Chem 115 Class Notes from 2/5/08 (kf)

THE ATOM

-They are made up of subatomic particles: **neutrons**, **protons**, and **electrons**. The atom's center holds the **nucleus** which consists of neutrons and protons while the electrons are located outside the nucleus in the **electron cloud** (see Model 2 under "What does an atom look like?" at bottom of page 24 of Lecture 3 class slides).

-Neutrons have no charge (0), protons have a positive charge (+), and electrons have a negative charge (-). *A neutral atom has an equal number of protons and electrons (neutrons have no effect on the charge).

-The nucleus contains most of an atom's mass but occupies very little space. Electrons have a lot less mass but take up most of the space in the atom (see Rutherford's experiment on alpha particles in book on page41-42).

-Bohr model: The atom has electron shells in which the electrons reside. The first shell can hold up to **2** electrons, the second shell can hold up to **8**, the third can hold up to **8**, and so on.

-An atom's mass is recorded in **atomic mass units (amu)**, One amu = 1.66054×10^{-24}g.

proton = 1.0073amu
neutron = 1.0087amu
electron = 5.486×10^{-4}amu

*Note: you can calculate the approximate mass of an atom by adding the protons and neutrons together because each of their value is very close to1.0amu. No need to include electrons because each has a mass of a lot less than 1.0amu.

ATOMIC NUMBERS, MASS NUMBERS, AND ISOTOPES

-**Atomic numbers** represent the # of protons in an atom. This # can be found in the periodic table above the element symbol. For example, hydrogen (H) has an atomic number of 1.

-The **mass number** signifies the # of protons and neutrons in an atom. Consider the example: ^{13}C. The thirteen is the mass #. The # of neutrons can be derived from this number by subtracting the atomic number from the mass number. Therefore, we can say there are 7 neutrons in this carbon (C) atom (13-6=7).

-An **isotope** is an atom with the same # of protons but different # of neutrons. There are three isotopes of C:

	# OF PROTONS	# OF NEUTRONS
^{12}C	6	6
^{13}C	6	7
^{14}C	6	8

- **atomic weight**, which is the average atomic mass, is found below an atomic symbol on the periodic table. This number represents the average atomic mass of all of a naturally occurring element's isotopes. Naturally occurring carbon is made up of 98.93% ^{12}C and 1.07% C^{13}. With masses of 12amu and 13.00335amu, respectively, we can calculate the atomic weight (this example is taken from page 47 in the book):

$$(0.9893)(12\text{amu}) + (0.0107)(13.00335\text{amu}) = 12.01\text{amu} = \text{atomic weight}$$