Ch. 4- The Atom

Early Models of the Atom:

Democritus: (~335 BC Greek Era)

- Atoms were indivisible and indestructible

Dalton's Atomic Theory (1808):

- Atom= simple sphere, no internal structure

JJ Thomson (1904):

- Plum Pudding Model

Rutherford (1911):

- Discovered the nucleus – central core of atom **(Gold-Foil experiment)**
- Planetary model- protons in a nucleus; electrons distributed around the nucleus

Structure of the Nuclear Atom:

Atom:

- Smallest particle of an element
- Made up sub-atomic particles:

  Protons: + charge (**Goldstein 1886, Rutherford 1917)**

  Electrons: – charge (**J.J. Thomson 1897, R.Millikan 1916)**

  Neutrons: no charge (**Chadwick 1932)**
- No net electric charge

Properties of Sub-atomic particles				
Particle	**Symbol**	**Relative Charge**	**Relative mass (mass of proton = 1)**	**Actual mass (g)**
Electron	e^-	-1	1/1840	9.109381×10^{-28} g
Proton	P^+	+1	1	1.67×10^{-24} g
Neutron	n^0	0	1	1.67×10^{-24} g

4.3 Distinguishing Among Atoms:

Atomic number:

- # of protons (and electrons)
- IDs the element

Mass number (atomic mass):

- # of protons (or atomic #) + neutrons

*Number of neutrons = mass # – # of protons (or atomic #)

Representing Atomic # and Mass # shorthand:

- mass # = superscript (A)
- atomic # = subscript (Z)
- $^{A}_{Z}X$

ex. $^{197}_{79}Au$, ^{197}Au
or Gold-197, Au-197

Isotopes:

- Atoms w/same # of protons (and electrons) different # of neutrons
- Led to the discovery of neutrons

ex. 3 Hydrogens isotopes:
Hydrogen-1 ($^{1}_{1}\mathrm{H}$) – 1 proton, no neutrons (mass # =1)
Hydrogen-2 ($^{2}_{1}\mathrm{H}$) – 1 proton, 1 neutron (mass #= 2)
Hydrogen-3 ($^{3}_{1}\mathrm{H}$) – 1 proton, 2 neutrons (mass# = 3)

Fill in Rest of chart

Element	Symbol	# of protons	# of electrons	# of neutrons	Atomic #	Mass #	Shorthand
Manganese		25		30			
Sodium			11	12			
Bromine		35		45			
Yttrium					39	89	
Arsenic			33			75	
Actinium						227	

Average Atomic Mass:

- A weighted average of the masses of the isotopes of an element (reflects both the mass and the relative abundance in nature)
- Atomic mass unit = (amu)

Calculating average atomic mass:

- Mass of each isotope x relative abundance (expressed as a decimal) then add the products.

Ex. 2 most abundant isotopes of carbon are **carbon 12** (mass= 12.000000amu) and **carbon-13** (mass = 13.003355 amu). Their percent abundances are 98.93% and 1.07%. Calculate the atomic mass of carbon.

12.000000 x .9893 = 11.872

13.003355 x .0107 = + .139

Average Atomic Mass = 12.011 amu