UNIT 9

ATOMIC AND NUCLEAR PHYSICS

All of physics is either impossible or trivial. It is impossible until you understand it, and then it becomes trivial

– Ernest Rutherford

LEARNING OBJECTIVES

In this unit, the students are exposed to

- electric discharge through the gases
- determination of specific charge of by an electron J.J. Thomson experiment
- determination of electronic charge by Millikan's oil drop experiment
- atom models – J.J. Thomson and Rutherford
- Bohr atom model and hydrogen atom
- atomic spectrum and hydrogen spectrum
- structure and properties of nucleus
- various classification of nuclei based on atomic number and mass number
- mass defect and binding energy
- relation between stability and binding energy curve
- alpha decay, beta decay and gamma emission
- law of radioactive decay
- nuclear fission and fusion
- elementary ideas of nuclear reactors
- qualitative idea of elementary particles

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9.1 INTRODUCTION

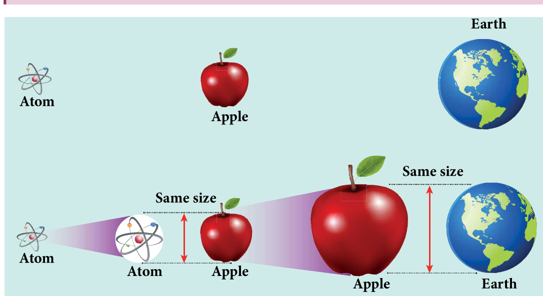

Figure 9.1 Comparision of size of an atom with that of an apple and comparision of size of an apple with that of the Earth

In earlier classes, we have studied that anything which occupies space is called matter. Matter can be classified into solids, liquids and gases. In our daily life, we use water for drinking, petrol for vehicles, we inhale oxygen, stainless steel vessels for cooking, etc. Experiences tell us that behaviour of one material is not the same as that of another, which means that the physical and chemical properties are different for different materials. In order to understand this, we need to know the fundamental constituents of materials.

When an object is divided repeatedly, the process of division could not be done beyond a certain stage in a similar way and we end up with a small speck. This small speck was defined as an atom. The word atom in Greek means 'without division or indivisible'. The size of an atom is very very small. For an example, the size of hydrogen atom (simplest among other atoms) is around 10^{-10} m. An American Physicist Richard P. Feynman said that if the size of an atom becomes the size of an apple, then the size of apple becomes the size of the earth as shown in Figure 9.1. Such a small entity is an atom.

In this unit, we first discuss the theoretical models of atom to understand its structure. The Bohr atom model is more successful than J. J. Thomson and Rutherford atom models. It explained many unsolved issues in those days and also gave better understanding of chemistry.

Later, scientists observed that even the atom is not the fundamental entity. It consists of electrons and nucleus. Around 1930, scientists discovered that nucleus is also made of proton and neutron. Further research discovered that even the proton and neutron are made up of fundamental entities known as quarks.

In this context, the remaining part of this unit is written to understand the structure and basic properties of nucleus. Further how the nuclear energy is produced and utilized are discussed.

9.2 ELECTRIC DISCHARGE THROUGH GASES

Gases at normal atmospheric pressure are poor conductors of electricity because they do not have free electrons for conduction. But by special arrangement, one can make a gas to conduct electricity.

A simple and convenient device used to study the conduction of electricity through gases is known as gas discharge tube. The arrangement of discharge tube is shown in Figure 9.2. It consists of a long closed glass tube (of length nearly 50 cm and diameter of 4 cm) inside of which a gas in pure form is filled usually. The small opening in the tube is connected to a high vacuum pump and a low-pressure gauge. This tube is fitted with two metallic plates known as electrodes which are connected to secondary of an induction coil. The electrode connected to positive of secondary is known as anode and the electrode to the negative of the secondary is cathode. The potential of secondary is maintained at about 50 kV.

