Early Atomic Models

Objectives

After going through the module the learner will be will be able to:

- Know the early models of atom
- Appreciate and understand alpha particle scattering and Rutherford's Nuclear model of atom
- Interpret the results of alpha particle scattering
- Deduce the Size of nucleus
- Explain limitations of Rutherford's model

Content Outline

- Unit syllabus
- Module wise distribution of unit syllabus
- Words you must know
- Introduction
- Rutherford's atomic model
- Alpha-particle scattering and Rutherford's nuclear model of atom
- Limitations of Rutherford model of the atom
- Summary

Unit Syllabus

Chapter 12 Atoms

Alpha particle scattering experiment, Rutherford's model of atom, Bohr model, energy levels, hydrogen spectrum

Chapter 13 Nuclei

Composition and size of nucleus, radioactivity, alpha, beta and gamma particles/rays and their properties, radioactive decay laws

Mass energy relations, mass defect, binding energy per nucleon and its variation with mass number, nuclear fission and nuclear fusion

Module Distribution of Unit Syllabus - 7 Modules

Module 1	• Introduction • Early models of atom • Alpha particle scattering and Rutherford's Nuclear model of atom • Alpha particle trajectory • Results and interpretations • Size of nucleus • What Rutherford's model could not explain
Module 2	• Bohr's model of hydrogen atom • Bohr's postulates • Electron orbits, what do they look like? • Radius of Bohr orbits • Energy levels, Energy states, energy unit eV • Lowest energy -13.6 eV interpretation • Velocity of electrons in orbits
Module 3	• The line Spectrum • m of hydrogen atom • de Broglie's explanation of Bohr's second postulate of quantisation • Departures from Bohr model energy bands • Pauli's Exclusion Principle and Heisenberg's uncertainty principle leading to energy bands
Module 4	• Atomic masses and composition of nucleus • Discovery of neutron • Size of nucleus • Nuclear forces • Energy levels inside the nucleus
Module 5	• Mass and energy, Einstein's equation $E = mc^2$ • Mass defect • MeV • Nuclear binding energy • Binding energy per nucleon as a function of mass number

	• Understanding the graph and interpretations from it
Module 6	• Radioactivity • Laws of radioactivity • Half life • Rate of decay -disintegration constant • Alpha decay • Beta decay • Gamma decay
Module 7	• Nuclear energy • Fission • Controlled fission reaction • Nuclear Reactor • India atomic energy programme • Nuclear Fusion – energy generation in stars • Controlled thermonuclear fusion

You Must Know

Atoms: Atoms are the fundamental building blocks of matter. The existence of different kinds of matter is due to different atoms constituting them.

Molecules: A molecule is in general a group of two or more atoms that are chemically bonded together, that is, tightly held together by attractive forces. A molecule can be defined as the smallest particle of an element or a compound that is capable of an independent existence and shows all the properties of that substance.

The molecules of an element are constituted by the same type of atoms.

Atoms of different elements join together in definite proportions to form molecules of compounds.

Electron: The electron is a particle, with a negative charge=1.6×10^{-19} C symbol e or β^-,

Proton: A *proton* is a subatomic particle, symbol p or p^+, with a positive electric charge of +1e elementary charge and a mass slightly less than that of a neutron

Neutron: A subatomic particle of about the same mass as a proton but without an electric charge, present in all atomic nuclei except those of ordinary hydrogen.

Law of conservation of mass: states that mass can neither be created nor destroyed, it may convert to energy.

Law of constant proportions or Law of Definite Proportions

"In a chemical substance the elements are always present in definite proportions by mass"

In a pure chemical compound, elements are always present in a definite proportion by mass. This is known as the **Law of Definite Proportions**.

Dalton's law: Dalton's atomic theory which suggested that the atom was indivisible and indestructible.

