

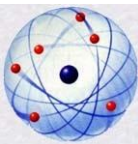
Chapter II : Atomic structure

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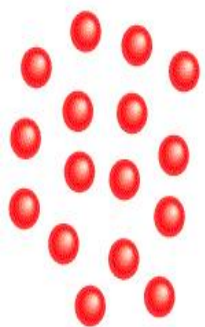
Brief History

- **Origin of the word “atom”**: Greek philosophers (400 B.C.) – *atomos* = indivisible.
- **17th–18th centuries**: First atomic theories; atoms combine to form molecules (Lavoisier, Dalton, ...).
- **1869** : Mendeleev introduced the **periodic table of elements** (classification by atomic mass).
- **1897** : Discovery of the **electron** – first subatomic model (Thomson).
- **1912** : : Discovery of the **nucleus** – Rutherford’s planetary model.
- Development of **quantum physics** (Bohr, Sommerfeld, ...).

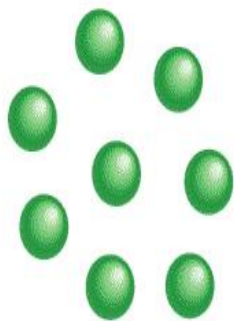


Dalton's Atomic Theory (1808)

Dalton's Atomic Theory

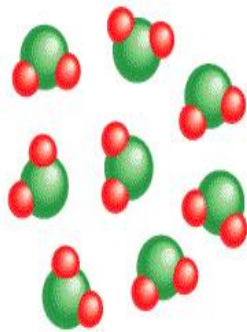


Atoms of element X



Atoms of element Y

(a)



Compound of elements X and Y

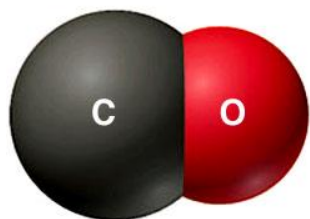
(b)

- All matter (Elements) is composed of extremely small particles called **atoms**.
- All atoms of a given element are **identical**.
- Atoms of one element are **different** from those of all other elements.
- **Compounds** are formed from more than one element.
- In any compound, the ratio of atoms of two elements is either a **whole number** or a **simple fraction**.

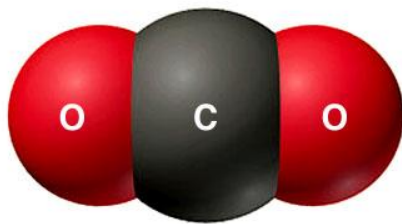
- A **chemical reaction** is only the separation, combination, or rearrangement of atoms; it does not involve the creation or destruction of atoms.
- N.B. Dalton did not know the internal structure of the atom (electrons, protons, neutrons, nucleus, etc.).
- Dalton imagined the atom as small and indivisible.

Dalton's Atomic Theory (1808) – Applications

Atomic Basis of the Law of Multiple Proportions



Carbon oxide I
(carbon monoxide)



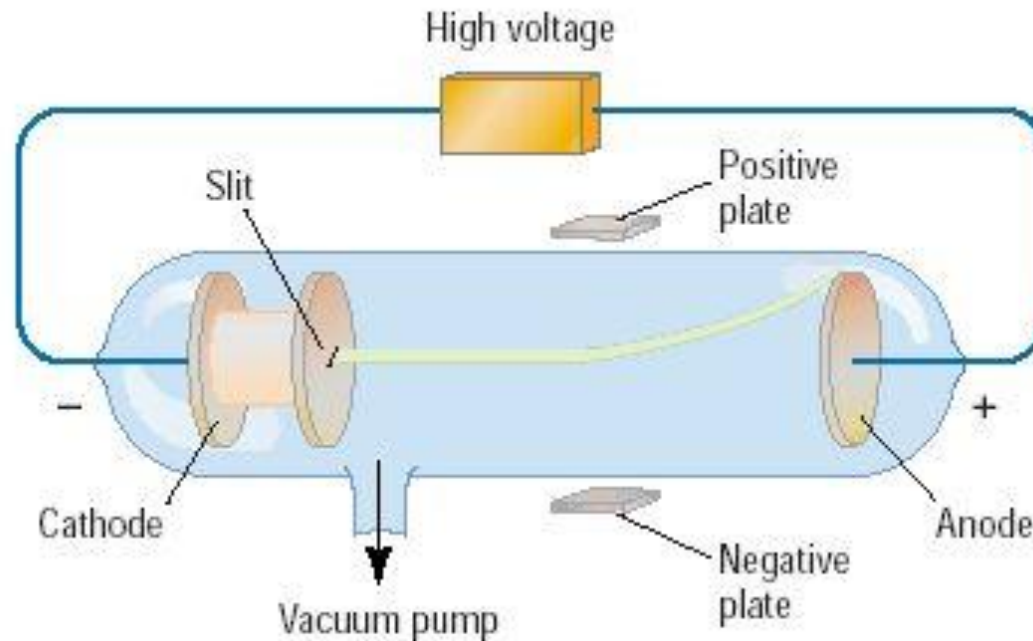
Carbon oxide II
(carbon dioxide)

