CHAPTER : 24

STRUCTURE OF ATOM

So far you have studied about mechanical, thermal, electrical and magnetic properties of matter. Have you ever thought as to why do different materials have different properties? That is, why does chalk break so easily but a piece of aluminium flattens on impact? Why do some metals start conducting current when light falls on them? And so on. To understand such properties of materials, we recall that atoms are building blocks of all forms of matter. That is, despite its appearance being continuous, matter has definite structure on microscopic level which is beyond the reach of our sense of seeing. This suggests that to discover answers to above said questions, you need to know the structure of the atom.

Our understanding of the structure of atom has evolved over a period of time. In this lesson, we have discussed different atomic models. Starting with Rutherford's model based on his classic scattering experiment, in this lesson we have discussed Bohr's model of atom that explains the electronic structure. Bohr's theory also helps us to explain the atomic spectrum of hydrogen atom.

OBJECTIVES

After studying this lesson, you should be able to:

- *describe Rutherford's scattering experiment and its findings;*
- *explain Rutherford's atomic model and state its shortcomings.*
- *calculate the radius of Bohr's first orbit and velocity of an electron in it;*
- *derive an expression for the energy of an electron in a hydrogen atom; and*
- *draw the energy level diagram of a hydrogen atom and explain its spectrum.*
- *describe about the production,properties, types and uses of x-rays; and*
- *define Mosley law and Duane-Hunt law.*

The Concept of Atom

The concept of atom is as old as human civilization. In ancient times Democritus in Greece and Kanad in India tried to explain the changes around us in terms of particles. But the exact theory of atom was presented by John Dalton, an English Chemist in 1808. He described atom as the smallest, indivisible particle endowed with all the properties of element and takes part in chemical reactions. Dalton's atom was an ultimate particle having no structure. This idea was accepted by the scientists in the nineteenth century as they knew nothing about the structure of atoms. The discovery of electrons by J.J. Thomson in 1897, while studying discharge of electricity through gases at low pressures, provided the first insight that atoms have a structure and negatively charged electrons are constituents of all atoms. Since the atom as a whole is neutral, it must also have equal amount of positive charge. Moreover, since electrons were thousands of times lighter than the atom, it was thought that the positively charged constituent of atoms carried the entire mass. On the basis of his experiments, Thomson suggested the *plum pudding model* of atom (Fig. 24.1). According to this, an atom is a tiny, uniformaly charged positive ball in which negatively charged electrons are suitably placed to make it neutral. It seemed perfectly reasonable then.

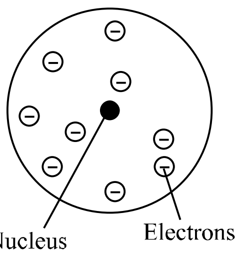


Fig. 24.1 : Plum-pudding model of atom

Our understanding of the structure of atom since the times of Thomson has improved considerably. Due to the pioneering works of Lord Rutherford, Niels Bohr, James Chadwick,Pauli, Schrodinger and others. In fact, our concept of new world of sub-atomic particles came into existence and has led to the invention of epoch making new technologies, like micro-electronics and nanotechnology.

24.1 RUTHERFORD'S EXPERIMENT ON SCATTERING OF A-PARTICLES

On the advice of Lord Rutherford, two of his students Geiger and Marsden performed an experiment in which a beam of α-particles was bombarded on a thin gold foil. The experimental arrangement used by them is shown in Fig. 24.2.

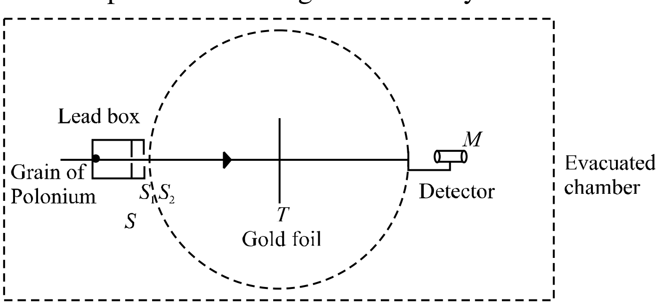


Fig. 24.2 : Schematics of experimental set up used for α-particle scattering

A well collimated fine pencil of α-particles from a source *S* was made to fall on a thin gold foil (*T*). The scattered α-particles were received on a ZnS fluorescent screen, which produced a visible flash of light when struck by an α-particles (and acted as detector), backed by a low power microscope (*M*). The detector was capable of rotation on a circular scale with *T* at the centre. The whole apparatus was enclosed in an evacuated chamber to avoid collisions of α-particles with air molecules. It was expected that if Thomson model was correct, most of the particles world go straight through the foil, with only minor deviation from the original path.

