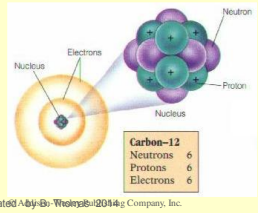
Ch. 3 - Atomic Structure

III. Counting Atoms

- Mass Number
- Isotopes
- Relative Atomic Mass
- Average Atomic Mass

A. Mass Number

- mass # = protons + neutrons
- always a whole number
- NOT on the Periodic Table!



B. Isotopes

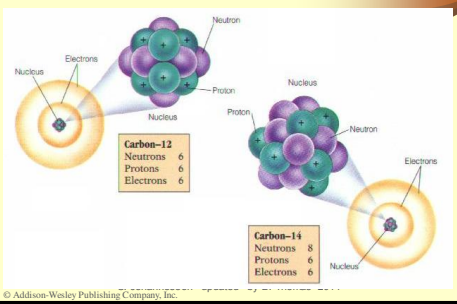
- Atoms of the same element with different mass numbers.
- Nuclear symbol:

Mass # → $^{12}_{6}C$

Atomic # →

- Hyphen notation: carbon-12

B. Isotopes



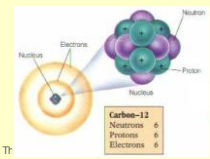
B. Isotopes

- **Chlorine-37**
  - atomic #: 17
  - mass #: 37
  - # of protons: 17
  - # of electrons: 17
  - # of neutrons: 20

$^{37}_{17}Cl$

C. Relative Atomic Mass

- ^{12}C atom = 1.992×10^{-23} g
- atomic mass unit (amu)
- 1 amu = $^{1}/_{12}$ the mass of a ^{12}C atom
- 1 p = 1.007276 amu
  1 n = 1.008665 amu
  1 e^- = 0.0005486 amu



D. Average Atomic Mass

- weighted average of all isotopes
- on the Periodic Table
- round to 2 decimal places

$$\text{Avg. Atomic Mass} = \frac{(\text{mass})(\%) + (\text{mass})(\%)}{100}$$

D. Average Atomic Mass

- EX: Calculate the avg. atomic mass of oxygen if its abundance in nature is 99.76% ^{16}O, 0.04% ^{17}O, and 0.20% ^{18}O.

$$\text{Avg. Atomic Mass} = \frac{(16)(99.76) + (17)(0.04) + (18)(0.20)}{100} = 16.00 \text{ amu}$$

D. Average Atomic Mass

- EX: Find chlorine's average atomic mass if approximately 8 of every 10 atoms are chlorine-35 and 2 are chlorine-37.

$$\text{Avg. Atomic Mass} = \frac{(35)(8) + (37)(2)}{10} = 35.40 \text{ amu}$$