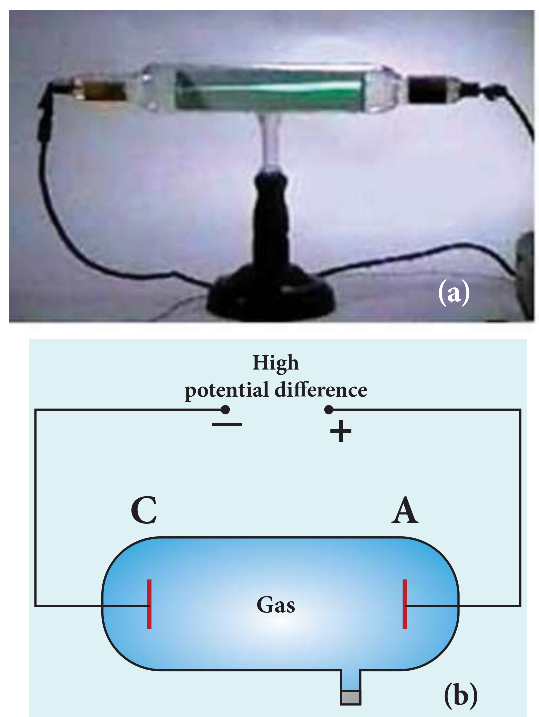

Figure 9.2 Discharge tube (a) real picture (b) schematic diagram

Suppose the pressure of the gas in discharge tube is reduced to around 110 mm of Hg using vacuum pump, it is observed that no discharge takes place. When the pressure is kept near 100 mm of Hg, the discharge of electricity through the tube takes place. Consequently, irregular streaks of light appear and also crackling sound is produced. When the pressure is reduced to the order of 10 mm of Hg, a luminous column known as positive column is formed from anode to cathode.

When the pressure reaches to around 0.01 mm of Hg, positive column disappears. At this time, a dark space is formed between anode and cathode which is often called Crooke's dark space and the walls of the tube appear with green colour. At this stage, some invisible rays emanate from cathode called cathode rays, which are later found be a beam of electrons.

Properties of cathode rays

(1) Cathode rays possess energy and momentum and travel in a straight line with high speed of the order of 10^7 ms^{-1}. It can be deflected by application of electric and magnetic fields. The direction of deflection indicates that they contain negatively charged particles.

(2) When the cathode rays are allowed to fall on matter, heat is produced. Cathode rays affect the photographic plates and also produce fluorescence when they fall on certain crystals and minerals.

(3) When the cathode rays fall on a material of high atomic weight, x-rays are produced.

(4) Cathode rays ionize the gas through which they pass.

(5) The speed of cathode rays is up to $\left(\frac{1}{10}\right)^{th}$ of the speed of light.

9.2.1 Determination of specific charge $\left(\frac{e}{m}\right)$ of an electron – Thomson's experiment

Thomson's experiment is considered as one among the landmark experiments for the birth of modern physics. In 1887, J. J. Thomson made remarkable improvement in the study of gases in discharge tubes. In the presence of electric and magnetic fields, the cathode rays were deflected. By the variation of electric and magnetic fields, the specific charge (charge per unit mass) of the cathode rays is measured.

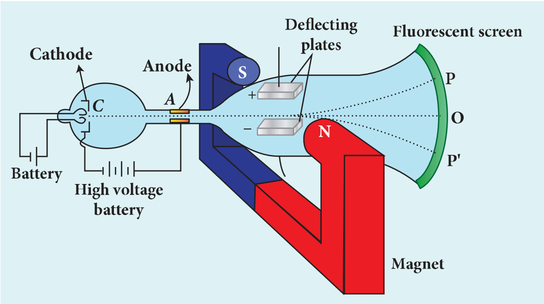

Figure 9.3 Arrangement of J.J. Thomson experiment to determine the specific charge of an electron

The arrangement of J. J. Thomson's experiment is shown in Figure 9.3. A highly evacuated discharge tube is used and cathode rays (electron beam) produced at cathode are attracted towards anode disc A. Anode disc is provided with pin hole in order to allow only a narrow beam of cathode rays. These cathode rays are now allowed to pass through the parallel metal plates which are maintained at high voltage as shown in Figure 9.3. Further, the gas discharge tube is kept in between pole pieces of magnet such that both electric and magnetic fields are acting perpendicular to each other. When the cathode rays strike the screen, they produce scintillation and hence bright spot is observed. This is achieved by coating the screen with zinc sulphide.

