Chapter 3 - Atoms: The Building Blocks of Matter

3-1 The Atom: From Philosophical Idea to Scientific Theory

Democritus (400 BC)

"The only existing things are atoms and empty space; all else is opinion"

I. Foundations of Atomic Theory
   - A. Quantitative Analysis
     1. Analysis of chemical rxns using improved balances and other measuring devices
   - B. Law of Conservation of Mass
     1. Mass is neither created nor destroyed during ordinary chemical rxns
   - C. Law of Definite Proportions
     1. A chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or the source of the sample
   - D. Law of Multiple Proportions
     1. If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers
        - a. CO_2 and CO
        - b. H_2O and H_2O_2

II. Atomic Theory
   - A. John Dalton (1766 -1844)
     1. All matter is made up of very tiny particles called atoms
     2. Atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass and other properties
     3. Atoms cannot be subdivided, created, or destroyed
     4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds
   - B. Shortcomings of Dalton's theory
     1. Could not explain isotopes
   - C. Modern Atomic Theory
     1. All matter is made up of very tiny particles called atoms
     2. Atoms of the same element are chemically alike
     3. Individual atoms of an element may not all have the same mass. However, the atoms of an element have a definite average mass that is characteristic of the element
     4. Atoms of different elements have different average masses
     5. Atoms are not subdivided, created, or destroyed in chemical rxns

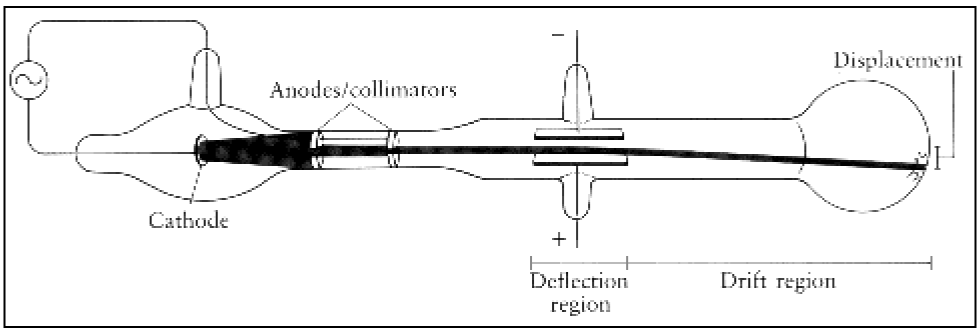
3-2 The Structure of the Atom

Atom - The smallest unit of an element that retains the properties of that Element

I. <u>Discovery of the Electron</u>

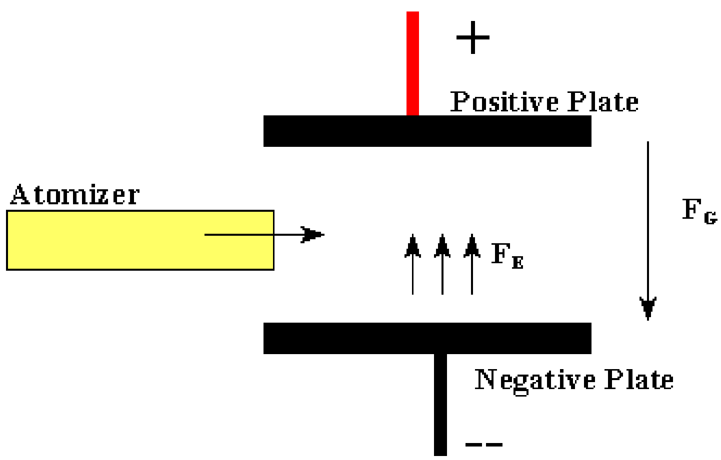
A. Cathode Rays and the Electron

1. Joseph John Thomson (1897)
    a. Cathode ray tube produces a ray with a constant charge to mass ratio
    b. All cathode rays are composed of identical negatively charged particles (electrons)



B. Charge and Mass of the Electron

1. Millikan's Oil Drop Experiment

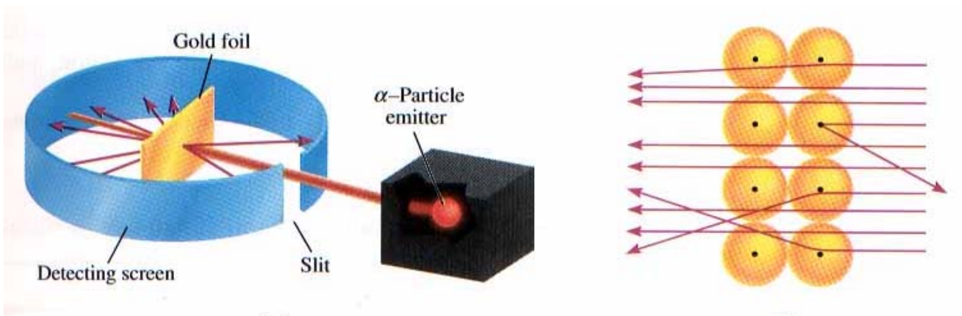


    a. Electron is negatively charged
    b. Mass is about 1/2000th of a hydrogen atom
        (a) Electron mass is 9.109×10^{-31}kg
2. Inferences from the properties of electrons
    a. Atoms are neutral, so there must be positive charges to balance the negatives
    b. Electrons have little mass, so atoms must contain other particles that account of most of the mass