Geiger and Marsden observed that most of the α-particles suffered only small deflections, as expected. But a few got deflected at large angles (90º or more). Some of them (1 in 8000) even got deflected at 180º. Fig. 24.3 presents the experimental results. The large angle scattering of α-particles could not be explained on the basis of Thomson model of atom.

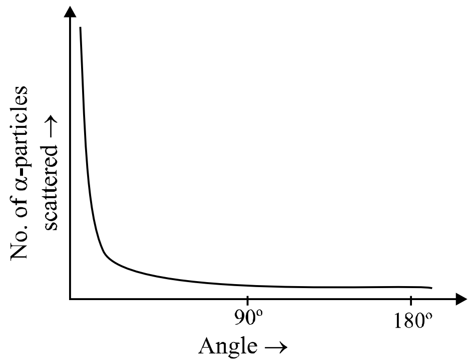


Fig. 24. 3 : Experimental result of Rutherford's experiment

According to Thomson model, α-particles should experience weak force due to electrons. However, since α-particles are about 7000 times heavier than electrons and travelled at high speed, large angle scattering strong repulsive force was required to be exerted.

To explain large angle scattering, Lord Rutherford suggested the nuclear model of atom. He argued that α-particles which pass at a large distance from the nucleus experience negligible coulombian repulsive force and hence, pass almost undeflected. However, closer to the nucleus a α-particle comes, greater force of repulsion it experiences and hence gets deflected at a greater angle. A few α-particles which proceed for a head-on collision towards the nucleus are scattered back by 180º along its direction of approach, as indicated by α-particle 4 in Fig. 24.4.

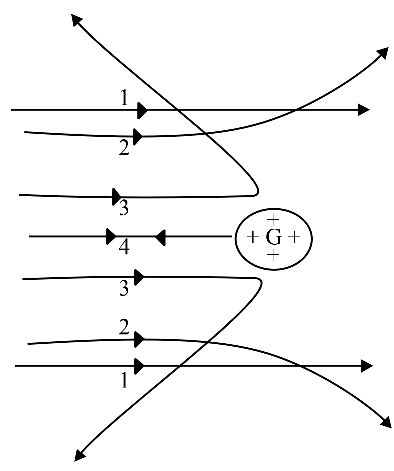
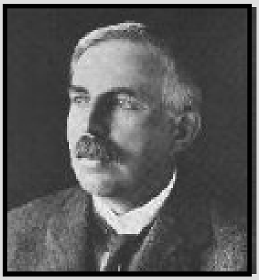


Fig. 24.4 : Paths traversed by α-particles scattered by a gold foil

Lord Rutherford

(1871–1937)



Born in New Zealand, Rutherford studied under J.J. Thomson at the Cavendish Laboratory in England. His poineering work on atom is a defining landmark. He developed Becquerel's discovery of radioactivity into an exact science and documented proof that the atoms of heavier elements, which had been thought to be immutable, actually disintegrate (decay) into various forms of radiation. In 1898, Rutherford discovered that two quite separate types of emissions came from radioactive atoms and he named them alpha and beta rays. Beta rays were soon shown to be high speed electrons. In 1907, he showed that the alpha particle was a helium atom stripped of its electrons. He and his assistant, Hans Geiger, developed Rutherford-Geiger detector to electrically detecting particles emitted by radioactive atoms. With this he could determine important physical constants such as Avogadro's number, the number of atoms or molecules in one gram-mole of material.

In 1911, Rutherford proposed the nuclear model of atom; that almost entire mass of an atom is concentrated in a nucleus 10^{-5} times the atom itself and electrons revolve around it. This second great work won him the **Nobel Prize** in chemistry in 1908.

Eminent Indian physicist, educationist and philosopher Dr. D.S. Kothari was one of his students and worked on pressure ionisation in stars.

24.1.1 Nuclear Model of Atom

Rutherford argued that large angle scattering of α-particles can be explained only by stipulating the presence of a hard, positively charged core of atom. Thus he proposed a new model of atom with following characteristics:

- The entire charge and most of the mass of the atom is confined in a very small (~10^{-15} m) central region, called the *nucleus.*
- The negatively charged electrons revolve at a distance around it such that the atom as a whole is electrically neutral and stable.

