fill order.
    i. Different from EC in that:
        1. Electrons & orbitals are always grouped by energy level, despite e^- fill order.
        2. Every electron is shown as a “spin up” or “spin down” arrow
        3. Every orbital is shown, even when skipped & left empty
    ii. Examples - See image at right