CHAPTER 2

ATOMS AND ATOMIC THEORY

KEY CONCEPTS:

1. THE DEVELOPMENT OF ATOMIC THEORY

2. ELEMENTS AND THEIR ISOTOPES

3. ATOMIC MASS

4. THE PERIODIC TABLE

5. AVOGADRO’S NUMBER AND THE MOLE

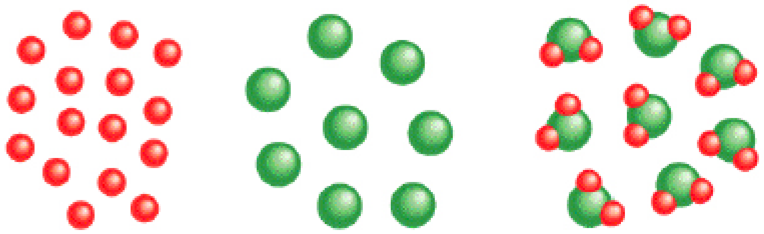
Dalton’s Atomic Theory

- All matter is made of small, indestructible particles called “atoms”.

- All atoms of a given element are identical and atoms differ for different elements.

- Compounds are formed when atoms of more than one element combine

- Atoms of different elements combine in simple, whole number ratios (e.g. 1:1 as in AB; or 1:2 as in AB_2)



Atoms of « X » Atoms of « Y » Compound « X_2Y »

- N.B. Dalton did not know the structure of the atom (i.e. electrons, protons, neutrons, etc.)
  - Dalton believed the atom was small and indivisible

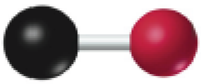
- Dalton’s work helped to explain several empirical observations:
  - the law of constant composition
  - the law of multiple proportions
  - the law of conservation of mass

The Law of Constant Composition

All samples of a given compound have the same composition

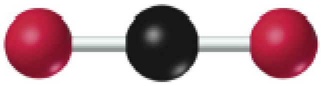
The Law of Multiple Proportions

Combinations of elements are in ratios of small whole numbers (to make molecules!)



Carbon monoxide (CO): 1.33 g of oxygen combines with 1.0 g of carbon or O/C = 1.33

Carbon dioxide (CO_2): the ratio of O/C is 2.667



2.667/1.33 = 2.0 which is the ratio of oxygen in CO_2/oxygen in CO

The Law of Conservation of Mass

The total mass during a chemical or physical process must remain constant

Atomic Structure: the electron

J.J. Thompson: 1898 – 1903

- the existence of a subatomic particle with a negative charge was demonstrated using a cathode-ray tube: *the electron*

$$\frac{\text{charge}}{\text{mass}} = -1.76 \times 10^{8}\ \text{C/g}$$

R. Millikan, 1909

- conceived an ingenious experiment to determine the exact charge of the electron
- his setup measured the speed of the fall of tiny droplets of oil

- with this experiment, he established the charge on a single electron

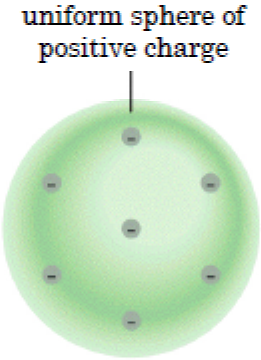
charge = $-1.60\text{x}10^{-19}$ C

- using the Z/m ratio found by Thompson, the electron's mass was then:

$$
\begin{aligned}
\text{mass} &= Z \bullet \frac{m}{Z} \\
&= -1.60\times10^{-19}\ \text{C} \bullet \frac{\text{g}}{-1.76\times10^{8}\text{C}} \\
&= 9.09\times10^{-28}\ \text{g}
\end{aligned}
$$

The structure of the atom

- before 1910, Thompson's "plum pudding" model of the atom was the most popular
- in this model, the small electrons are dispersed in a much larger uniform sphere of positive charge
- this model seemed logical, given the incredibly small mass of the electron