The nuclear model of atom proposed by Rutherford faced some difficulties. Some of the consequences of Rutherford's model contradicted experimental observations.

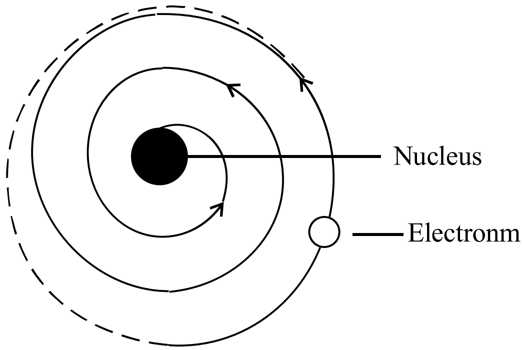


Fig. 24.5 : Motion of electrons in Rutherford's nuclear model of atom

(i) **Stability of the atom :** We know that electrons are negatively charged. These are attracted by the nucleus and get accelerated. An accelerated charged particle, according to clasical wave theory, emits elecromagnetic radiations. Hence, the revolving electrons should lose energy eventually and spiral into the nucleus (Fig. 24.5). This would have made the atom short-lived and contradicted the observed stability of matter.

(ii) **Frequency of electromagnetic radiation :** The electron spiralling towards the nucleus will emit electromagnetic radiations of all frequencies giving rise to a continuous spectrum. But experiments show that atoms emit radiations of certain well defined frequencies only (line spectra).

From the above discussion, you may be tempted to conclude that nuclear model of atom could not explain the experimental facts. Nevertheless, it contributed significantly to our understanding and was the first landmark in the right direction.

INTEXT QUESTIONS 24.1

1. Choose the correct answers :

   (a) In Rutherford's scattering experiment, target was bombarded with (i) β-rays (ii) γ -rays (iii) α-rays.

   (b) The nucleus is surrounded by :
   (i) electrons (ii) protons (iii) α-particles

   (c) The large angle scattering of α-particles indicated the presence of (i) some positively charged hard core inside the atom (ii) some porous core inside the atom (iii) negatively charged core.

2. Name two experimental observations that could not be explained by Rutherford's model.

24.2 BOHR'S MODEL OF HYDROGEN ATOM

To overcome the difficulties of Rutherford model, in 1913, Neils Bohr proposed a model of atomic structure, based on quantum ideas proposed by Max Planck. Bohr's model not only described the structure of atom but also accounted for its stability. It proved highly successful in explaining the observed spectrum of the hydrogen atom. Let us learn about it now.

Bohr's Postulates

Bohr quantised energy as well as angular momentum to explain hydrogen spectrum; Planck quantised only energy to explain black body radiation.

Bohr started with the planetary model of atom. However, to overcome the problems that plagued Rutherford model, Bohr made several assumptions. These are known as Bohr's postulates. There are four postulates.

(i) *Electrons in an atom move in circular orbits around the nucleus with the centripetal force supplied by the Coulomb force of attraction between the electron and the nucleus. Mathematically, we can write*

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r^2} \tag{24.1}$$

where Z denotes the number of positive charges in the nucleus.

(ii) *Of the infinite number of possible circular orbits, only those orbits are allowed for which the value of orbital angular momentum of the electron is an integral multiple of $h/2\pi$:*

$$|\mathbf{L}| = mvr = \frac{nh}{2\pi} \tag{24.2}$$

where **L** is the orbital angular momentum, equal to *mvr* for a circular orbit. Here *h* is Planck's constant and *n* is an integer.

(iii) *An electron moving in an allowed orbit does not radiate any energy*. In these allowed orbits, the energy of the electron is constant. These orbits are called *stationary states.*

Note that an electron can move in a stationary state but its energy is constant.

(iv) *Energy is emitted by an atom only when its electron "falls" from an allowed higher energy level E_f to another allowed lower level E_i. The change in energy is the energy of the emitted photon*. Similarly, an electron only absorbs radiation when it "jumps" to a higher energy level from a lower energy level. The change in energy of an electron can be related to the frequency or wavelength of the emitted or absorbed photon:

For emission

$$\Delta E = E_i - E_f = hv \qquad (24.3a)$$

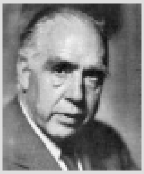
For absorption

$$\Delta E = E_f - E_i = hv \qquad (24.3b)$$

where v is the frequency of the emitted photon.

