Unit 4: Introduction to Atomic Structure; Isotopes & Average Atomic Mass

See also pp. 73-80 in the text

- Structure of the Atom (what Dalton got wrong) — Part 1
  - 3 Fundamental Particles:

Particle	Symbols	Charge	Mass	Location
proton	p^+ $_1^1p$ $_1^1H$	1+	1.0073 amu	Nucleus
neutron	n^0 $_0^1n$	0	1.0087 amu	Nucleus
electron	e^- $_{-1}^0e$	1−	1/1823 amu	"Cloud"

*amu = atomic mass unit. 1 amu ≡ $\frac{1}{12}$ mass of C-12 atom. 1 g = 6.022×10^{23} amu

- Identity of atom: Atomic Number (Z) — Charge on nucleus
  - Number of protons
    - Must also be number of electrons since atoms are electrically neutral (No net charge)
    - How many protons does C have? 6 electrons? 6

      What about gold? Au 79, 79

      Uranium? U 92, 92

- Isotopes & Mass Number
  - Isotopes are different atoms of same element with different #'s of neutrons
  - Mass number A = #p + #n (whole #)
    - NOT avg. atomic mass (Can't tell Mass # or # Neutrons from periodic table)
  - Two ways of indicating: Element-A or $^{A}_{Z}X$
    - C-12 or $^{12}_{6}C$: # protons? 6 neutrons? 6 (12−6) Atomic #? 6 Mass #? 12
    - C-13 — 6 protons, 7 neutrons, 6, 13
    - C-14 — 6 protons, 8 neutrons, 6, 14

N-14 — 7 protons, 7 neutrons, 14 — Not an isotope of C

- Relative Isotopic Abundance & Average Atomic Mass
  - Atomic mass is weighted average of the masses of all of an element's naturally occurring isotopes
    - Weighted average takes into account relative percentages of each isotope
  - For element, average atomic mass = $\sum_{\text{All Isotopes of Element}} \left(\text{Isotope Mass} \times \frac{\%\ \text{Abundance}}{100\%} \right)$ (sum; amu)
    - Example: Chlorine has two natural isotopes: Cl-35 (mass = 34.97 amu; abundance = 75.77%) and Cl-37 (36.97 amu, 24.23%):
      - Notice that mass number (A) is approximately mass of that isotope

Avg. At. Mass = $(34.97\text{ amu})\left(\frac{75.77\%}{100\%}\right) + (36.97\text{ amu})\left(\frac{24.23\%}{100\%}\right)$

Cl-35 Contribution: 26.50 amu + Cl-37: 8.958 amu

= 35.46 amu

- Average mass is closer ("weighted") to more abundant isotope

- Identify the element by calculating the atomic mass from the isotopic mass and abundance:

Isotope	Mass(amu)	% Abundance	Contribution (amu)
X-20	19.9924	90.48 / 100 =	18.09
X-21	20.9938	0.27	0.057
X-22	21.9914	9.25	2.03
			20.18

- At. Mass = 20.18; matches atomic mass of Ne