1) Dalton’s Atomic Theory and Law of Multiple Proportions p 82 – 85
   a) Law of Multiple Proportions states that if 2 elements combine to form different compounds, the ratio of the masses of the second element that react with a fixed mass of the first element will be a simple, whole-number ratio.
      i) Example: 12 g C + 32 g Oxygen = 44 g CO₂ and 12 g C and 16 g Oxygen = 28 g CO (carbon monoxide) which is a completely different compound. The “fixed” amount of the first element is 12 g for both compounds. The ratio of oxygen in CO₂ to oxygen in CO is 32:16 or 2 to 1. A simple, whole-number ratio.
      ii) Example: 36 g C + 8 g Hydrogen = 44 g Propane and 36 g C + 12 g Hydrogen = 48 g Methane. We have same 2 elements reacting in different amounts to make different compounds. The ratio of hydrogen in methane to hydrogen in propane is 12:8 or 3:2, a simple whole-number ratio.
      iii) This law is the main evidence for Dalton’s Atomic Theory
   b) Dalton’s Atomic Theory
      i) Combined Law of Mult Proportions and Law of Mass Conservation
      ii) Purpose was to better understand the nature of matter
      iii) Theory
         (1) **All elements composed of small, indivisible particles called atoms
         (2) **All atoms of same element have exactly same properties
         (3) Atoms of different elements have different properties
         (4) Compounds are formed when atoms join together and form in simple, whole-number ratios
            **Assumption 1 has been disproven under specific circumstances, although not in chemistry
            **Assumption 2 is not quite right as certain isotopes of an element are heavier than others

2) Atomic Structure p 201 – 208
   a) Read the history, but I won’t quiz or test you on it
   b) Electrical Charge
      i) Two types of charge—positive and negative
      ii) Absence of overall charge—electrically neutral
      iii) Like charges repel; unlike charges attract
      iv) Electrically neutral means that the sum of the positive and negative charges = 0.
      v) If the sum of the charges is not 0, then the object has a positive or negative charge.
   c) Atoms
      i) Nucleus
         (1) Contains positively charge protons and neutral neutrons
         (2) Nucleus doesn’t fly apart even though the protons must repel each other; a Strong force holds nucleus together
      ii) Electrons
         (1) Negatively charged
         (2) Outside of the nucleus
         (3) Attracted to the protons in the nucleus
   d) Determining the number of protons and electrons in an atom
      i) Z number—the number of protons in the nucleus
      ii) Called the Atomic number
      iii) All neutral atoms have the same number of protons and electrons
         (1) Sulfur has atomic number of 16. Thus, there are 16 protons and 16 electrons.
      iv) Mass number—total number of neutrons and protons in an atom
      v) Elements are made of isotopes—atoms with same number of protons but different # of neutrons
(1) Carbon has 3 isotopes: C-12 (6 p and 6 n), C-13 (6 p and 7 n) and C-14 (6 p and 8 n). notice that all isotopes have the same number of protons—they have to! They are all carbon and it is the proton number that defines the atom.

(2) Example: An atom has 10 p, 10 e, and 11 n. What is the atom? The only impt info in this question is the 10p. 10 protons tells us that the atomic number is 10 and on the Periodic Table that is Neon. Specifically we can say Ne-21 (this identifies the isotope, too)

3) Atomic Mass, AMU’s, Avogadro’s number, Mole, Molar Mass p 142 – 152

a) Atomic mass increases as you go across the PT and as you go down the groups. But what are the units?

   i) AMU—atomic mass unit which is equal to 1.66 x 10^{-24} g
      (1) Very small
      (2) Allows us to determine the mass of certain atoms
      (3) The atomic mass on the PT is not in grams, it is in amu’s which we can convert to grams. We will learn an easier relationship later
         (a) Example: What is the mass of a fluorine atom in grams?
             F has atomic mass of 19 amu. We set up a t grid! You are given 19 amu and want to convert it to g.