Niels Henrik David Bohr

(1885-1962)



Niels Bohr was born in Copenhagen, Denmark. He grew up in an atmosphere most favourable to the development of his genius. His father was an eminent physiologist and was largely responsible for awakening his interest in physics while he was still at school. In the spring of 1912, he worked in Rutherford's laboratory in Manchester. He studied the structure of atoms on the basis of Rutherford's nuclear model of atom. He succeeded in working out and presenting a picture of atomic structure that explained atomic spectra of hydrogen atom.

In 1916, he was appointed Professor of Theoretical Physics at Copenhagen University, and in 1920 (until his death in 1962), he became head of the Institute for Theoretical Physics, established for him at that university.

Recognition of his work on the structure of atom came with the award of the **Nobel Prize** in Physics in 1922.

Note that these postulates beautifullycombine classical and quantum ideas. For example, the first postulate is in accordance with classical physics while other postulates use quantum physics.

24.2.1 Energy Levels

To calculate the energy of an electron in nth orbit of radius r_n, we rewrite Eqn. (24.1) as

$$\frac{mv_n^2}{r_n} = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r_n^2}$$

where v_n is speed of the electron in its orbit.

On multiplying both sides of this equation by mr_n^3 we get

$$m^2 v_n^2 r_n^2 = \frac{1}{4\pi\varepsilon_0} m Z e^2 r_n$$

On combining this result with Eqn. (24.2), we get

$$m^2 v_n^2 r_n^2 = n^2 \frac{h^2}{4\pi^2} = \frac{m}{4\pi\varepsilon_0} Ze^2 r_n \qquad (24.4)$$

On re-arranging terms, we get an expression for the radius of the nth orbit :

$$r_n = 4\pi\varepsilon_o \frac{n^2h^2}{4\pi^2 mZe^2}$$

$$= \frac{n^2h^2\varepsilon_0}{Ze^2 m\pi} \qquad n = 1, 2, 3, ... \qquad (24.5)$$

Note that radius of an orbit is directly proportional to second power of the number of orbit. It means that radius is more for higher orbits. Moreover, the relative values of the radii of permitted orbits are in the ratio $1^2, 2^2, 3^2, 4^2, .. ,$ ie. in the ratio 1 : 4 : 9 : 16 and so on. For hydrogen atom (Z = 1), the radius of its inner most orbit is called **Bohr radius**. It is denoted by a_0 and its magnitude is 5.3×10^{-11}m. In terms of a_0, the radii of other orbits are given by the relation

$$r_n = n^2 a_0$$

In shows that the spacing between consecutive orbits increases progressively. On inserting the value of r_n from Eqn. (24.5) in Eqn. (24.2), we get an expression for the speed of the electron in the nth orbit :

$$\upsilon_n = \frac{nh}{2\pi m r_n} = \frac{nh}{2\pi m} \cdot \frac{Ze^2 m\pi}{n^2h^2\varepsilon_0}$$

$$= \frac{1}{2}\frac{Ze^2}{\varepsilon_0 nh} \qquad (24.6)$$

From Lesson 16 you will recall that potential energy of a negative charge (electron in this case) in bringing it from infinity to a point at a distance r in a field of positive charge (nucleus in this case) is obtained by summing (integrating) the product of Coulomb force and distance :

$$U = -\frac{1}{4\pi\varepsilon_0}\int_{r_n}^{\infty} \frac{Ze^2}{r^2} dr$$

$$= \frac{1}{4\pi\varepsilon_0} \left.\frac{Ze^2}{r}\right]_{r_n}^{\infty}$$

$$= -\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r_n} \qquad (24.7)$$

since potential energy of the electron at infinity will be zero.