(i) Determination of velocity of cathode rays

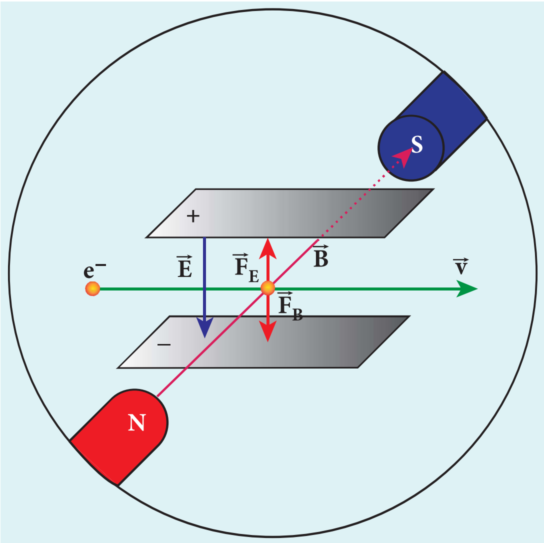

Figure 9.4 Electric force balancing the magnetic force – the path of electron beam is a straight line

For a fixed electric field between the plates, the magnetic field is adjusted such that the cathode rays (electron beam) strike at the original position O (Figure 9.3). This means that the magnitude of electric force is balanced by the magnitude of force due to magnetic field as shown in Figure 9.4. Let e be the charge of the cathode rays, then

$$eE = eB\upsilon$$

$$\Rightarrow \upsilon = \frac{E}{B} \quad (9.1)$$

(ii) Determination of specific charge

Since the cathode rays (electron beam) are accelerated from cathode to anode, the potential energy of the electron beam at the cathode is converted into kinetic energy of the electron beam at the anode. Let V be the potential difference between anode and cathode, then the potential energy is eV. Then from law of conservation of energy,

$$eV = \frac{1}{2}m\upsilon^2 \Rightarrow \frac{e}{m} = \frac{\upsilon^2}{2V}$$

Substituting the value of velocity from equation (9.1), we get

$$\frac{e}{m} = \frac{1}{2V}\frac{E^2}{B^2} \quad (9.2)$$

Substituting the values of E, B and V, the specific charge can be determined as

$$\frac{e}{m} = 1.7\times10^{11}\,\mathrm{C\,kg^{-1}}$$

(iii) Deflection of charge only due to uniform electric field

When the magnetic field is turned off, the deflection is only due to electric field. The deflection in vertical direction is due to the electric force.

$$F_e = eE \quad (9.3)$$

Let m be the mass of the electron and by applying Newton's second law of motion, acceleration of the electron is

$$a_e = \frac{1}{m}F_e \quad (9.4)$$

Substituting equation (9.4) in equation (9.3),

$$a_e = \frac{1}{m}eE = \frac{e}{m}E$$

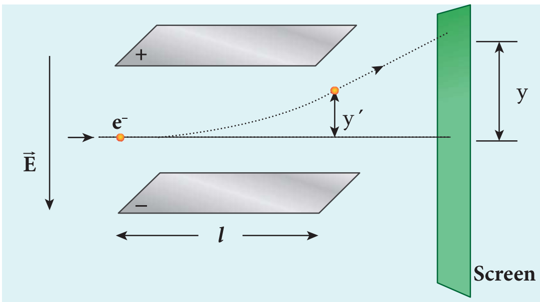

Figure 9.5 Deviation of path by applying uniform electric field

Let y be the deviation produced from original position on the screen as shown in Figure 9.5. Let the initial upward velocity of cathode ray be $u = 0$ before entering the parallel electric plates. Let t be the time taken by the cathode rays to travel in electric field. Let l be the length of one of the plates, then the time taken is