Stated as follows:

- All matter is made of very tiny particles called atoms, which participate in chemical reactions.
- Atoms are indivisible particles, which cannot be created or destroyed in a chemical reaction.
- Atoms of a given element are identical in mass and chemical properties.
- Atoms of different elements have different masses and chemical properties.
- Atoms combine in the ratio of small whole numbers to form compounds.
- The relative number and kinds of atoms are constant in a given compound.

Atomic mass: One atomic mass unit is a mass equal to **exactly one-twelfth (1/12th) the mass of one atom of carbon-12**. The relative atomic masses of all elements have been found with respect to an atom of carbon-12.

Molecular mass: of a substance is the sum of the atomic masses of all the atoms in a molecule of the substance. It is therefore the relative mass of a molecule expressed in atomic mass units (u).

The mass of 1 mole of a substance: is equal to **its relative atomic or molecular mass in grams.** The atomic mass of an element gives us the mass of one atom of that element in atomic mass units (u)

Avogadro constant: 6.022×10^{23} is defined as the number of atoms in exactly 12 g of carbon-12.

The mole:the amount of substance that contains the same number of particles (atoms/ ions/ molecules, units etc.) as there are atoms in exactly 12 g of carbon-12.

Thomson's model of an atom: Thomson proposed the model of an atom to be similar to that of a Christmas pudding. The electrons, in a sphere of positive charge, were like currants (dry fruits) in a spherical Christmas pudding.

Thomson proposed that:

- An atom consists of a positively charged sphere and the electrons are embedded in it.
- The negative and positive charges are equal in magnitude. So, the atom as a whole is electrically neutral.

Rutherford's model of an atom: Ernest Rutherford was interested in knowing how the electrons are arranged within an atom. Rutherford designed an experiment for this. In this experiment, fast moving alpha (a)-particles were made to fall on a thin gold foil.

Introduction

By the nineteenth century, enough evidence had accumulated in favour of the atomic hypothesis of matter. Our current understanding of the atomic structure is very different from the evolution of quantum theory. However, we will see its development through years of scientific persistence.

Atoms and molecules are the fundamental building blocks of matter. The existence of different kinds of matter is due to different atoms constituting them. Now the questions arise:

- What makes the atom of one element different from the atom of another element?
- Are atoms really indivisible, as proposed by Dalton, or are there smaller constituents inside the atom?

It was known by 1900 that the atom was an indivisible particle but contained at least one subatomic particle – the electron identified by J.J. Thomson. Even before the electron was identified, E. Goldstein in 1886 discovered the presence of new radiations in a gas discharge and called them **canal rays**. These rays were positively charged radiations which ultimately led to the discovery of another subatomic particle. This subatomic particle had a charge, equal in magnitude but opposite in sign to that of the electron. Its mass was approximately 2000 times that of the electron. It was given the name of proton. In general, an electron is represented as 'e^-' and a proton as 'p^+'. The mass of a proton is taken as one unit and its charge as plus one. The mass of an electron is considered to be negligible and its charge is -1.6×10^{-19}C.

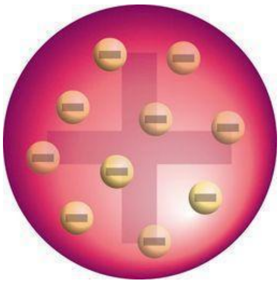
It seemed highly probable that an atom was composed of protons and electrons, mutually balancing their charges and keeping the atom neutral. It also appeared that the protons were in the interior of the atom, for electrons could easily be removed but not protons.

However, atoms on a whole are electrically neutral. .

But what is the arrangement of the positive charge and the electrons inside the atom? In other words, what is the structure of an atom?

Early Modules

The first model of atom was proposed by **J. J. Thomson in 1898**. According to this model, the positive charge of the atom is uniformly distributed throughout the volume of the atom and the negatively charged electrons are embedded in it like seeds in a watermelon. This model was picturesquely **called the plum pudding model of the atom**. In a positively charged cake the electrons were distributed randomly like raisins.



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Thomson proposed that

- An atom consists of a positively charged sphere and the electrons are embedded in it.
- The negative and positive charges are equal in magnitude. So, the atom as a whole is electrically neutral.