It readily follows from Eqn. (24.1) that

$$\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r_n} = mv_n^2$$

Hence, potential energy of an electron in n^{th} orbit

$$U = -mv_n^2 \tag{24.8}$$

Since kinetic energy

$$K.E = \frac{1}{2}mv_n^2 \tag{24.9}$$

the total energy of the electron in nth orbit is given by

$$\begin{aligned} E &= K.E + U \\ &= \frac{1}{2}mv_n^2 - mv_n^2 \\ &= -\frac{1}{2}mv_n^2 \end{aligned}$$

Combining this result with Eqn. (24.6), we get

$$E = -\frac{m}{2}\left(\frac{2\pi Ze^2}{4\pi\varepsilon_0 nh}\right)^2$$

$$= -\frac{m}{8\varepsilon_0^2}\frac{Z^2e^4}{n^2h^2} \tag{24.10}$$

$$= \frac{RZ^2}{n^2}\ ;\ n = 1, 2, 3... \tag{24.11}$$

where

$$R = \frac{me^4}{8\varepsilon_0^2h^2} \tag{24.12}$$

is called Rydberg constant, From Eqn. (24.11) we note that

- energy of an electron in various allowed orbits is inversely proportional to the square of the number of orbit; and
- the energy in an orbit is negative, which implies that the electron is bound to the nucleus.

Putting the standard values of $m = 9.11 \times 10^{-31}$kg, $e = 1.6 \times 10^{-19}$C, $\varepsilon_0 = 0.85 \times 10^{-11}\ C^2N^1m^{-2}$, and $h = 6.62 \times 10^{-34}$ Js in Eqn. (24.12), we obtain $R = 2.17 \times 10^{-18}$ J = 13.6eV, since 1eV = 1.6×10^{-19}J. On using this result in Eqn. (24.11), we find that the energy of an electron in n^{th} orbit of hydrogen atom (in eV) is given by

$$E_n = -\frac{13.6}{n^2} \tag{24.13}$$

Thus every orbit can be specified with a definite energy; the energy of the first orbit being the lowest:

$$E_1 = -13.6 \text{ eV}$$

and the the highest energy state

$$E_\infty = 0$$

It means that different orbits represent different energy levels –13.6eV to 0. This is depicted in Fig. 24.6. $E = 0$ signifies that the electron is free.

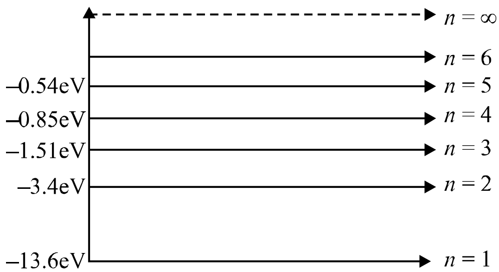


Fig. 24.6: Energy levels in hydrogen atom

According to Bohr's fourth postulate, the frequency υ_{mn} of the emitted (absorbed) radiation when the electron falls (jumps) from the nth state to the mth state is given by

$$v_{\mathrm{mn}} = \frac{RZ^2}{h}\left(\frac{1}{m^2} - \frac{1}{n^2}\right) \tag{24.14}$$

Fraunhoffer Lines

The spectrum of sunlight, when examined carefully by a high power spectroscope, is found to be crossed by a large number of dark lines spread over the length of the continuous spectrum. Wollaston observed these lines in the year 1802. But their existence was studied by Fraunhoffer on the basis of Kirchoff's laws and named these as Fraunhoffer lines. The main body of the sun emits continuous spectrum but the atomosphere of comparatively much cooler vapours and gases in the Sun's atmosphere, called the chromosphere (~6000º C), absorb light corresponding to certain wavelengths. These appear as dark lines in the continuous spectrum of the sun.

Kirchhoff compared the absorbed wavelengths with the wavelengths emitted by various elements present on the earth and identified 60 terrestrial (existing on earth) elements present in the outer atomsphere of sun, e.g. oxygen, hydrogen, sodium, iron, calcium etc.

INTEXT QUESTIONS 24.2

1. Which of Bohr's postulates "fit"with classical physics and which support the ideas of quantum physics?

2. According to Bohr, why did an atom not collapse while its electrons revolved around the nucleus?

3. According to Bohr, what is happening in the atom when a photon of light is (i) emitted (ii) absorbed?

4. Write the energy of the first three orbits of hydrogen atom on the basis of Bohr's model.

5. An atom is excited to an energy level E_1 from its ground state energy level E_0. What will be the wavelength of the radiation emitted?

