STRUCTURE OF THE ATOM

There are three subatomic particles that are important in Chemistry:

1. Electrons (negatively charged)
2. Protons (positively charged)
3. Neutrons (no charge)

There are two main sections of the atom:

1. Nucleus
   - centre of the atom
   - contains the protons and neutrons
   - very small (ant on a football field)
   - has almost all the mass
   - contains all protons (p^+)
   - contains all neutrons (n^0)

2. Area outside the nucleus:
   - where orbitals are found
   - makes up almost all the volume (football field)
   - has almost no mass (protons and neutrons are about 1840 times the mass of electrons)
   - where the electrons (e^-) are found

Descriptors of the Atom:

1. Atomic Number
   - refers to the number of protons (p^+) in the nucleus
   - never changes
   - if the atom is neutral, as all atoms are, this will also equal the number of electrons (e^-)

ATOMIC NUMBER = NUMBER OF PROTONS = NUMBER OF ELECTRONS

2. Atomic Mass
   - describes masses of atoms relative to carbon
   - used for large groups of atoms
   - average atomic mass is provided on periodic tables

3. Mass Number
   - number of protons <u>plus</u> the number of neutrons (total mass in nucleus)
   - mass numbers and numbers of neutrons are <u>not found on your periodic tables</u>
   - used to distinguish among **<u>isotopes</u>**

Isotopes

Isotopes are atoms that have the same number of protons, but different numbers of neutrons, than other atoms in the same element. All elements but four have more than one naturally occurring isotope

Since neutrons have significant mass, this means the entire atom will have a different mass and also slightly different properties.

Some isotopes are naturally occurring, others are man-made, like plutonium.

Radioisotopes contain two few or too many neutrons to remain stable. They emit energy and/or subatomic particles to become more stable. They are used as fuel, like uranium, or in medicine, or can cause cancer in living things.

The average atomic mass given on your periodic tables includes an average of all naturally occurring isotopes, along with their relative abundances in nature.

Example: Silver has two isotopes, one with an atomic mass of 106.9 and a relative abundance of 51.8 %, the other 108.9, with a relative abundance of 48.2%.

To calculate the average atomic mass for silver multiply the atomic mass of each by its relative abundance (of course the relative abundance must be converted from a percent to a decimal number) and add them together.

Average atomic mass of silver = (106.9)(0.518) + (108.9)(0.482)
= 107.9

ISOTOPE NOTATION

$${}_{Z}X$$

X = element symbol

A = mass number

Z = atomic number

To name an isotope use the element name and its mass number:

example silver-107