Although Thomson's model explained that atoms are electrically neutral.Subsequent studies on atoms showed that the distribution of the electrons and positive charges are very different from that proposed in this model.

Failure of Thomson model

It could not explain the origin of several spectral lines in the case of hydrogen and other atoms. We know that condensed matter (solids and liquids) and dense gases at all temperatures emit electromagnetic radiation in which a continuous distribution of several wavelengths is present, though with different intensities.

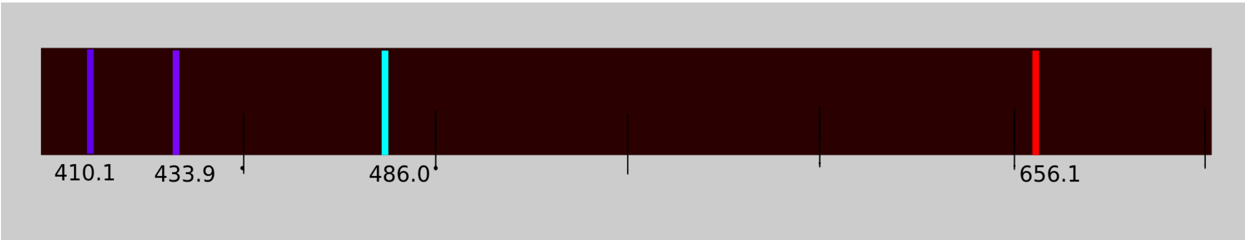
In contrast, light emitted from rarefied gases heated in a flame, or excited electrically in an evacuated tube, such as the familiar neon sign or mercury vapour light has only certain discrete wavelengths.

The spectrum appears as a series of bright lines.

In such gases, the average spacing between atoms is large. Hence, the radiation emitted can be considered due to individual atoms rather than because of interactions between atoms or molecules.

In the early nineteenth century it was also established that each element is associated with a characteristic spectrum of radiation, for example,

Hydrogen always gives a set of lines with fixed relative position between the lines.



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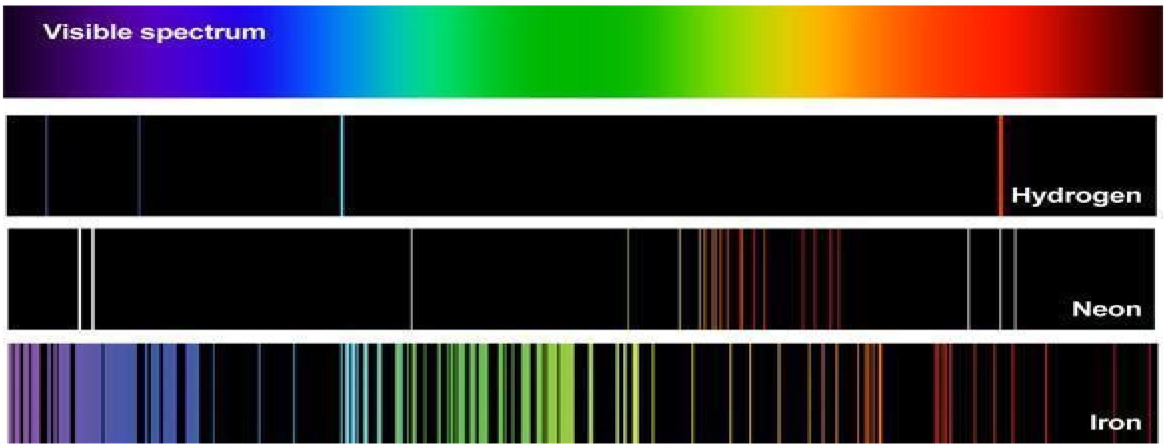
See comparative spectrum characteristic of hydrogen, neon and iron



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This fact suggested an intimate relationship between the internal structure of an atom and the spectrum of radiation emitted by it.