II. The Nucleus

A. The Rutherford Experiment (1911)

1. Alpha particles (helium nuclei) fired at a thin sheet of gold
    a. Assumed that the positively charged particles were bounced back if they approached a positively charged atomic nucleus head-on
    (Like charges repel one another)



2. Very few particles were greatly deflected back from the gold sheet
    a. nucleus is very small, dense and positively charged
    b. most of the atom is empty space

B. Structure of the Nucleus

1. Protons
    a. Positive charge, mass of 1.673×10^{-27}kg
    b. The number of protons in the nucleus determines the atom's identity and is called the atomic number
2. Neutrons
    a. No charge, mass of 1.675×10^{-27}kg
3. Nuclear Forces
    a. Short range attractive forces:
    neutron-to-neutron, proton-to-proton, proton-to-neutron

Particle	Symbols	Relative charge	Mass Number	Relative mass (amu*)	Actual mass (kg)
Electron	e^- $\ {}^{0}_{-1}e$	-1	0	0.000 5486	9.109×10^{-31}
Proton	p^+ $\ {}^{1}_{1}\mathrm{H}$	+1	1	1.007 276	1.673×10^{-27}
Neutron	n° $\ {}^{1}_{0}n$	0	1	1.008 665	1.675×10^{-27}

C. Sizes of Atoms

1. Atomic radius
    a. 40 to 270 picometers (pm)
        (1) 1 pm = 10^{-12}m
    b. Most of the atomic radius is due to the electron cloud
2. Nuclear radius
    a. 0.001 pm
    b. density is 2x108 metric tons/cm^3
        (1) 1 metric ton = 1000kg

3-3 Counting Atoms

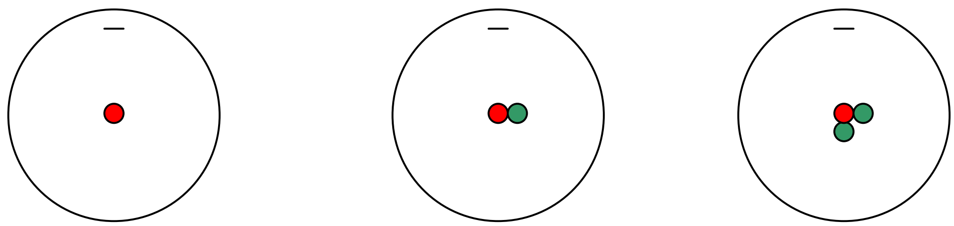
I. <u>Atomic Number, Isotopes and Mass Number</u>

A. Atomic Number (Z)
   1. The number of protons in the nucleus of each atom of that element
   2. Atoms are identified by their atomic number
   3. Because atoms are neutral,

      # protons = # electrons

   4. Periodic Table is in order of increasing atomic number

B. Isotopes
   1. Atoms of the same element that have different masses
   2. All elements of the same element have the same # of protons, but may vary in the number of neutrons
   3. Although isotopes have different masses, they do not differ significantly in their chemical behavior
   4. Hydrogen as an example:



Protium (hydrogen-1) Deuterium (hydrogen-2) Tritium (hydrogen-3)

C. Mass Number
   1. The total number of protons and neutrons in the nucleus of an isotope

D. Designating Isotopes
   1. Hyphen notation
      a. Mass number is written after the name of the element
         (1) hydrogen-2, helium-4
   2. Nuclear Symbol
      a. Composition of the nucleus using the element's symbol
         (1) $^{2}_{1}H$ Mass number =2 Atomic number = 1
         (2) $^{4}_{2}He$ Mass number = 4 Atomic number = 2

II. <u>Using Atomic Mass</u>

A. Relative Atomic Masses
   1. Atomic mass unit (amu)
      a. 1 amu = 1/12 the mass of a carbon-12 atom
   2. All masses on the periodic table are relative to the carbon-12 standard
   3. Approximate masses of atomic particles
      a. proton = 1 amu
      b. neutron = 1 amu
      c. electron = 0.000 5486 amu

B. Average Atomic Masses
   1. The weighted average of the atomic masses of the naturally occurring isotopes of an element
      a. Atomic masses on the periodic table are average masses
      b. In calculations using atomic mass, we will round the masses to two decimal places before doing calculations

Examples:	Mg = 24.3050	we use: 24.31
	O = 15.9994	we use: 16.00
	N = 14.00674	we use: 14.01

III. <u>Avogadro's Number and the Mole</u>

A. The Mole
   1. The amount of substance that contains as many particles are there are in exactly 12 grams of carbon-12
   2. The amount of substance that contains the Avogadro number of particles

B. Avogadro's Number
   1. The number of particles in exactly one mole of a pure substance
   2. Avogadro's number = 6.022×10^{23}

C. Molar Mass
   1. The mass of one mole of a pure substance
      a. Units are grams/mole (or g/mol)
      b. Molar mass of an element equals the average atomic mass in gram units

D. Gram/Mole Conversions
   1. Convert from grams to moles

$$grams \quad x \quad \left(\frac{1}{molar\ mass} \right) = moles$$

   2. Convert from moles to grams

$$moles \quad x \quad molar \quad mass = grams$$

E. Conversions with Avogadro's Number
   1. Review sample problems on pages 84 and 85 - solutions vary depending on the given data and the desired quantity