Dalton's work explained several empirical laws:

- ✓ The **Law of Definite Proportions**.
- ✓ The **Law of Multiple Proportions**.
- ✓ The **Law of Conservation of Mass**.

Discovery of the Electron

In 1897, J.J. Thomson used a cathode ray tube to deduce the presence of a negatively charged particle.



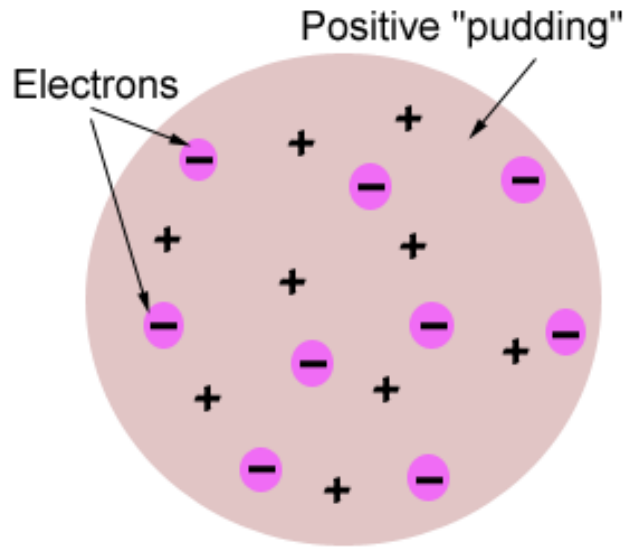
Cathode ray tubes pass electricity through a gas that is contained at a very low pressure.

Conclusions from the study of the Electron

- ❑ Cathode rays have identical properties regardless of the element used to produce them. All elements must contain identically charged electrons.
- ❑ Atoms are neutral, so there must be positive particles in the atom to balance the negative charge of the electrons
- ❑ Electrons have so little mass that atoms must contain other particles that account for most of the mass