             \[
             \begin{array}{ccc}
             19 \text{ amu F} & 1.66 \times 10^{-24} \text{ g F} & = 3.15 \times 10^{-23} \text{ g F; this is the weight of 1 individual F atom} \\
             \hline
             1 \text{ amu F} & \hline
             \end{array}
             \]

   b) Molecular mass—mass, in amu’s, of a molecule
      i) To determine a molecules mass, simply add up the individual masses of the atoms
         (1) Example: H$_2$O  H—1.0 amu x 2 atoms = 2 amu, O—16.0 amu x 1 atom = 16 amu. Therefore, 2 amu + 16 amu = 18 amu. The molecular mass of water is 18 amu.
         (2) Example: H$_2$SO$_4$
             (a) H—1.0 amu x 2 = 2.0 amu
             (b) S—32 amu x 1 = 32 amu
             (c) O—16.0 amu x 4 = 64 amu
             (d) Total molecular mass: 2 + 32 + 64 = 98 amu

      ii) To determine the molecular mass in grams, convert your amu answer to grams using the t-grid.

   c) Mole Concept  p 145
      i) A Mole is a unit of measure in chemistry that is equal to Avogadro's number of particles. In every 1 mole, there will always be 6.022 x 10^{23} particles. Regardless of the material.
      ii) 1 mole of silver is 6.022 x 10^{23} atoms of silver
      iii) 1 mole of water is 6.022 x 10^{23} molecules of water
      iv) Each mole will have a different weight based on the matter it represents. This is where the mass comes in.
      v) The atoms’ mass in amu tells you how many grams it takes to have 1 mole of these atoms.
      vi) A molecule’s mass in amu tells you how many grams it takes to have 1 mole of these atoms
         (1) Example: How many moles of fluorine are there in a 50.0 g sample of fluorine? T-grid!
             (a) First, 1 mole F = 19.0 g F (based on v above)
                \[
                \begin{array}{ccc}
                50.0 \text{ g F} & 1 \text{ mole F} & = 2.63 \text{ moles F} \\
                \hline
                19.0 \text{ g F} & \hline
                \end{array}
                \]
         (2) Example: How many moles of HCl are contained in a 800.0 g sample:
            (a) First, find the molecular weight of HCl: (1.0 x 1) + (35.5 x 1) = 36.5 g = 1 mole
                \[
                \begin{array}{ccc}
                800.0 \text{ g HCl} & 1 \text{ mole HCl} & = 21.92 \text{ mol HCl} \\
                \hline
                36.5 \text{ g HCl} & \hline
                \end{array}
                \]
            (4) 3 mol Na = ? atoms of Na
                \[
                \begin{array}{ccc}
                6.022 \times 10^{23} \text{ atoms} & 3 \text{ mol Na} & = 9 \times 10^{23} \text{ atoms Na} \\
                \hline
                1 \text{ mol Na} & \hline
                \end{array}
                \]
(3) Example: What is the mass of a NO sample if it contains 0.15 moles? Molecular weight of NO is (14 x 1) + (16 x 1) = 30 g = 1 mol

\[
\begin{array}{c|c|c}
0.15 \text{ mol NO} & 30 \text{ g NO} & 4.5 \text{ g NO} \\
\hline
1 \text{ mol NO} & 
\end{array}
\]

4) Bohr Model of the Atom p 224 - 228
a) The ability of atoms to emit individual wavelengths of light helped Bohr create his model; all light travels at the same speed
b) Used a planetary model a central nucleus (containing protons and neutrons) with electrons in orbits around it
c) Electrons can jump into and out of specific orbits, but cannot be anywhere in-between orbits
i) To jump from one orbit to an orbit farther away from nucleus, electron must absorb energy—excited
ii) To drop down to an orbit closer to the nucleus the electron must get rid of extra energy and it does this by emitting light—de-excited
d) Bohr’s model has been updated and we have a new model