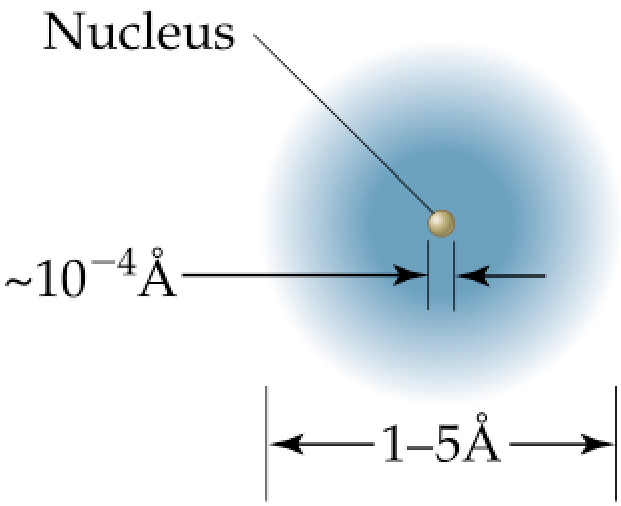
E. Rutherford, 1911 (in Montreal)

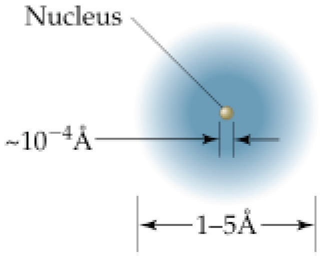
- in his experiment, a very thin foil of gold was bombarded with α particles

The Modern View of Atomic Structure

Nucleus

- Rutherford and Chadwick later discovered that the nucleus contains protons and neutrons
- nuclear radius is 20,000 times smaller than the atomic radius!





The mass of the atom comes from the nucleus, the volume from the orbits of the electrons

Particle	Symbol	Mass	Charge
Proton	p^+	1.67×10^{-24} g	$+1.60 \times 10^{-19}$ C
Neutron	n	1.67×10^{-24} g	0
Electron	e^-	9.11×10^{-28} g	-1.60×10^{-19} C

Scale of Atoms

Useful units:

- 1 amu (atomic mass unit) = 1.66054×10^{-24} kg
- 1 pm (picometer) = 1×10^{-12} m
- 1 Å (Angstrom) = 1×10^{-10} m

= 100 pm = 1×10^{-8} cm

typical C-C single bond length = 1.54 Å

Isotopes, Atomic Numbers, and Mass Numbers

- atomic number **Z = number of protons**
- atomic mass **A = Sum of protons and neutrons**

- For the element X, we write: $^{A}_{Z}X$

1. each element has a unique value of Z

2. isotopes of an element will have the same value of Z, but different values of A (due to number of neutrons)

Isotopes

- the different isotopes of an element usually display similar chemistry, since the number of neutrons has little influence on the reaction’s outcome
- we use the atomic mass to identify isotopes:

$^{235}_{92}U$ is uranium-235

$^{238}_{92}U$ is uranium-238

- the isotopes of hydrogen have special names:

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
hydrogen	deuterium	tritium

Atomic masses

In the SI system, the standard for atomic mass is pure carbon-12:

$$\boxed{\text{the mass of } {}^{12}\text{C} = 12 \text{ amu}}$$

where 12 amu is an *exact value* (i.e. = 12.0000000000000...)

∴ 1 amu (« u ») = 1/12 the mass of a C atom
= 1.66054×10^{-24} g
or 1 g = 6.02214×10^{23} u

- The masses of all other elements are *measured relative* to the ^{12}C atom
- the mass of a single atom of carbon-12 is therefore

$$\frac{12 \text{ g}}{6.022045 \times 10^{23} \text{ atoms}} = 1.992679 \times 10^{-23} \text{ g}/\text{atom}$$

- a single atomic mass unit is 12 times smaller, so:

$\boxed{1 \text{ u} = 1.660565 \times 10^{-24} \text{ g}}$ and $\boxed{1 \text{ g} = 6.022045 \times 10^{23} \text{ u}}$

Average atomic mass

IMPORTANT: in nature, the elements exist as a *mixture of isotopes*
∴ we must account for the relative proportions of each isotope in our calculations

Atomic mass = weighted average of the respective atomic masses of the natural isotopes of a given element

