1.1 Introduction

- The Greek philosopher, Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named atomos (meaning uncuttable or indivisible).
- His idea was not accepted by other Greece philosophers (notably Plato and Aristotle).

1.2 Modern Atomic Theory

1.2.1 Dalton's Atomic Theory

The atomic model diagram provided below illustrates the evolution of atomic models.

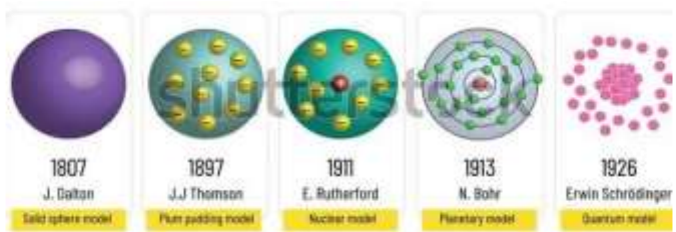

Fig.1.1 Atomic models

Postulates of Dalton's atomic theory

- Elements are made of small particles called atoms.
- Atoms are indivisible.
- Atoms can neither be created nor destroyed.
- All atoms of the same element are identical and have the same mass and size.
- Atoms of different elements have different masses and size.
- Atoms combine in small whole numbers to form compounds.

Drawbacks Of Dalton's Atomic Theory

Which of Dalton's postulates are contradicted by later observations?

- **The indivisibility of an atom was proved wrong**: an atom can be further subdivided into protons, neutrons, and electrons.
- **Atoms of the same element are similar in all respects**. However, atoms of some elements vary in their mass. These atoms of different masses are called isotopes. For example, carbon has three isotopes with mass numbers 12,13 & 14
- **Atoms of different elements differ in mass**. This has been proven wrong because atoms of different elements may have the same mass which are called **isobar**s. For Example, Argon and Calcium atoms each have an atomic mass of 40 amu.
- According to Dalton, **atoms of different elements combine in simple whole-number ratios to form compounds**. This is not observed in complex organic compounds like sugar ($C_{12}H_{22}O_{11}$) and protein molecule.

1.2.2 Modern Atomic Theory

Postulates of The Modern Atomic Theory

- All Matter Is Made-up of tiny, indivisible particles called atoms
- Atoms can not be created or destroyed during ordinary chemical reactions.
- All atoms of the same element have the same atomic number but may vary in mass number
- Atoms of different elements are different.
- Atoms combine in small whole numbers to form compounds

1.3 Early Experiments

(Discovery of Sub-atomic particle)

1.3.1 Discovery of Electron

- William Crookes was the first scientist who designed the discharge tube which was called the Crooke's discharge tube or Cathode ray tube.
- The discharge tube consists of a glass tube from which most of the air has been evacuated having two metal plates sealed at both ends. These metal plates are called electrodes.
- These electrodes are connected to positive and negative terminals of a battery.
- The electrode is connected the positive and negative terminal of the battery.

- When both electrodes are connected to high voltage, current starts flowing and a green glow was observed at the anode.
- The rays are emitted from the direction of the cathode, and are called Cathode rays.
- J. J. Thomson conducted some experiments with a discharge tube for studying the properties of cathode rays.

J. J. Thomson concluded the following properties;

- cathode rays travel in straight lines.
- cathode rays have particle nature.
- cathode rays are made up of negatively charged particles called electrons.
- Electrons are found in all atoms.

In 1909, American physicist Robert A. Millikan determined the charge of the electron by observing how an electric field influenced the rate at which charged oil droplets fell under gravity.

- Based on careful experiments, Millikan determined the charge on an electron as $e = -1.602 \times 10^{-19}$ Coulomb(c).
- He used this value and Thomson's mass/charge ratio to calculate the mass of electron as 9.109×10^{-31} Kg.

1.3.2 Radio Activity

Does Radioactivity Support Dalton's idea of atoms?

- **Radioactivity:**- Is the spontaneous disintegration of the nuclei of unstable atoms through the emission of sub-atomic particles called: **Alpha particles, Beta particles and Gamma Rays .**
- Radioactivity is a phenomenon associated with atoms, independent of their physical or chemical state

α **–Particles**	β **–Particles**	γ **–Particles**
Is a positively charged Nuclear particle	Are very high speed electron	Are high energy electromagnetic waves
Consist of two protons and two Neutrons. simply they are ${}_2^4He^{2+}$ ion	Consist of electrons	Are mass less and charge less
Have very low penetrating power	Have higher penetrating power than α particles	Are so powerful rays similar to x-rays
Their charge is twice the charge of a proton	Are formed when the neutron is changed into proton & electron	Are found in association with alpha and Beta.