In 1885, Johann Jakob Balmer (1825 – 1898) obtained a simple empirical formula which gave the wavelengths of a group of lines emitted by atomic hydrogen. Since hydrogen is the simplest of the elements known, its spectrum makes very important study.

Rutherford's Atomic Model

Ernest Rutherford (1871–1937), a former research student of J. J. Thomson, was engaged in experiments on α-particles emitted by some radioactive elements. In **1906,** he proposed a classic experiment of scattering of these α-particles by atoms to investigate the atomic structure. This experiment was later performed around **1911 by Hans Geiger (1882–1945) and Ernst Marsden (1889–1970,** who was a 20-year-old student and had not yet earned his bachelor's degree.

Ernest Rutherford was interested in knowing how the electrons are arranged within an atom. Rutherford designed an experiment for this. In this experiment, fast moving alpha (α)-particles were made to fall on a thin gold foil.

- He selected a gold foil because he wanted as thin a layer as possible. This gold foil was about 1000 atoms thick.
- α-particles are doubly-charged helium ions. Since they have a mass of 4 u, the fast-moving α-particles have a considerable amount of energy.
- It was expected that α-particles would be deflected by the subatomic particles in the gold atoms. Since the α-particles were much heavier than the protons, he did not expect to see large deflections.

The explanation of the results led to the birth of Rutherford's **planetary model of atom** (also called the **nuclear model of the atom**).

According to this the entire positive charge i.e. most of the mass of the atom is concentrated in a small volume called the nucleus with electrons revolving around the nucleus just as planets revolve around the sun.

Rutherford's nuclear model was a major step towards how we see the atom today. However, it could not explain

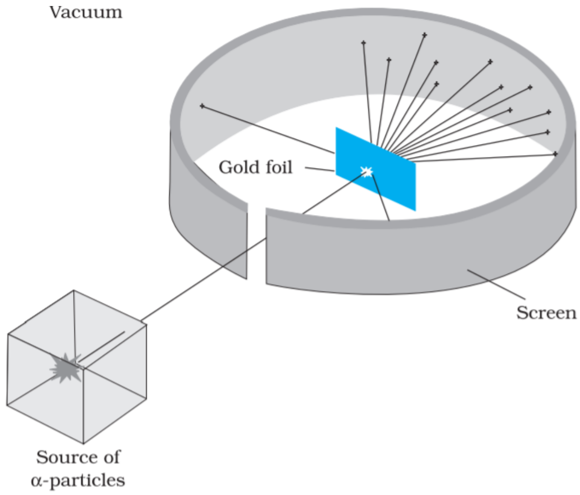
Why do atoms emit light of only discrete wavelengths?

How could an atom as simple as hydrogen, consisting of a single electron and a single proton, emit a complex spectrum of specific wavelengths?

How did Rutherford propose the model?

Alpha- Particle Scattering and Rutherford's Nuclear Model of Atom

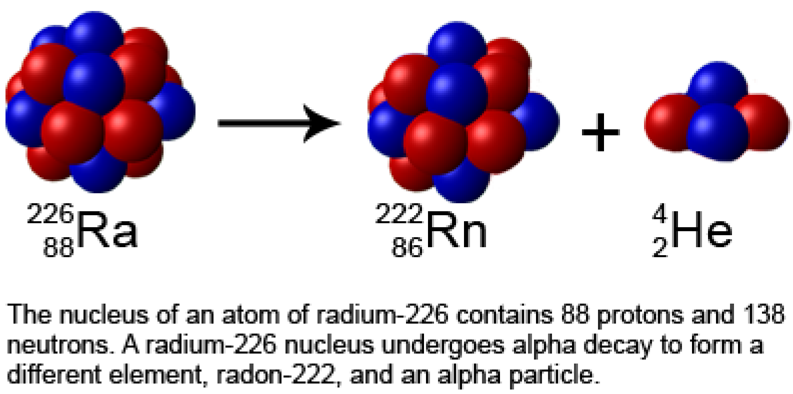
At the suggestion of Ernest Rutherford, in 1911, H. Geiger and E. Marsden performed some experiments. In one of their experiments, as shown in the figure



Geiger-Marsden scattering experiment. The entire apparatus is placed in a vacuum chamber (not shown in this figure)

An α particle is a helium ion ie a helium atom from which both the electrons have been removed. It has a charge of +2 e and its mass is four times the mass of a proton .