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

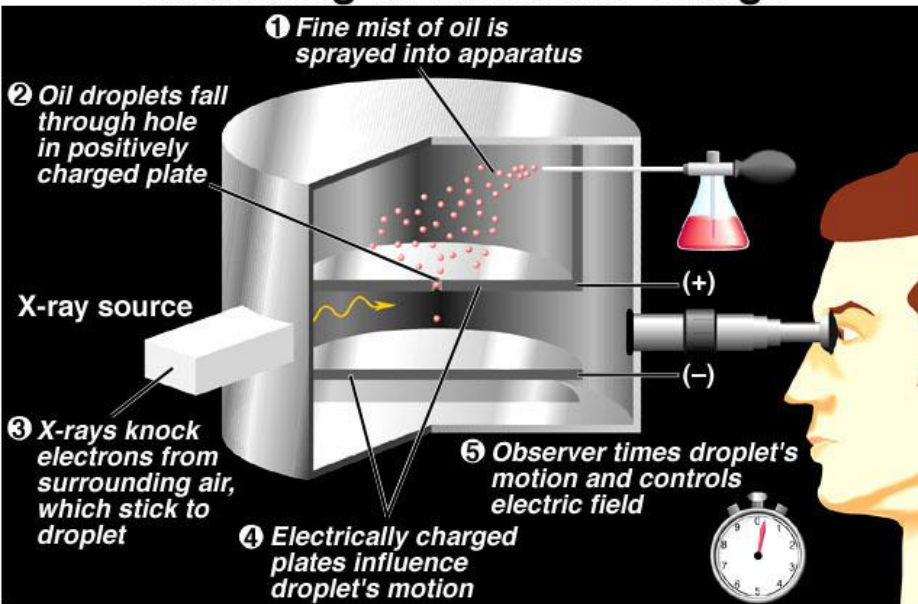
Thomson's Atomic Model



Thomson believed that the electrons were like plums embedded in a positively charged “pudding,” thus it was called the “plum pudding” model.

Structure of the Atom: The Electron (Millikan)

Millikan's Oil-drop Experiment for Measuring an Electron's Charge

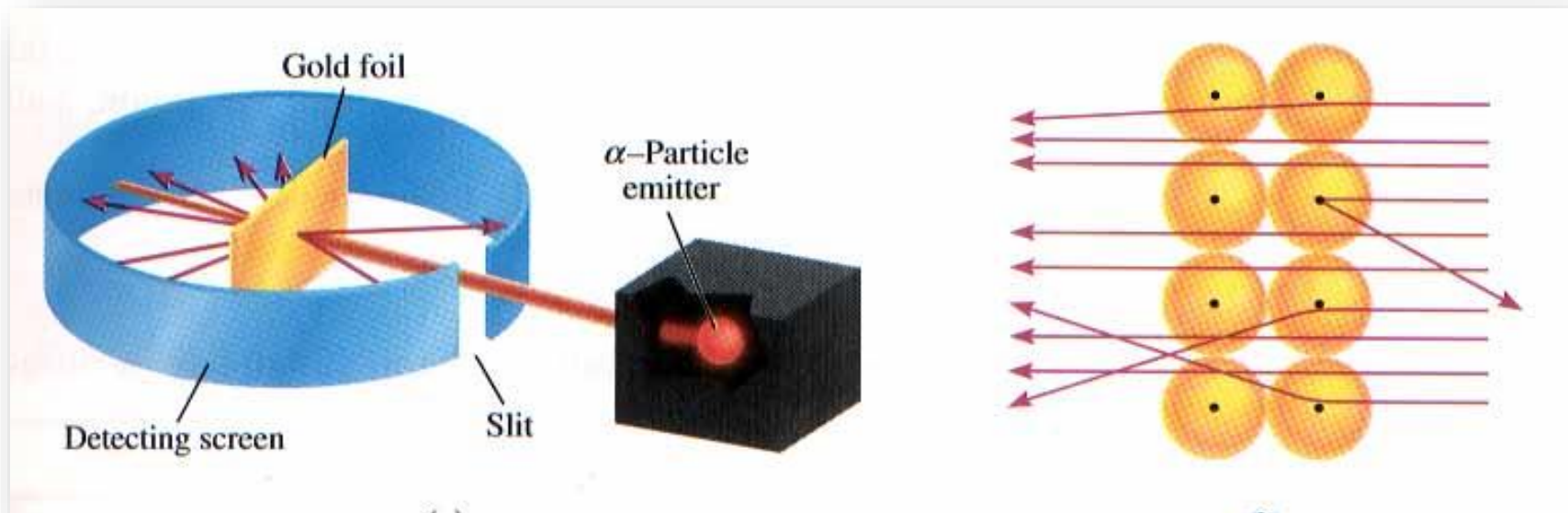


- Around 1900, **Millikan** measured the **charge of the electron**:
 $q = -1.60 \times 10^{-19} \text{ C}$
- With Thomson's work, Millikan deduced the **mass of the electron**.

$$\begin{aligned} \text{masse} &= \frac{\text{charge}}{\text{charge} / \text{masse}} \\ &= \frac{-1.60 \times 10^{-19} \text{ C}}{-1.76 \times 10^8 \text{ C/g}} = 9.09 \times 10^{-28} \text{ g} \end{aligned}$$

Rutherford's Gold Foil Experiment

Rutherford bombarded thin gold foils with **alpha particles**.