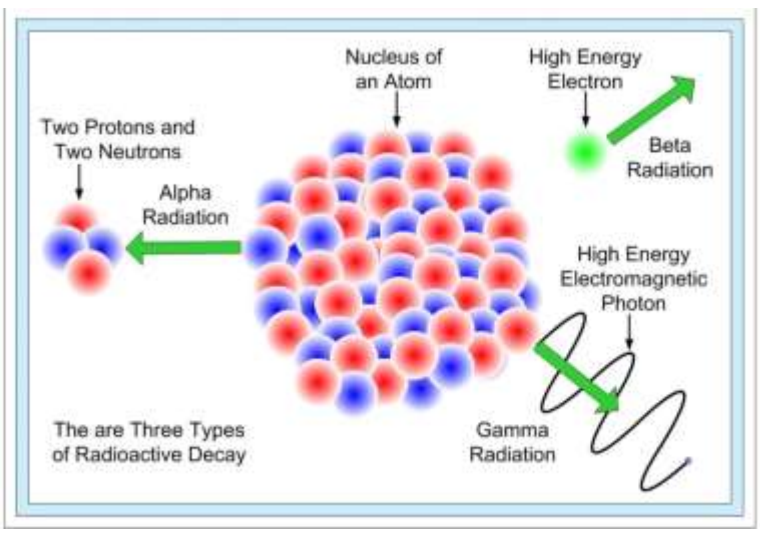

Fig.1.4 -a The Types of Radioactive Decay.

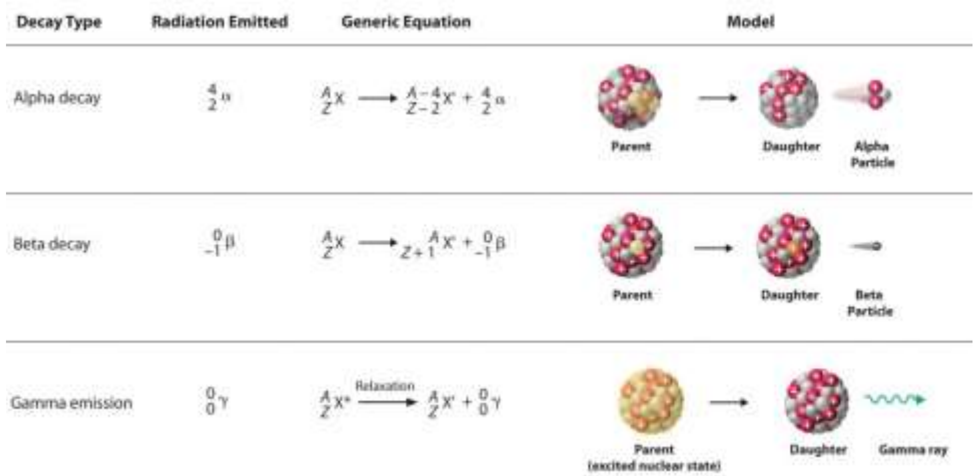

Fig.1.4 -b The Types of Radioactive Decay.

1.3.3 The Discovery of Nucleus

In 1914, Ernest Rutherford performed an experiment which involves directing α particles at a thin sheet of metal foil.

- From his experiment Rutherford concluded that all the positive charge in the atom is concentrated in the center, and he called this central positive mass of the **Nucleus**.
- The nucleus is believed to contain Proton and Neutron.
- The mass of the atom is concentrated in the nucleus because it contains the denser particles, proton and neutron.
- The charge of the nucleus is positive and its numerical value is equal to the number of protons.

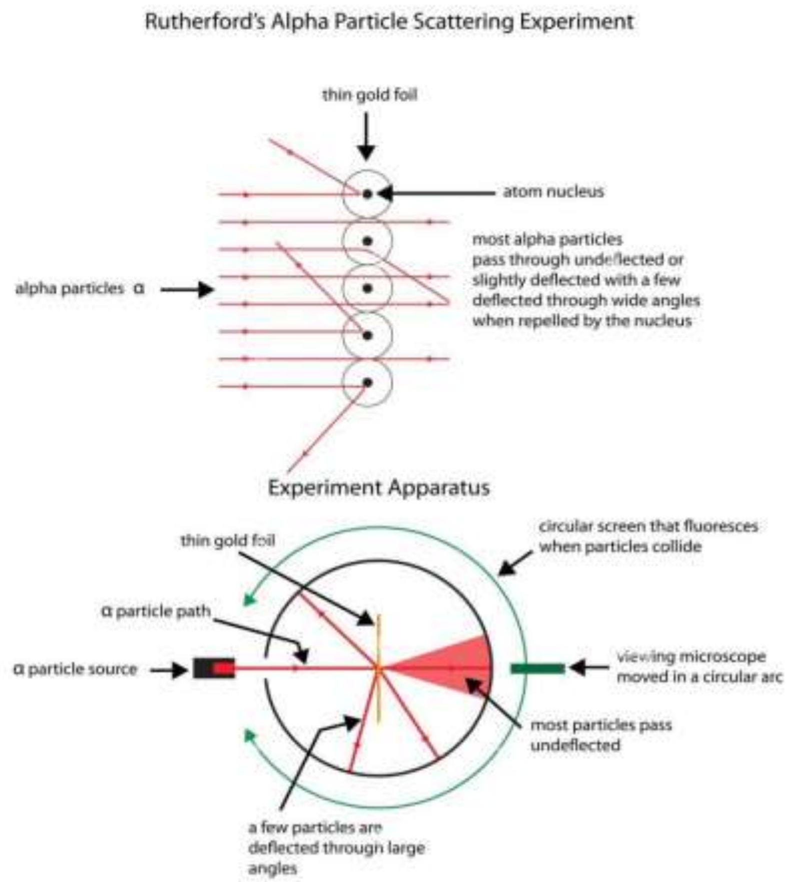

Figure 1.3.3 labelled-diagram-to-illustrate-Rutherford's-alpha-particle-scattering

1.3.4 The Discovery of Neutron

- In 1932, James Chadwick discovered the neutron by conducting experiments that involved bombarding beryllium with alpha particles. This process resulted in the emission of neutral particles, which Chadwick identified as neutrons.