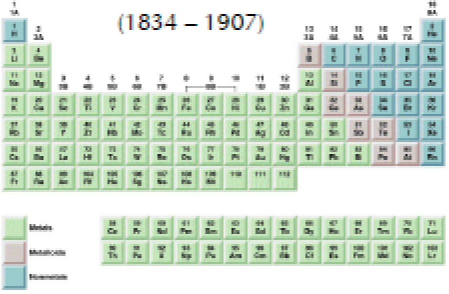
ex. carbon has two naturally occurring isotopes:

	^{12}C	^{13}C
Mass	12 u	13.00335 u
Abundance	98.892 %	1.108 %

The average atomic mass of carbon is:
(0.98892)(12) + (0.0108)(13.00335) = **12.011 u**

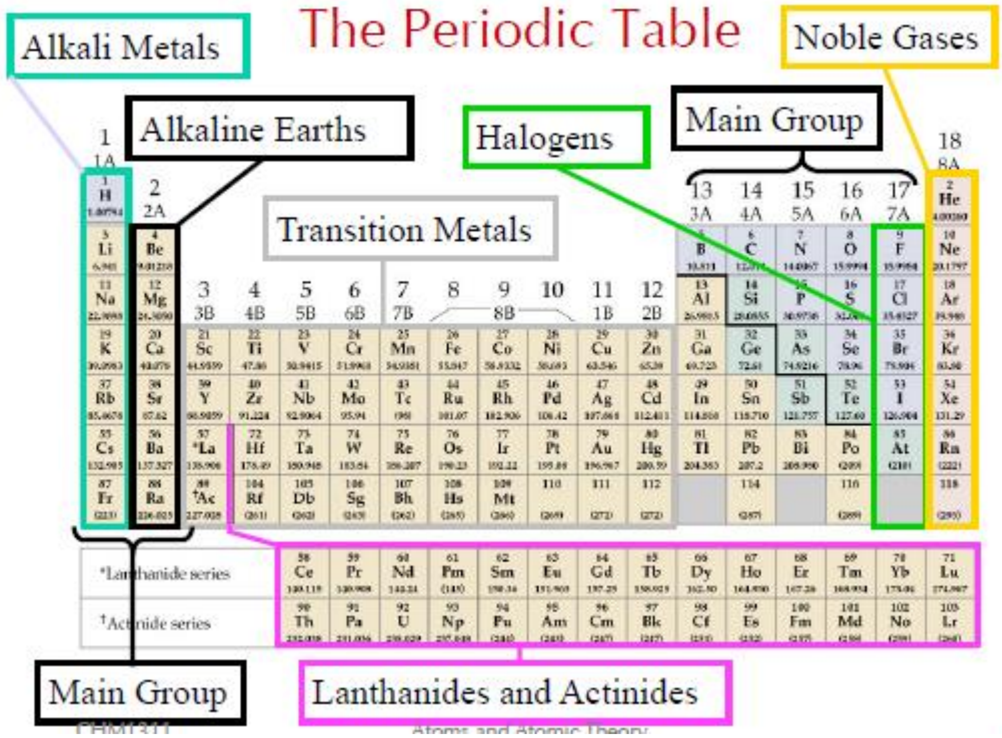
The Periodic Table

(1834 – 1907)



- table in which the elements are grouped together according to their similar physical and chemical properties
- the horizontal rows are called periods
- the vertical columns are called groups (or families)
- three categories of elements:
  - metals (good conductors)
  - non-metals (poor conductors)
  - metalloids or semi-metals

The Periodic Table



Alkali Metals

Alkaline Earths

Transition Metals

Halogens

Main Group

Noble Gases

Main Group

Lanthanides and Actinides

Avogadro’s number

Amedeo
Avogadro
(1776-1856)

The **mole** is an amount of substance that contains the same number of elementary entities as there are carbon-12 atoms in *exactly* 12 g of carbon-12.

$$N_A = 6.02214199 \times 10^{23}\ \mathrm{mol}^{-1}$$

Molar Mass

- The molar mass, M, is the mass of one mole of a substance.

$$M\ (\mathrm{g/mol}\ ^{12}\mathrm{C}) = \mathrm{A}\ (\mathrm{g/atom}\ ^{12}\mathrm{C}) \times N_A\ (\mathrm{atoms}\ ^{12}\mathrm{C}\ /\mathrm{mol}\ ^{12}\mathrm{C})$$

- the numerical value of the atomic mass and the molar mass of an element are identical, but they have ***different units!***