6. In case of hydrogen atom, the radius of the electron in its nth orbit is proportional to

   (i) $1/n$ (ii) $1/n^2$ (iii) n (iv) n^2

7. The total energy E_n of the electron in the nth orbit of hydrogen atom is proportional to

   (i) e^4 (ii) e^3 (iii) e^2 (iv) e

24.3 HYDROGEN SPECTRUM

Refer to Fig. 24.7. It shows frequency spectrum of hydrogen atom. As may be noted, the line spectrum of hydrogen consists of many lines in different regions of the spectrum. The various lines in a particular region of spectrum are found to have a pattern and may be represented by a common formula. So they are said to form a series. Let us about the series of hydrogen spectrum

Lyman series was discovered by in 1906. Accoding to Bohr, this series arises when an electron jumps to the first orbit ($m = 1$) from an higher orbit ($n = 2, 3, 4$...). The frequencies of various spectral lines of this series are given by

$$v_{1n} \,\bar{}= \frac{R}{h}\left(\frac{1}{1^2} \quad \frac{1}{n^2}\right)$$

where n is natural number greater than one.

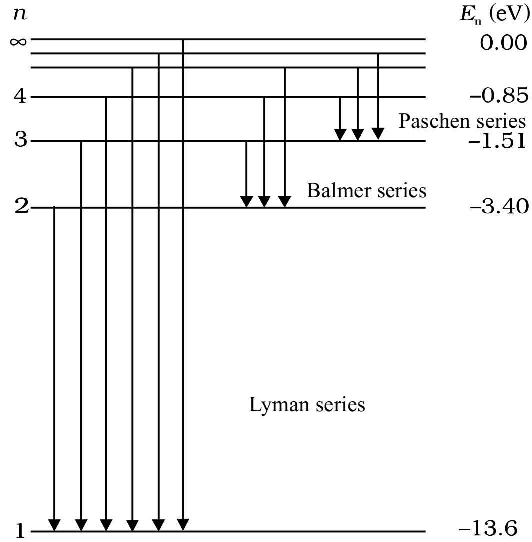


Fig. 24.7 : Energy Level diagram showing emission of various spectral series in hydrogen atom

Balmer series was discovered in 1885 in the visible region. According to Bohr, in this series, electron jumps to the second orbit ($m = 2$) from higher orbits ($n = 3, 4, 5...$). The frequencies of various spectral lines of the series are given by

$$v_{2n} = \frac{R}{h}\left(\frac{1}{2^2} - \frac{1}{n^2}\right); n > 2$$

Paschen series was discovered in 1908 in the near infra-red region. The existence of this series can be explained by assuming that electrons jump to third orbit ($m = 3$) from higher orbits ($n = 4, 5, 6...$). The frequencies of various spectral lines in the region are given by

$$v_{3n} = \frac{R}{h}\left(\frac{1}{3^2} - \frac{1}{n^2}\right); n_2 > 3$$

Brackett series was discovered in mid infra-red region. In this series, electrons jump to fourth orbit ($n = 4$) from higher orbits ($n = 5, 6...$). Therefore, the frequencies of various spectral lines in the region are given by

$$v_{4n} = \frac{R}{h}\left(\frac{1}{4^2} - \frac{1}{n^2}\right); n > 4$$

Pfund series was discovered in far infra-red region. According to Bohr, this series is obtained when electron jumps to fifth orbit ($n_1 = 5$) from any higher orbit ($n = 6, 7,...$). The frequencies of various spectral lines of the series are given by

$$v_{5n} = \frac{R}{h}\left(\frac{1}{5^2} - \frac{1}{n^2}\right); n > 5$$

The ingenuity of Bohr's model lies in the fact that it not only explained the already known spectrum but also predicted the existence of a number of series, which were observed later on. In fact, a new physics was born! Transition of the electrons from higher orbits to lower orbits showing emission of different series of spectral lines is shown in Fig. 24.8.

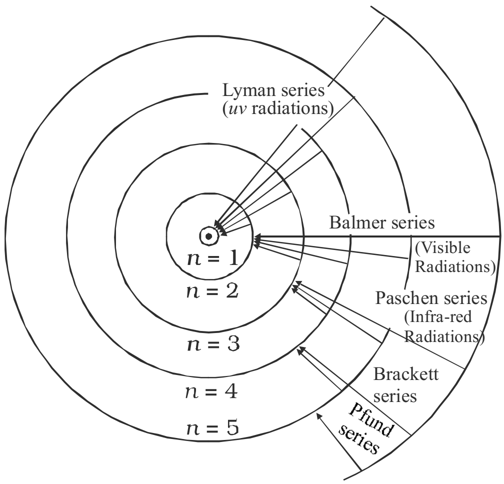


Fig. 24.8 : Permitted orbits in an atom of hydrogen and transitions leading to spectral lines of various series.