- Alpha (α) particles are helium nuclei
- Particles were fired at a thin sheet of gold foil
- Particle hits on the detecting screen (film) are recorded

Interpretation

- The deflections were due to strong **repulsions**.
- Since alpha particles are positively charged, the positive charge must be concentrated in a very small, massive **nucleus** at the center of the atom.
- The nucleus occupies about **10^{-13}** of the atomic volume, but contains nearly **all the mass**.
- The positive charges inside the nucleus are **protons**.
- The charge of a proton is equal in magnitude but opposite in sign to that of an electron.
- Proton mass: 1.67252×10^{-24} g, (about 1840 times the electron's mass).

Rutherford's Findings

- Most of the particles passed right through
- A few particles were deflected
- VERY FEW were greatly deflected



Conclusions:

- The nucleus is small
- The nucleus is dense
- The nucleus is positively charged

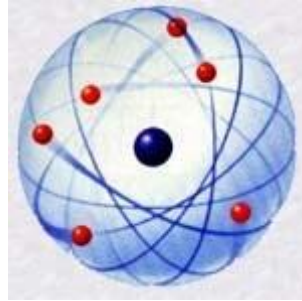
Structure of the Atom: The Neutron (Chadwick, 1932)

- Before 1932, a mystery remained: hydrogen has one proton, while helium has two protons but is **four times more massive**.
- The existence of a neutral particle was hypothesized.
- **Chadwick discovered the neutron in 1932.**
- The neutron has no charge and is slightly more massive than the proton:
 - Neutron mass: 1.67493×10^{-24} g
 - Proton mass: 1.67252×10^{-24} g.

Atomic Particles

Particle	Charge	Mass #	Location
Electron	-1	0	Electron cloud
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus

Model of the Atom: 1 Nucleus + 1 Electron Cloud



<u>Nucleus</u> :	Z protons	+	N neutrons	➔	A nucleons
	charge $+e$ ($+e = 1,60219 \cdot 10^{-19}$ C)		neutre		
	diameter $\sim 10^{-14}$ m		diameter $\sim 10^{-14}$ m		
	$m_p = 1,6727 \cdot 10^{-27}$ kg		$m_n = 1,6750 \cdot 10^{-27}$ kg		

Nuclear cohesion due to **strong nuclear forces** (much stronger than electrostatic or gravitational interactions).

Electron Cloud : Z electrons

charge $-e$
($-e = -1,60219 \cdot 10^{-19}$ C)
diameter $\sim 10^{-18}$ m
 $m_e = 9,1091 \cdot 10^{-31}$ kg

Gravitate at about 1 \AA
(10^{-10} m) from the nucleus

Atomic Number

Atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element.

Element	# of protons	Atomic # (Z)
Carbon	6	6
Phosphorus	15	15
Gold	79	79

Mass number is the number of protons and neutrons in the nucleus of an isotope.

$$\text{Mass \#} = p^+ + n^0$$

Nuclide	p^+	n^0	e^-	Mass #
Oxygen - 18	8	10	8	18
Arsenic - 75	33	42	33	75
Phosphorus - 31	15	16	15	31

General notation

Z

1 atom $X \Leftrightarrow {}^A_Z X$



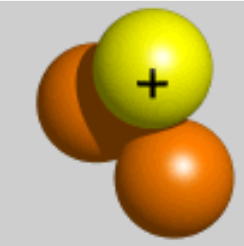
$A = Z + N$ A : mass number
 Z : atomic number

Example : ${}^{35}_{17}\text{Cl}$ ou ${}^{37}_{17}\text{Cl}$

- The atom is **electrically neutral**.
- Nearly all the mass is concentrated in the **nucleus**.
- Matter is mostly **empty space**.

Isotopes

Isotopes are atoms of the same element having different masses due to varying numbers of neutrons.