- The experiment revealed an electrically neutral particles having a mass slightly greater than that of protons. Chadwick named these particles neutrons

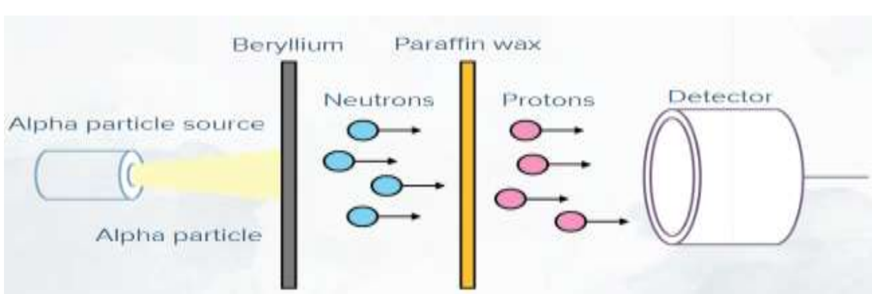

Figure 1.3.4 Discovery of the NEUTRON.

1.4 Make up of Nucleus

In 1919, Rutherford discovered that hydrogen nuclei, now known as protons, are produced when alpha particles strike lighter elements such as nitrogen. A proton is a nuclear particle with a positive charge equal in magnitude to that of an electron. It has a mass of $m_p = 1.67262\times10^{-27}$ Kg, approximately 1840 times the mass of an electron.

The protons in a nucleus contribute to the positive charge of the nucleus. Table 1.1 compares the relative masses and charges of the three subatomic particles, with "amu" denoting "atomic mass unit," which is equal to 1.

Table 1.1. Properties of subatomic particles

PARTICLE	ACTUAL MASS	RELATIVE ATOMIC MASS (a.m.u)	ACTUAL CHARGE	RELATIVE CHARGE
PROTON(p)	1.672622×10^{-27}	1.007276	$+1.602\times 10^{-19}$	+1
NEUTRON(n)	1.674927×10^{-27}	1.008665	0	0
ELECTRON(e)	9.109383×10^{-31}	5.485799×10^{-4}	-1.602×10^{-19}	-1

1.4.1 Atomic Mass & Isotope

- The number of protons found in the nucleus of an atom is called Atomic number.
- Atomic number is designated by the letter Z.
- Atomic number (Z) is Number of proton=Number of electron
- The sum of the number of protons and neutrons is known as **mass number** and designated by the letter A.
  Mass number (A)= Number of proton + Number of neutrons

ISOTOPES

- Isotopes are atoms of the same element having the same number of protons and electrons, but different number of neutrons.
- Isotopes differ in physical properties like mass, boiling point, melting point but they have similar chemical properties.
- Most elements found in nature are mixtures of isotopes.
- If an element has n isotopes with relative masses A_1, A_2 --- A_n , and fractional abundances of f_1, f_2 --- f_n , then the average relative atomic mass (A) of the element is: $\boldsymbol{A = A_1 f_1 + A_2 f_2 + --- + A_n f_n}$

EXAMPLES

1. **What is the mass number of an isotope of tin that has 66 neutrons and 50 protons?**

Solution:- A= P+ n = 66+50 = 116 amu

2. **Carbon exists as a fractional abundance of 0.9890 and a mass of exactly 12 amu, and carbon -13, with a fractional abundance of 0.0110 and a mass of 13.00335 amu. Calculate the average atomic mass of carbon?**

Solution:- $A = A_1 f_1 + A_2 f_2$ Ac **= (12 x0.9890) + (13.00335x0.0110**) = 12.006 ≈ 12.01 amu

3. Naturally occurring Boron has isotopes $_{10}B^5$ *and* 10 $_{11}B^5$ and its relative atomic mass is 10.8amu. Which isotope is more abundant?

Solution:- $Relative\ atomic\ mass = B_1 f1 + B_2 f_2$

$$10.8 = 10f_1 + 11f_2$$
$$f_1 + f_2 = 1$$
$$10.8 = 10f_1 + 11(1 - f_1)$$
$$10.8 = 10f_1 + 11 - 11f_1$$
$$10.8 - 11 = -f_1$$
$$f1 = 0.2$$

which is the abundance for $_{10}B^5$ thus,

$f_2 = 1 - 0.2 = 0.8$ is the abundance for $_{11}B^5$ therefore, $_{11}B^5$ is more abundant