5) Quantum Mechanical Model of the atom p 228
a) Bohr’s assumptions we still use
i) Electrons orbit the nucleus
ii) There are many different orbits the electrons jump between
iii) Electrons need energy to go into orbits farther from nucleus
iv) Electrons release energy to drop down into orbits closer to the nucleus
v) There are only certain energies that an electron in an atom can have.
b) Additions
i) Electrons do not orbit in fixed circles, but in orbital clouds
   (1) Last paragraph on p 228 gives good analogy
ii) Orbitals are not all shaped the same
   (1) 4 different shapes (s, p, d, f)
c) 4 Quantum Numbers: numbers that specify the properties of atomic orbitals and their electrons
i) Principal Quantum Number: “n”—indicates the main energy levels surrounding a nucleus (shells)
   (1) These are the “orbits” to Bohr
   (2) Whole numbers only
   (3) Correspond to the periods on the Periodic Table
   (4) N = 1 is closest to the nucleus
   (5) As N increases, the distance of the main energy levels increases and so does their energy
ii) Orbital Quantum Number: “\ell” indicates the shape of the orbital
   (1) After the first main energy level, orbitals with different shapes occupy the different regions
   (2) Regions are called sublevels or subshells
   (3) To find “\ell”, simply take n -1. For 1st energy level N = 1, so S orbital would be 0. For N = 2, S sublevel is 0 and P sublevel is 1. For N = 3, there are 3 possibilities: S is 0, P is 1 and D is 2. For N = 4, S is 0, P is 1, D is 2 and D is 4.
   (4) The number of sublevels (or possible shapes) in each main energy level is = to the value of the principal quantum number (n)
   (5) “s” orbital—lowest energy; “p” orbital—higher energy; “d” orbital—still higher energy; and “f” orbital—highest energy
   (a) S orbital: spherical, only one sublevel in the S orbital
   (b) P orbital: dumbbell shapes oriented on the x-axis, y-axis, z-axis
   (c) D orbital: 4 lobed (mostly))
(d) F orbital: too complex to worry about shape

iii) Magnetic Quantum number: “mᵋ” indicates the orientation of an orbital about the nucleus
(1) S orbital—has only 1 orientation; a sphere with nucleus as the center; there is only 1 S sublevel for each S orbital; mᵋ = 0
(2) P orbital—3 different orientations; Pₓ, Pᵧ, Pᵦ; each p sublevel has 3 different p orbitals; mᵋ = ±1, 0, 1
(3) D orbital—5 different orientations; mᵋ = ±2, ±1, 0, 1, 2
(4) F orbital—7 different orientations; mᵋ = ±3, ±2, ±1, 0, 1, 2, 3

iv) Spin Quantum Number—“mₛ” only two possible values, +1/2 and −½, which indicates two possible states of an electron in an orbital; mₛ = +1/2 or −1/2
(1) Each orbital can only hold 2 electrons which must have opposite spins

v) Rules to know before writing Quantum numbers and electron configurations
(1) Aufbau Principle: an electron occupies the lowest energy level orbital that can receive it
(2) Hund’s Rule: orbitals of equal energy are each occupied by one electron before any one orbital is occupied by a second electron, and all single occupied orbitals must have same spin
(3) Pauli Exclusion principle: no 2 electrons in the same atom can have the same set of 4 quantum numbers
(4) Heisenberg Uncertainty principle: can’t know position and velocity of electron at same time

<table>
<thead>
<tr>
<th>Principal Quantum Number &quot;N&quot;</th>
<th>Orbital Quantum Number ( \ell )</th>
<th>Magnetic Quantum Number</th>
<th>Spin Quantum Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0</td>
<td>0</td>
<td>1 orientations</td>
</tr>
<tr>
<td></td>
<td>S</td>
<td></td>
<td>+1/2, -1/2</td>
</tr>
<tr>
<td>2</td>
<td>0,1</td>
<td>-1, 0, 1</td>
<td>3 orientations</td>
</tr>
<tr>
<td></td>
<td>S, P</td>
<td></td>
<td>+1/2, -1/2</td>
</tr>
<tr>
<td>3</td>
<td>0,1,2</td>
<td>-2, -1, 0, 1, 2</td>
<td>5 orientations</td>
</tr>
<tr>
<td></td>
<td>S, P, D</td>
<td></td>
<td>+1/2, -1/2</td>
</tr>
<tr>
<td>4</td>
<td>0,1,2,3</td>
<td>-3, -2, -1, 0, 1, 2, 3</td>
<td>7 orientations</td>
</tr>
<tr>
<td></td>
<td>S, P, D, F</td>
<td></td>
<td>+1/2, -1/2</td>
</tr>
<tr>
<td>5</td>
<td>0,1,2,3,4</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Example: What are the Quantum numbers for Li.
We have to start at the lowest energy level and fill it full of electrons before moving to the next energy level (Lithium is on N = 2). Starting with the first level, the only sublevel it contains is S and S can only hold 2 electrons, so 1s² is how we write this. There are no more sublevels in N = 1, so we move to N = 2. The lowest sublevel is S so we put in one electron. There are only 3 electrons and we have used all three electrons.