Isotope	Protons	Electrons	Neutrons	Nucleus
Hydrogen-1 (protium)	1	1	0	
Hydrogen-2 (deuterium)	1	1	1	
Hydrogen-3 (tritium)	1	1	2	

Atomic Masses

Atomic mass is the average of all the naturally occurring isotopes of that element.

Isotope	Symbol	Composition of the nucleus	% in nature
Carbon-12	^{12}C	6 protons 6 neutrons	98.89%
Carbon-13	^{13}C	6 protons 7 neutrons	1.11%
Carbon-14	^{14}C	6 protons 8 neutrons	<0.01%

Carbon = 12.011

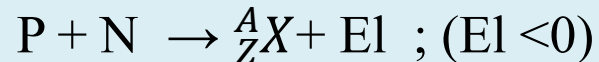
The Mole

- The mass of an atom and its diameter are too small to be counted one by one.
- The **mole (mol)** is the unit of amount of substance.
- **Avogadro's number:** $N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$.
- **Atomic mass unit (a.m.u.):** 1/12 of the mass of a ^{12}C atom.
 $1 \text{ a.m.u.} = 1.66057 \times 10^{-27} \text{ kg}$.

Nuclear Stability

NUCLEUS BINDING AND COHESION ENERGY:

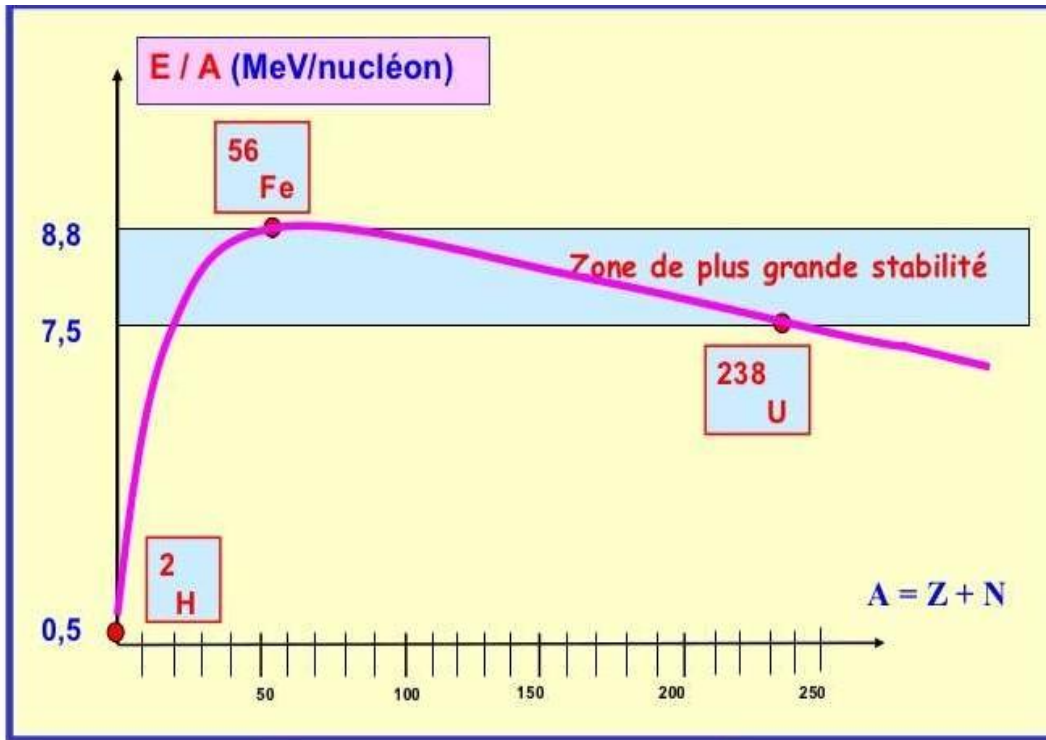
1. Binding energy (EB): Energy required for the formation of any nucleus from these nucleons (Protons (P) + Neutrons (N)). The formation of the nucleus is described by the following reaction:



The formation of a nucleus generally requires negative energy.