$$t = \frac{l}{\upsilon} \qquad (9.5)$$

Hence, the deflection y' of cathode rays is (note: $u = 0$ and $a_e = \frac{e}{m}E$)

$$y' = ut + \frac{1}{2}at^2 \Rightarrow y' = ut + \frac{1}{2}a_e t^2$$

$$= \frac{1}{2}\left(\frac{e}{m}E\right)\left(\frac{l}{\upsilon}\right)^2$$

$$y' = \frac{1}{2}\frac{e}{m}\frac{l^2B^2}{E} \qquad (9.6)$$

Therefore, the deflection y on the screen is

$$y \propto y' \Rightarrow y = Cy'$$

where C is proportionality constant which depends on the geometry of the discharge tube and substituting y' value in equation 9.6, we get

$$y = C\frac{1}{2}\frac{e}{m}\frac{l^2B^2}{E} \qquad (9.7)$$

Rearranging equation (9.7) as

$$\frac{e}{m} = \frac{2yE}{Cl^2B^2} \qquad (9.8)$$

Substituting the values on RHS, the value of specific charge is calculated as $\frac{e}{m} = 1.7 \times 10^{11}\,\mathrm{C\,kg^{-1}}$.

Note

The specific charge is independent of
(a) gas used
(b) nature of the electrodes

9.2.2 Determination of charge of an electron – Millikan's oil drop experiment

Millikan's oil drop experiment is another important experiment in modern physics which is used to determine one of the fundamental constants of nature known as charge of an electron (Figure 9.6 (a)).

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By adjusting electric field suitably, the motion of oil drop inside the chamber can be controlled – that is, it can be made to move up or down or even kept balanced in the field of view for sufficiently long time.

The experimental arrangement is shown in Figure 9.6 (b). The apparatus consists of two horizontal circular metal plates A and B each with diameter around 20 cm and are separated by a small distance 1.5 cm. These two parallel

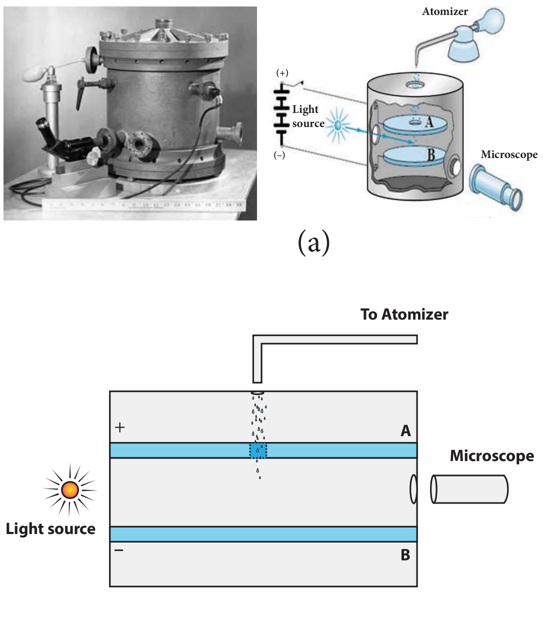


Figure 9.6 Millikan's experiment (a) real picture and schematic picture (b) Side view picture

plates are enclosed in a chamber with glass walls. Further, plates A and B are maintained at high potential difference around 10 kV such that electric field acts vertically downward. A small hole is made at the centre of the upper plate A and an atomizer is kept exactly above the hole to spray the liquid. When a fine droplet of the highly viscous non volatile liquid (like glycerine) is sprayed using atomizer, they fall freely downward through the hole of the top plate only under the influence of gravity.