It is obtained during radioactive decay



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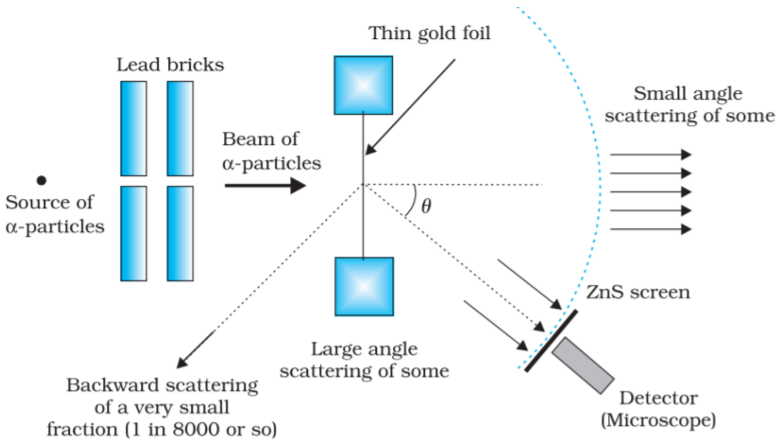
The Experiment

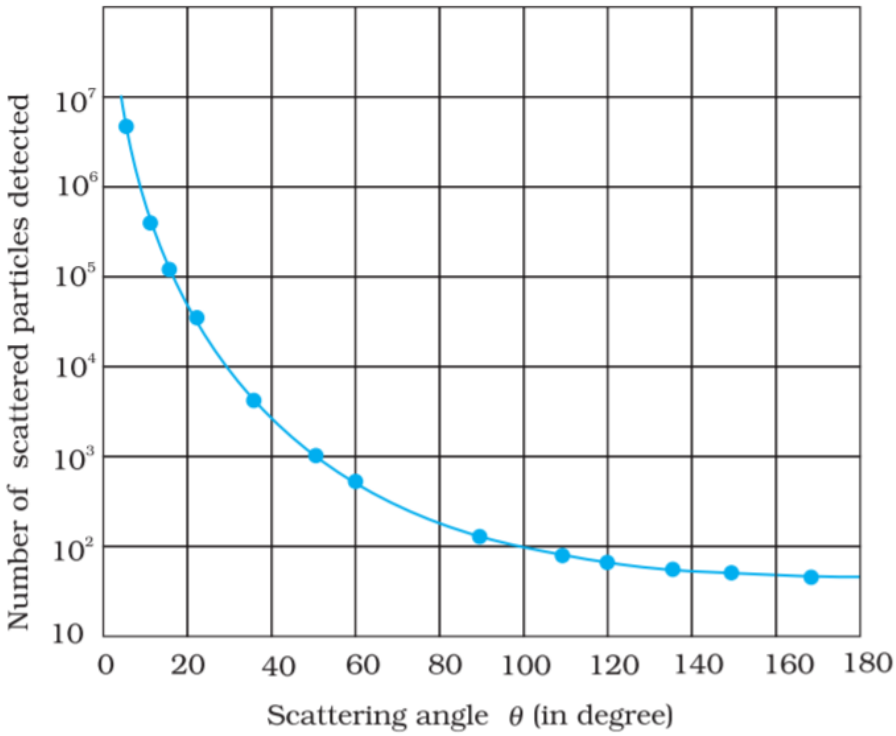
- He selected a gold foil because he wanted as thin a layer as possible. This gold foil was about 1000 atoms thick.
- α-particles are doubly-charged helium ions. Since they have a mass of 4 u, the fast-moving α-particles have a considerable amount of energy.
- It was expected that α-particles would be deflected by the subatomic particles in the gold atoms, since the α-particles were much heavier than the protons