2. Cohesion energy (E_{co}): Energy necessary for the destruction of a nucleus into its constituents (N+P) according to the reaction: ${}^A_ZX + E_{co} \rightarrow P + N$, this energy is positive and we can write: $E_{co} = -EB$.

In general, the cohesive energy per nucleon is less than 8.8 MeV, whatever the considered element. This average cohesion energy can be plotted as a function of the number of nucleons A on a diagram called the **Aston curve**.



Aston curve

This cohesion energy is of the order of MeV/nucleon ($1\text{MeV} = 10^6\text{ eV} = 1.6 \cdot 10^{-13}\text{ J}$).

The curve obtained shows a maximum at around $A = 60$, the corresponding atoms being the most stable atoms in existence.

The slope of the Aston curve is very steep in the region of “light” atoms with $A < 15$. For “heavy” atoms with $Z > 15$, the slope is much gentler.

Light and heavy atoms (whose average binding energy is low ($E / A < 7.5$ MeV / nucleon), will seek to stabilize and approach the zone of maximum stability around $Z = 60$.

Two different processes are possible: fusion for light atoms and fission (disintegration) for heavy atoms.

Calculating the binding energy EB or cohesion energy Eco of nuclei: Theory of relativity (EINSTEIN equation) :

The concepts of mass and energy are always the same: disappearance of mass
=> addition of energy or disappearance of energy.

Hence : $E = m * C^2$; (C: celerity or speed of light; $C = 3 \cdot 10^8$ m/s).

Note: The mass of the nucleus is always less than the sum of the masses of these constituents, there is a loss of mass Δm which is transformed into energy with : $\Delta E = \Delta m * C^2$;

Δm = mass disappearance = mass defect.

According to the reaction of nucleus formation : $Z + N \rightarrow {}^A_Z X + \Delta E_B$

$$\Delta m = [\text{nucleus mass } (\frac{A}{Z}X) - (Z \cdot \text{proton mass} + N \cdot \text{neutron mass})] < 0;$$

When we create a nucleus, we talk about binding energy ($\Delta E_B < 0$).

When the nucleus is broken down (destroyed) into these nucleons : $\frac{A}{Z}X + \Delta E_{CO} \rightarrow Z + N$

$\Delta m = [(Z \cdot \text{mass of proton} + N \cdot \text{mass of neutron}) - \text{nucleus mass } (\frac{A}{Z}X)] > 0$, Δm is positive, This is known as cohesive energy. ($\Delta E_{CO} < 0$).

➤ **The greater the cohesion energy per nucleon (or bond per nucleon), the more stable the nucleus.**

So, to compare the stability of two nuclei, we need to compare the cohesion energy per nucleon (or bond per nucleon) of the two nuclei. The more stable nucleus will therefore have the greater cohesion energy per nucleon (or bond per nucleon).

- **Definition of electron volt:** The electron volt (eV) is the energy acquired by an electron accelerated by a potential difference (p.d.r.) of 1 volt. .

Hence : **1 eV = 1.6 10⁻¹⁹ Coulomb (c) * 1 Volt (V) = 1.6 10⁻¹⁹ joule (J).**

- **Energy of an A.M.U. (1 A.M.U.)**

$$\Delta E = \Delta m \cdot c^2, \Delta m = 1 \text{ a.m.u} = 1.66 \cdot 10^{-27} \text{ kg} \Rightarrow \Delta E = 1.66 \cdot 10^{-27} \text{ kg} * (3 \cdot 10^8)^2 \text{ (m/s)}^2 \Rightarrow$$

$$\Delta E = 14.94 \cdot 10^{-11} \text{ J.}$$

$$1 \text{ eV} = 1.6 \cdot 10^{-19} \text{ J} \Rightarrow \Delta E = (14.94 \cdot 10^{-11}) / (1.6 \cdot 10^{-19}) = 933 \cdot 10^6 \text{ eV} \Rightarrow \Delta E = 933 \text{ MeV,}$$

Hence : **1 a.m.u represents 933 MeV.**