Few oil drops in the chamber can acquire electric charge (negative charge) because of friction with air or passage of x-rays in between the parallel plates. Further the chamber is illuminated by light which is passed horizontally and oil drops can be seen clearly using microscope placed perpendicular to the light beam. These drops can move either upwards or downward.

Let m be the mass of the oil drop and q be its charge. Then the forces acting on the droplet are

(a) gravitational force $F_g = mg$
(b) electric force $F_e = qE$
(c) buoyant force F_b
(d) viscous force F_v

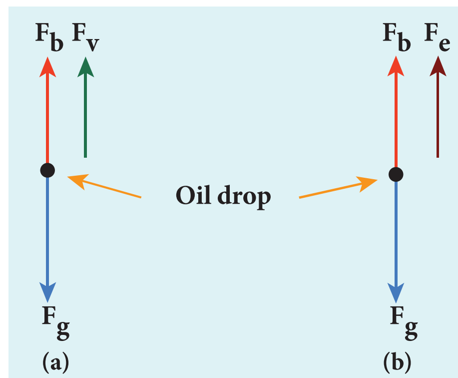


Figure 9.7 Free body diagram of the oil drop – (a) without electric field (b) with electric field

(a) Determination of radius of the droplet

When the electric field is switched off, the oil drop accelerates downwards. Due to the presence of air drag forces, the oil drops easily attain its terminal velocity and moves with constant velocity. This velocity can be carefully measured by noting down the time taken by the oil drop to fall through a predetermined distance. The free body diagram of the oil drop is shown in Figure 9.7 (a), we note that viscous force and buoyant force balance the gravitational force.

Let the gravitational force acting on the oil drop (downward) be $F_g = mg$

Let us assume that oil drop to be spherical in shape. Let ρ be the density of the oil drop, and r be the radius of the oil drop, then the mass of the oil drop can be expressed in terms of its density as

$$\rho = \frac{m}{V}$$

$$\Rightarrow m = \rho\left(\frac{4}{3}\pi r^3\right) \qquad \left(\because \text{volume of the sphere, } V = \frac{4}{3}\pi r^3\right)$$

The gravitational force can be written in terms of density as

$$F_g = mg \Rightarrow F_g = \rho\left(\frac{4}{3}\pi r^3\right)g$$

Let σ be the density of the air, the upthrust force experienced by the oil drop due to displaced air is

$$F_b = \sigma\left(\frac{4}{3}\pi r^3\right)g$$

Once the oil drop attains a terminal velocity υ, the net downward force acting on the oil drop is equal to the viscous force acting opposite to the direction of motion of the oil drop. From Stokes law, the viscous force on the oil drop is

$$F_v = 6\pi r\upsilon\eta$$

From the free body diagram as shown in Figure 9.7 (a), the force balancing equation is

$$F_g = F_b + F_\upsilon$$

$$\rho\left(\frac{4}{3}\pi r^3\right)g = \sigma\left(\frac{4}{3}\pi r^3\right)g + 6\pi r\upsilon\eta$$

$$\frac{4}{3}\pi r^3(\rho - \sigma)g = 6\pi r\upsilon\eta$$

$$\frac{2}{3}\pi r^3(\rho - \sigma)g = 3\pi r\upsilon\eta$$

$$r = \left[\frac{9\eta\upsilon}{2(\rho - \sigma)g}\right]^{\frac{1}{2}} \qquad (9.9)$$

Thus, equation (9.9) gives the radius of the oil drop.