- They directed a beam of 5.5 MeV α-particles emitted from a $85214Bi$ radioactive source at the thin gold foil. Figure shows a schematic diagram of this experiment.
- Alpha-particles emitted by $85214Bi$ radioactive sources were collimated into a narrow beam by their passage through lead bricks. The beam was allowed to fall on a thin foil of gold of thickness 2.1×10^{-7} m.
- The scattered alpha-particles were observed through a rotatable detector consisting of a zinc sulphide screen and a microscope.
- The scattered alpha-particles striking the screen produced brief light flashes or scintillations.
- These flashes were viewed through a microscope and the distribution of the number of scattered particles may be studied as a function of angle of scattering.

Analysis of The Results



Schematic arrangement of the Geiger-Marsden experiment.co relates the two figures.
A typical graph of the total number of α-particles scattered at different angles, in a given interval of time, is shown in Figure



The dots in this graph represent the data points and the solid curve is the theoretical prediction based on the assumption that the target atom has a small, dense, positively charged nucleus.

The graph shows

- Many of the α-particles pass through the foil. It means that they do not suffer any collisions.
- Only about 0.14% of the incident α-particles scatter by more than 1°; and
- About 1 in 8000 deflects by more than 90°.
- Occasionally an alpha particle gets rebounded from the gold foil, suffering a deflection of 180^0

Significance of the result

a) Rutherford argued that, to deflect the α-particle backwards, it must experience a large repulsive force.

This force could be provided if the greater part of the mass of the atom and its positive charge were concentrated tightly at its centre. Then the incoming α-particle could get very close to the positive charge without penetrating it, and such a close encounter would result in a large deflection. This agreement supported the hypothesis of the nuclear atom.

This is why Rutherford is credited with the discovery of the nucleus.

In Rutherford's nuclear model of the atom, the entire positive charge and most of the mass of the atom are concentrated in the nucleus with the electrons some distance away.

b) As most of the alpha particles pass straight through the foil, so most of the space within atoms must be empty.
c) The electrons would be moving in orbits about the nucleus just as the planets do around the sun. Rutherford's experiments suggested the size of the nucleus to be about 10^{-15} m to 10^{-14} m. From kinetic theory, the size of an atom was known to be 10^{-10} m, about 10,000 to 100,000 times larger than the size of the nucleus.

Thus, the electrons would seem to be at a distance from the nucleus of about 10,000 to 100,000 times the size of the nucleus itself.

Thus, **most of an atom is empty space.With the atom being largely empty space, it is easy to see why most α-particles go right through a thin metal foil.**

Explanation In Support Of The Theory

When an α-particle happens to come near a nucleus, the intense electric field there scatters it through a large angle. The atomic electrons, being so light, do not appreciably affect the α-particles. The scattering data shown in the graph can be analysed by employing Rutherford's nuclear model of the atom.

As the gold foil is very thin, it can be assumed that α-particles will suffer not more than one scattering during their passage through it.

Therefore, computation of the trajectory of an alpha-particle scattered by a single nucleus is enough.

Quantitatively

Alpha particles are nuclei of helium atoms and, therefore, carry two units, 2e, of positive charge and have the mass of the helium atom.

The charge of the gold nucleus is Ze, where Z is the atomic number of the atom; for gold Z = 79.

Since the nucleus of gold is about 50 times heavier than an α-particle, it is reasonable to assume that it remains stationary throughout the scattering process.

Under these assumptions, the trajectory of an alpha-particle can be computed employing Newton’s second law of motion and Coulomb's law for electrostatic force of repulsion between the alpha-particle and the positively charged nucleus.