INTEXT QUESTIONS 24.3

1. The negative total energy of an orbital electron means that it a) has emitted a photon, b) is bound to the nucleus, c) is in stable equilibrium, d) satisfies Bohr's postulate

   $L = \dfrac{nh}{2\pi}$.

2. An electron jumps to the fourth orbit. When the electron jumps back to the lower energy level, the number of spectral lines emitted will be a) 6, b) 8, c) 5, d) 3.
3. Lymann series of spectral lines are emitted when electron jump from higher orbits to the .... orbit

   a) first, b) second, c) third, d) fourth.
4. Which physical property of electron was quantized by Bohr?
5. An electron jumps from third orbit to first orbit. Calculate the change in angular momentum of an electron?

24.4 X–RAYS

X–rays are produced when fast moving electrons are suddenly stopped by a heavy metal in a glass tube having extremely low pressure. The electrons emitted by the hot filament are focussed on the target which is made up of metal of high melting point and high atomic number as shown in Fig. 24.4.

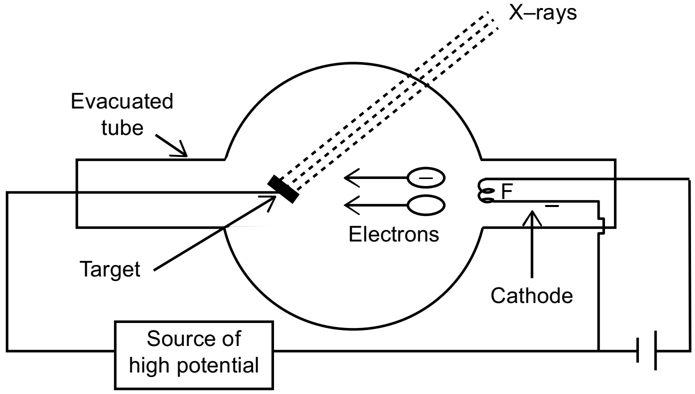


Fig. 24.4

When the electrons approach the target, 5% of their energy gets converted into X–rays and rest of the energy gets converted into heat, which is kept under control by the circulating cold water. The tube has extremely low pressure so that the electrons emitted from the hot filament (*F*) may directly hit the target without suffering collisions in between.

The intensity of the X–rays is controlled by adjusting the filament current and the quality is controlled by the accelerating voltage applied between the filament and target. This voltage usually ranges between 10 kV and 1 MV.

Properties of X–rays

X–rays show the following properties:

(i) They affect the photographic plate

(ii) They cause fluorescence in certain chemical compounds.

(iii) They ionize the gases.

(iv) They show no reflection in mirrors, no refraction in glass, no diffraction with the conventional gratings but when refined techniques are used with atomic layers of crystals, they show all these familiar phenomena of light.

(v) They do not get deviated by electric or magnetic field.

X– Rays Spectra: The element whose X–ray spectra is studied is placed at the place of target of the X–rays tube. The X–ray wavelengths are determined by the Bragg's spectrometer.

X–rays are of two types :

1. **Continuous X–rays:** All the X–ray tubes emit X–rays of all wavelengths beyond certain wavelength λ_{min}. Some of the important features of the spectrum of continuous X–rays are as follows:

   (i) The intensity of X–rays increases at all wavelengths as the voltage across the tube is increased.

   (ii) The shortest wavelength λ_{min} emitted is sharply defined and it depends on the voltage applied.

   (iii) As the voltage is increased, the wavelength at which the maximum emission occurs shifts towards the shortest wavelength side as shown in Fig. 24.5.

   Continuous X–rays are produced when the kinetic energy of the incident electrons is transformed into electromagnetic radiation upon collision with atoms. Before being stopped, electrons make several collisions and produce photons of all frequencies.

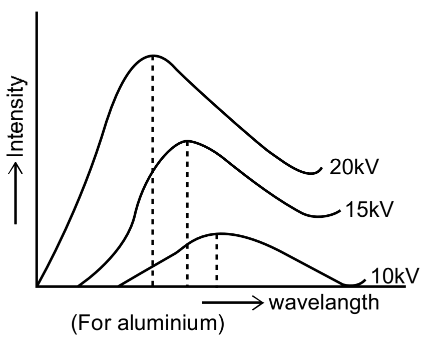


Fig. 24.5

   Photon of largest frequency is produced when an electron makes head -on collision with the atom and loses all its energy at once. For such a collision, the photon frequency ν_{max} is highest and corresponding wavelength λ_{min} is lowest.