(b) Determination of electric charge

When the electric field is switched on, charged oil drops experience an upward electric force (qE). Among many drops, one particular drop can be chosen in the field of view of microscope and strength of the electric field is adjusted to make that particular drop to be stationary. Under these circumstances, there will be no viscous force acting on the oil drop. Then, from the free body diagram shown Figure 9.7 (b), the net force acting on the oil droplet is

$$F_e + F_b = F_g$$

$$\Rightarrow qE + \frac{4}{3}\pi r^3\sigma g = \frac{4}{3}\pi r^3\rho g$$

$$\Rightarrow qE = \frac{4}{3}\pi r^3(\rho - \sigma)g \qquad (9.10)$$

$$\Rightarrow q = \frac{4}{3E}\pi r^3(\rho - \sigma)g \qquad (9.11)$$

Substituting equation (9.9) in equation (9.11), we get

$$q = \frac{18\pi}{E}\left(\frac{\eta^3\upsilon^3}{2(\rho - \sigma)g}\right)^{\frac{1}{2}} \qquad (9.12)$$

Millikan repeated this experiment several times and computed the charges on oil drops. He found that the charge of any oil drop can be written as integral multiple of a basic value, -1.6×10^{-19} C, which is nothing but the charge of an electron.

9.3 ATOM MODELS

Introduction

Around 400 B.C, Greek philosophers Leucippus and Democretus proposed the concept of atom, 'Every object on continued subdivision ultimately yields atoms'. Later, many physicists and chemists tried to understand the nature with the idea of atoms. Many theories were proposed to explain the properties (physical and chemical) of bulk materials on the basis of atomic model.

For instance, J. J. Thomson proposed a theoretical atom model which is based on static distribution of electric charges. Since this model fails to explain the stability of atom, one of his students E. Rutherford proposed the first dynamic model of an atom. Rutherford gave atom model which is based on results of an experiment done by his students (Geiger and Marsden). But this model also failed to explain the stability of the atom.

Later, Niels Bohr who is also a student of Rutherford proposed an atomic model for hydrogen atom which is more successful than other two models. Niels Bohr atom model could explain the stability of the atom and also the origin of line spectrum. There are other atom models, such as Sommerfeld's atom model and atom model from wave mechanics (quantum mechanics). But we will restrict ourselves only to very simple (mathematically simple) atom model in this section.

9.3.1 J. J. Thomson's Model (Water melon model)

In this model, the atoms are visualized as homogeneous spheres which contain uniform distribution of positively charged particles (Figure 9.8 (a)). The negatively charged particles known as electrons are embedded in it like seeds in water melon as shown in Figure 9.8 (b).

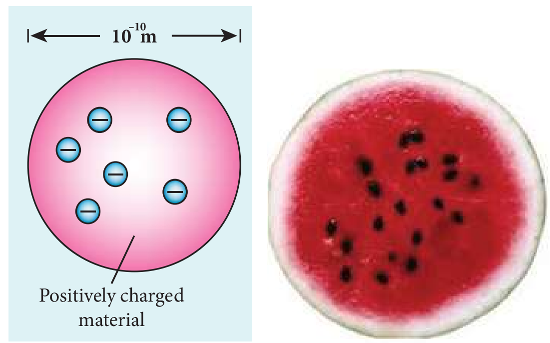


Figure 9.8 (a) Atom (b) Water melon

The atoms are electrically neutral, this implies that the total positive charge in an atom is equal to the total negative charge. According to this model, all the charges are assumed to be at rest. But from classical electrodynamics, no stable equilibrium points exist in electrostatic configuration (this is known as Earnshaw's theorem) and hence such an atom cannot be stable. Further, it fails to explain the origin of spectral lines observed in the spectrum of hydrogen atom and other atoms.

9.3.2 Rutherford's model

In 1911, Geiger and Marsden did a remarkable experiment based on the advice of their teacher Rutherford, which is known as scattering of alpha particles by gold foil.