Orbital notation: \( \left( \begin{array}{c} \uparrow \downarrow \\ 1s^2 \end{array} \right) \left( \begin{array}{c} \uparrow \\ 2s^2 \end{array} \right) \) The quantum numbers are 2,0,0, ½ (this is the set for the last electron)

Example: What are the quantum numbers for fluorine? Fluorine has an atomic number of 9, so it has 9 electrons.

\( \left( \begin{array}{c} \uparrow \downarrow \\ 1s^2 \end{array} \right) \left( \begin{array}{c} \uparrow \downarrow \\ 2s^2 \end{array} \right) \left( \begin{array}{c} \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \\ 2p^5 \end{array} \right) \) (remember, p has 3 shapes and each holds 2 electrons)
The quantum number of the last electron (#9) is 2, 1, 0, ½

Example: Quantum numbers for phosphorus? Atomic number is 15, so we have 15 electrons to place before we can write the numbers.

\( \left( \begin{array}{c} \uparrow \downarrow \\ 1s^2 \end{array} \right) \left( \begin{array}{c} \uparrow \downarrow \\ 2s^2 \end{array} \right) \left( \begin{array}{c} \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \\ 2p^5 \end{array} \right) \left( \begin{array}{c} \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \\ 3s^2 \end{array} \right) \left( \begin{array}{c} \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \\ 3p^3 \end{array} \right) \) Quantum #: 3, 1, 1, ½
Notice under each orbital diagram are a series of letters and numbers. These are the electron configurations!
The electron configuration for phosphorus is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Example: Find the quantum numbers and electron configuration for Calcium, Ca. Atomic # is 20.

\[
\begin{array}{cccccc}
(\uparrow \downarrow) & (\uparrow \downarrow) & (\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) & (\uparrow \downarrow) & (\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) & (\uparrow \downarrow) \\
1s^2 & 2s^2 & 2p^6 & 3s^2 & 3p^6 & 4s^2 \\
\end{array}
\]
Quantum numbers: 4, 0, 0, -1/2
Electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Example: Find the quantum numbers and electron configuration for Nickel. Atomic number is 28.
This one is trickier. The “d” sublevel is out of place. It actually is higher in energy than you expect.

\[
\begin{array}{cccccc}
(\uparrow \downarrow) & (\uparrow \downarrow) & (\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) & (\uparrow \downarrow) & (\uparrow \downarrow)(\uparrow \downarrow)(\uparrow \downarrow) & (\uparrow \downarrow) \\
1s^2 & 2s^2 & 2p^6 & 3s^2 & 3p^6 & 4s^2 \\
\end{array}
\]
Quantum numbers: 3, 2, 0, -1/2
Electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ Notice that the 3d sublevel has more energy than the 4s!! The d sublevels will always be at (n-1).

6) Electron Configuration and the Periodic Table p 247 - 255
   a) Group 1 & 2 correspond to the S sublevel
      i) Notice that S can only hold 2 electrons and their are only 2 blocks horizontally. Hmmmm.
      ii) N = period number
   b) Group 3 – 12 correspond to the D sublevel
      i) Remember there are 5 shapes of D with 2 electrons in each shape for a total of 10 electrons and there are
         10 blocks horizontally. Hmmmm.
      ii) This group is always N-1. So, for period 4 then you would have 3d. for period 5, 4d. etc.
   c) Group 12 – 18 correspond to the P sublevel
      i) There are 3 shapes p can have and each shape can only hold 2 electrons for a total of 6 electrons. Hmmmm.
         There are 6 blocks horizontally per period.
   d) Lanthanides and Actinides—represent the F sublevels
      i) F sublevel has 7 orientations and each can hold 2 electrons for a total of 14 electrons. There are exactly 14
         blocks horizontally.
      ii) F sublevel is n – 2. So, for period 6, you would have 4f. Period 7 would be 5f.
   e) ALL YOU HAVE TO DO IS COUNT OFF THE BLOCKS!
   f) Valence electrons—electrons farthest from the nucleus and generally have highest energy level
      i) Determines the atoms chemistry
      ii) Atoms in same group (column) on the Table have the same valence electrons and similar chemistry
      iii) Noble gases—have a full octet (8 electrons in outer shell) so these do not like to react and are sometimes
          called Inert gases.
7) Lewis Structures
   a) Uses the valence electrons
   b) Also called electron dot diagrams
   c) Allows chemists to see the valence electrons and determine how it will react
Example: Write the Lewis structure for the following atoms:

Magnesium, Mg. First, determine the valence using the Periodic table. Mg is in Group 2 so it has 2 valence electrons!

\[
\bullet \text{Mg} \bullet
\]

Gallium, Ga. Valence on the Table is group 3, so 3 valence electrons.

\[
\bullet \text{Ga} \bullet
\]

Tellurium, Te. Valence on the Table is group 6, so 6 valence electrons.

\[
\bullet \text{Te} \bullet
\]

8) More on Lewis Structures  p 251-255  & p 263-272
   a) Ionic compounds (remember these are the ones made from a metal + a nonmetal)
      i) Nonmetals try to gain electrons to complete their octet (8 valence electrons)
      ii) Metals try to lose their electrons so their Lewis strux should have no dots extra
      iii) See the examples from the board
   b) Covalent molecules (nonmetal + nonmetal, sharing electrons rather than gain/lose)
      i) Atoms in covalent molecules share electrons in a covalent bond (2 shared electrons)
         (1) Shared pairs are represented by a dash
         (2) See rules p 267
         (3) Bonds can be double (sharing 4 electrons) or triple bonds (sharing 6 electrons)
         (4) See examples from the board

9) Periodic Properties p 258 – 276
   a) Periodic property: characteristic of atoms that varies regularly across the periodic chart
      i) Ionization Potential: amount of energy needed in order to take an electron away from an atom; increases as you go right on the table and up
         (1) Depends on electron configuration
            (a) Group 1A needs to lose 1 electron to have full octet, so takes very little energy to remove the electron compared to Group 7 which only needs 1 electron to be full (it is easier to gain 1 than lose 7 so group 7 holds its electrons tightly)
            (b) Within a group, increases as you go from bottom to top. It is easier to take one electron from francium (its valence electron is 7 levels from the nucleus so weak attraction) than to take 1 from Lithium (valence electron is in level 2 so much closer to nucleus and stronger attraction b/n protons and electrons)
      ii) Electronegativity: a measure of how strongly an atom attracts extra electrons to itself
         (1) Increases as you go to the right and as you go up
This is based on filling the octet. The closer the atom is to filling the octet, the more strongly it wants an electron. 

(a) The closer the valence electrons are to the nucleus, the stronger the attraction between protons and electrons.

iii) Atomic Radius: average distance from center of nucleus to outer edge (approximation at best)

1. Increases as you go left on the chart and down
2. The closer the electrons are to the nucleus, the stronger the attraction with the protons. This shrinks the overall size. And, the more protons in the nucleus, the stronger the pull on electrons.

10) Polyatomic ions p 285
   a) Ions which are formed when a group of atoms gains or loses electrons
   b) Study the ions on the handouts. These will help in the next unit!!!

11) VSEPR Theory p 289
   a) Valence Shell Electron Pair Repulsion
      i) Electrons being shared are called shared pairs
      ii) Electrons not being shared are called unshared electron pairs
   b) Molecules will attain whatever shape keeps the valence electrons of the central atom as far apart from one another as possible
   c) You will use your handout to work some problems as well as the rules on page 295

12) Purely Covalent and Polar Covalent bonds p 298
   a) Purely covalent bond—a covalent bond (sharing electrons) in which the electrons are shared equally between the atoms involved
      i) Each atom in a bonded pair wants the shared electrons for itself so they pull on the electrons. If the electronegativities are the same, then the atoms will share the electrons equally (think joint custody)
   b) Polar Covalent bond—a covalent bond in which the electrons are shared unequally between the atoms involved
      i) The atom with the higher electronegativity will get more “time” with the electrons. This will result in slightly charged areas of the molecule (the atom with the higher electronegativity will be slightly negative and the other atom will be slightly positive. This happens for a fraction of a time)
   c) Polar molecules must have:
      i) Polar bonds (unequally shared electrons)
      ii) Polar bonds cannot be of equal strength and in opposite directions
   d) Polar covalent compounds can dissolve other polar covalents and ionic compounds (both contain electrical charges)
   e) Purely covalent compound (nonpolar) can only dissolve other purely covalent compounds (because no charges are present)