The magnitude of this force is

$$F = \frac{1}{4\pi\epsilon_0}\frac{(2e)(Ze)}{r^2}$$

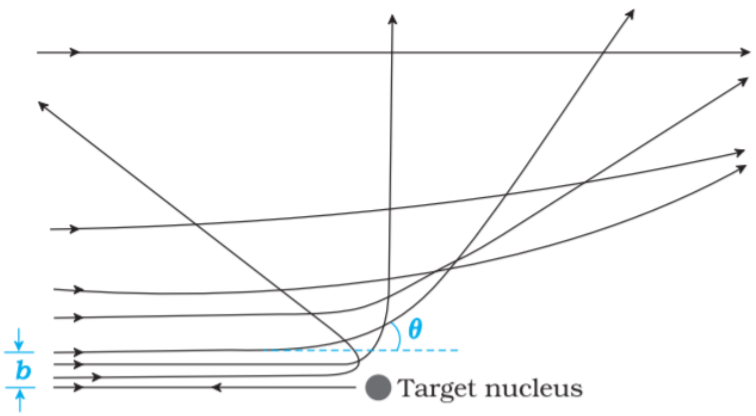
where r is the distance between the α-particle and the nucleus.

The force is directed along the line joining the α-particle and the nucleus.

The magnitude and direction of the force on an α-particle continuously changes as it approaches the nucleus and recedes away from it.

Alpha Particle Trajectory

The trajectory traced by an α-particle depends on the impact parameter, b of collision. The impact parameter is the perpendicular distance of the initial velocity vector of the α-particle from the centre of the nucleus.



Trajectory of α-particles in the coulomb field of a target nucleus. The impact parameter, b and scattering angle θ are also depicted.

A given beam of α-particles has a distribution of impact parameters b, so that the beam is scattered in various directions with different probabilities.

Impact parameter

The scattering of an alpha particle from a nucleus depends upon its distance of closest approach to the nucleus or on an equivalent length called impact parameter

In a beam, all particles have nearly the same kinetic energy. It is seen that an α-particle close to the nucleus (small impact parameter) suffers large scattering.

In case of head-on collision, the impact parameter is minimum and the α-particle rebounds back ($\theta \cong \pi$).

For a large impact parameter, the α-particle goes nearly undeviated and has a small deflection ($\theta \cong 0$). The fact that only a small fraction of the number of incident particles rebound back indicates that the number of α-particles undergoing head on collision is small. This, in turn, implies that the mass of the atom is concentrated in a small volume.

Rutherford scattering therefore, is a powerful way to determine an upper limit to the size of the nucleus.

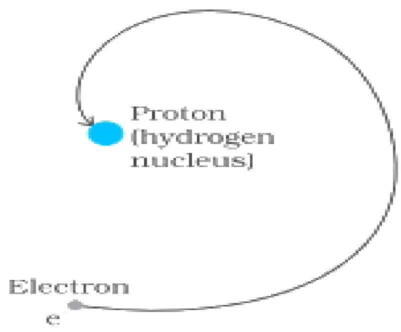
Rutherford's Model

On the basis of alpha particle scattering Rutherford proposed the following model of an atom

- An atom consists of a small and massive central core in which the entire positive charge and almost entire mass of the atom are concentrated. The core is called the nucleus.
- The size of the nucleus is very small $\sim 10^{-15}$m as compared to the size of the atom $\sim 10^{-10}$ m.
- The nucleus is surrounded by suitable number of electrons so that the atom remains neutral.
- The electrons revolve around the nucleus in orbits as the planets around the sun the centripetal force is provided by the electrostatic attraction between the electrons and the nucleus.

Limitations of Rutherford Model of the Atom

- According to electromagnetic theory an accelerated charged particle must radiate electromagnetic energy. An electron revolving around the nucleus is under continuous acceleration towards the centre as they are revolving in a circle. It should thus continuously lose energy and move in orbits with gradually decreasing radii and finally collapse into the nucleus. But the Nucleus is stable.



- An accelerated atomic electron must spiral into the nucleus as it loses energy. In Rutherford's model, an electron can revolve in orbits of all possible radii. So it should emit a continuous spectrum. But a hydrogen atom with one electron emits discrete lines in its spectrum.