The experimental arrangement is shown in Figure 9.9. A source of alpha particles (radioactive material, example polonium) is

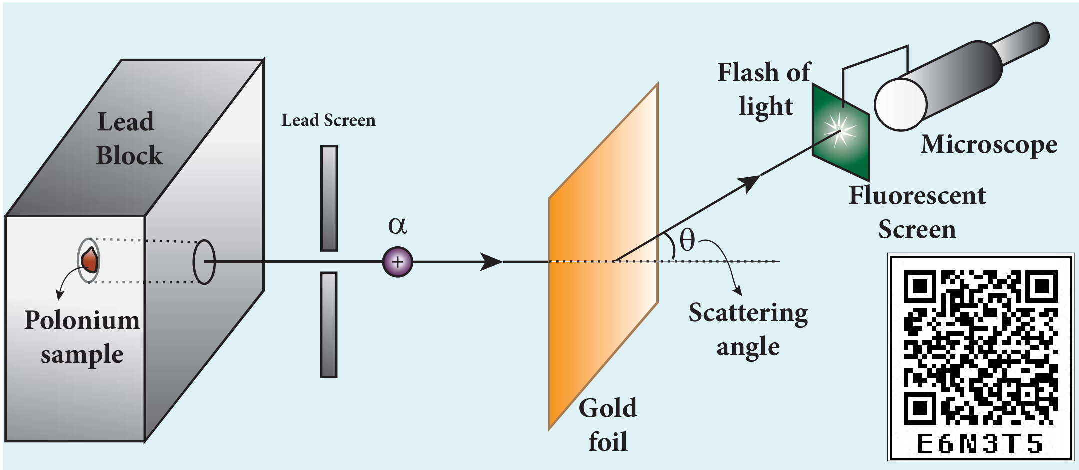


Figure 9.9 Schematic diagram for scattering of alpha particles experiment by Rutherford

kept inside a thick lead box with a fine hole as seen in Figure 9.9. The alpha particles coming through the fine hole of lead box pass through another fine hole made on the lead screen. These particles are now allowed to fall on a thin gold foil and it is observed that the alpha particles passing through gold foil are scattered through different angles. A movable screen (from 0° to 180°) which is made up of zinc sulphide (ZnS) is kept on the other side of the gold foil to collect the scattered alpha particles. Whenever alpha particles strike the screen, a flash of light is observed which can be seen through a microscope.

Rutherford proposed an atom model based on the results of alpha scattering experiment. In this experiment, alpha particles (positively charged particles) were allowed to fall on the atoms of a metallic gold foil. The results of this experiment are given below and are shown in Figure 9.10, Rutherford expected the atom model to be as seen in Figure 9.10 (a) but the experiment showed the model as in Figure 9.10 (b).

(a) Most of the alpha particles were un-deflected through the gold foil and went straight.

(b) Some of the alpha particles were deflected through a small angle.

(c) A few alpha particles (one in thousand) were deflected through the angle more than 90°

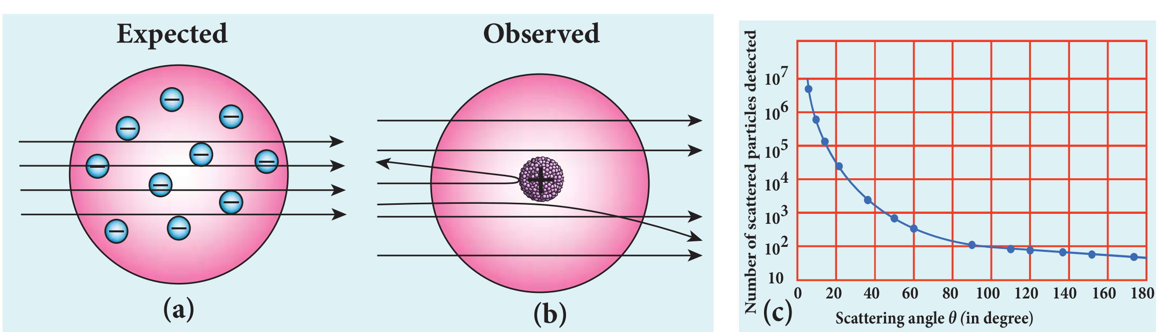


Figure 9.10 In alpha scattering experiment – (a) Rutherford expected (b) experiment result (c) The variation of alpha particles scattered N(θ) with scattering angle θ

(d) Very few alpha particles returned back (back scattered) –that is, deflected back by 180°

In Figure 9.10 (c), the dotted points are the alpha scattering experiment data points obtained by Geiger and Marsden and the solid curve is the prediction from Rutherford's nuclear model. It is observed that the Rutherford's nuclear model is in good agreement with the experimental data.