   Thus, $$eV = h\nu_{max} = \frac{hc}{\lambda_{min}}$$

   This is known as Duane – Hunt Law.

2. **Characteristic X–rays:** In addition to continuous X–rays, X–rays tubes emit radiations which are characteristic of the target used. It is observed

that on the continuous spectrum, at certain wavelengths, large amounts of energy are radiated. The positions of these lines do not depend on the voltage applied but depend only on the nature of the target used.

Mosley's law

Mosley investigated the characteristic X–rays of a large number of elements. He found that some specific characteristic lines appeared in the spectra of all elements but at slightly differing wavelengths (Fig. 24.6). Each characteristic line obeyed a specific equation. For example, K_2 lines obey the following relation

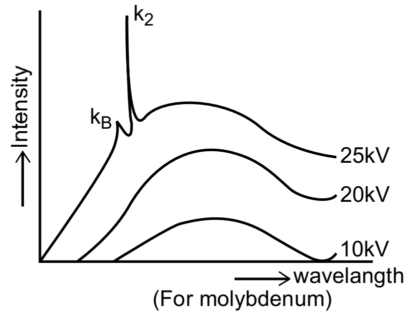


Fig. 24.6

$$\bar{\nu} = R\left[\frac{1}{1^2} - \frac{1}{2^2}\right](Z-1)^2$$

WHAT YOU HAVE LEARNT

- Rutherford's scattering experiment indicated the presence of small central region inside the atom where all the positive charge and most of the mass of the atom is concentrated. The region was named as the nucleus.
- Electrons revolve around the nucleus and total negative charge is equal to the total positive charge of the nucleus.
- Rutherford's model of atom could not explain satisfactorily the observed stability of the atom and the electromagnetic radiation emitted by the atoms.
- A satisfactory model of an atom was suggested by Niels Bohr based on four postulates.
- Permissible orbits for elecrons are those for which angular momentum $(I\omega) = nh/2\pi$
- Emission (absorption) of energy takes place when electron jumps from a higher orbit to a lower orbit (from a lower to a higher orbit).
- The radii of the permitted orbits in which the electron is free to revlove around the nucleus of the hydrogen atom are given by

$$a_n = \frac{n^2h^2}{4\pi^2\, mke^2} = \frac{n^2h^2\varepsilon_0}{Ze^2m\pi}$$

For hydrogen atom, the radius of the first permitted orbit is $a = 0.53$Å.

- The energy of the electron in the n^{th} orbit of the hydrogen atom is given by

$$E_n = -\frac{e^4 m}{8h^2\varepsilon_0^2 n^2}$$

The negative sign of total energy indicates that the electron is bound to the nucleus.

- The frequency of the photon emitted when the electron moves from the energy level E_i to E_f is given by :

$$v_{mn} = \frac{R}{h}\left[\frac{1}{m^2}-\frac{1}{n^2}\right]$$

- x-rays are produced when fast moving electrons are suddenly stopped by a heavy metal.
- x-rays are of two types (i) continuous and (ii) characteristic.
- Duane-Hunt law $eV = h\nu_{max} = \dfrac{hc}{\lambda_{min}}$
- Mosley law $\overline{\nu} = R\left(\dfrac{1}{1^2}-\dfrac{1}{2^2}\right)(z-1)^2$

ANSWERS TO INTEXT QUESTIONS

24.1

1. a (iii), b (ii), c (i)
2. It could not explain the large angle scattering of particles from the gold foils as observed by Rutherford.

24.2

1. Bohr's first postulate is from classical physics; remaining three are from quantum physics.

2. Because the orbits are stationary.

3. (i) Electron falls from higher to lower energy state.

   (ii) Electron is excited to some higher energy state.

4. $E_1 = -13.6\text{eV}$, $E_2 = 3.4$ eV, $E_3 = -1.51\text{eV}$

5. $\lambda = \dfrac{hc}{E_i - E_0}$

6. (iv)

24.3

1. (b)

2. (a) Number of spectral lines emitted $= \dfrac{1}{2} n\,(n-1) = \dfrac{1}{2} \times 4\,(4-1) = 6$

3. (a)

4. Angular momentum of revolving electron.

5. From the nth states with principal quantum number n calculate the number of wavelengths observed in the spectrum from a hydrogen sample.