Example

In Rutherford's nuclear model of the atom, the nucleus (radius about 10^{-15} m) is analogous to the sun about which the electron moves in orbit (radius ≈ 10^{-10} m) like the earth orbits around the sun. If the dimensions of the solar system had the same proportions as those of the atom, would the earth be closer to or farther away from the sun than actually it is? The radius of earth's orbit is about 1.5×10^{11} m. The radius of the sun is taken as 7×10^{8} m.

Solution

The ratio of the radius of an electron's orbit to the radius of the nucleus is $(10^{-10}\text{ m})/(10^{-15}\text{ m}) = 10^{5}$, that is, the radius of the electron's orbit is 10^5 times larger than the radius of the nucleus. If the radius of the earth's orbit around the sun were 10^5 times larger than the radius of the sun, the radius of the earth's orbit would be $10^5 \times 7 \times 10^8$ m = 7×10^{13} m.

This is more than 100 times greater than the actual orbital radius of earth. Thus, the earth would be much farther away from the sun. It implies that an atom contains a much greater fraction of empty space than our solar system does.

Example

In a Geiger-Marsden experiment, what is the distance of closest approach to the nucleus of a 7.7 MeV α-particle before it comes momentarily to rest and reverses its direction?

Solution

The key idea here is that throughout the scattering process, the total mechanical energy of the system consisting of an α-particle and a gold nucleus is conserved.

The system's initial mechanical energy is E_i, before the particle and nucleus interact, and it is equal to its mechanical energy E_f when the α-particle momentarily stops.The initial energy E_i is just the kinetic energy K of the incoming α- particle. The final energy E_f is just the electric potential energy U of the system.The potential energy U can be calculated.

Let d be the centre-to-centre distance between the α-particle and the gold nucleus when the α-particle is at its stopping point.

Then we can write the conservation of energy $E_i = E_f$ as

$$K = \frac{1}{4\pi\epsilon_0}\frac{(2e)(Ze)}{d}$$

Thus the distance of closest approach d is given by

$$d = \frac{2Ze^2}{4\pi\epsilon_0 K}$$

The maximum kinetic energy found in α-particles of natural origin is 7.7 MeV or 1.2×10^{-12} J.

Since $1/4\pi\varepsilon0 = 9.0 \times 10^9$ N m^2 /C^2.

Therefore with $e = 1.6 \times 10^{-19}$ C, we have,

$$d = \frac{(2)\left(9\times10^9 Nm^2C^{-2}\right)\left(1.6\times10^{-19}C\right)\ Z}{1.2\times10^{-12}J}$$

$$= 3.84\times10^{-16} Zm$$

The atomic number of foil material gold is Z = 79, so that d (Au) = 3.0×10^{-14} m = 30 fm.

(1 fm (i.e. fermi) = 10^{-15} m.)

The radius of the gold nucleus is, therefore, less than 3.0×10^{-14} m.

This is not in very good agreement with the observed result as the actual radius of the gold nucleus is 6 fm. The cause of discrepancy is that the distance of closest approach is considerably larger than the sum of the radii of the gold nucleus and the α-particle.

Thus, the α-particle reverses its motion without ever actually touching the gold nucleus

Summary

- Atoms, as a whole are electrically neutral and therefore contain equal amounts of positive and negative charges.
- In Thomson's model, an atom is a spherical cloud of positive charges with electrons embedded in it.
- In Rutherford's model, most of the mass of the atom and all its positive charge are concentrated in a tiny nucleus (typically one by ten thousand the size of an atom), and the electrons revolve around it.
- Rutherford nuclear model has two main difficulties in explaining the structure of atom:
  - It predicts that atoms are unstable because the accelerated electrons revolving around the nucleus must spiral into the nucleus. This contradicts the stability of matter.
  - It cannot explain the characteristic line spectra of atoms of different elements.
  - Atoms of each element are stable and emit a characteristic spectrum. The spectrum consists of a set of isolated parallel lines termed as line spectrum. It provides useful information about the atomic structure.