Conclusion made by Rutherford based on the above observation

From the experimental observations, Rutherford proposed that an atom has a lot of empty space and contains a tiny matter at its centre known as nucleus whose size is of the order of 10^{-14}m. The nucleus is positively charged and most of the mass of the atom is concentrated in the nucleus. The nucleus is surrounded by negatively charged electrons. Since static charge distribution cannot be in a stable equilibrium, he suggested that the electrons are not at rest and they revolve around the nucleus in circular orbits like planets revolving around the sun.

(a) Distance of closest approach

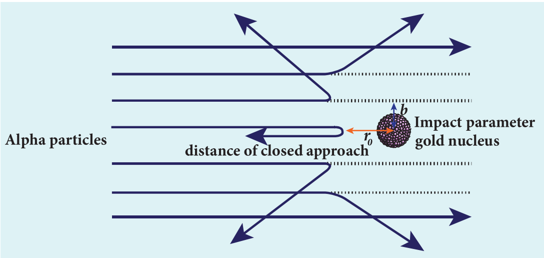


Figure 9.11 Distance of closest approach and impact parameter

When an alpha particle moves straight towards the nucleus, it reaches a point where it comes to rest momentarily and returns back as shown in Figure 9.11. **The minimum distance between the centre of the nucleus and the alpha particle just before it gets reflected back through 180° is defined as the distance of closest approach r_0 (also known as contact distance).** At this distance, all the kinetic energy of the alpha particle will be converted into electrostatic potential energy (Refer unit 1, volume 1 of +2 physics text book).

$$\frac{1}{2}mv_0^2 = \frac{1}{4\pi\varepsilon_0}\frac{(2e)(Ze)}{r_0}$$

$$\Rightarrow r_0 = \frac{1}{4\pi\varepsilon_0}\frac{2Ze^2}{\left(\frac{1}{2}mv_0^2\right)} = \frac{1}{4\pi\varepsilon_0}\frac{2Ze^2}{E_k}$$

where E_k is the kinetic energy of the alpha particle.This is used to estimate the size of the nucleus but size of the nucleus is always lesser than the distance of closest approach. Further, Rutherford calculated the radius of the nucleus for different nuclei and found that it ranges from 10^{-14} m to 10^{-15} m.

(b) Impact parameter

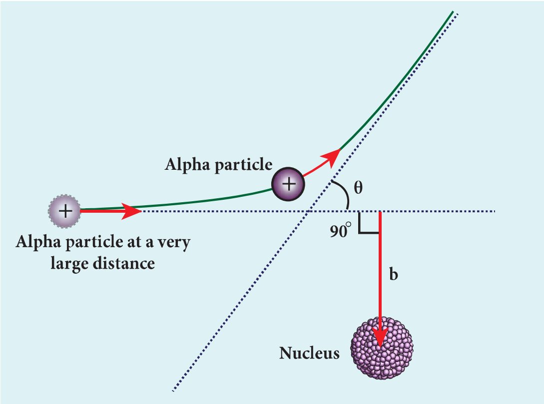


Figure 9.12 Impact parameter

The impact parameter (b) (see Figure 9.12) **is defined as the perpendicular distance between the centre of the gold nucleus and the direction of velocity vector of alpha particle when it is at a large distance**. The relation between impact parameter and scattering angle can be shown as

$$b \propto \cot\left(\frac{\theta}{2}\right) \Rightarrow b = K\cot\left(\frac{\theta}{2}\right) \qquad (9.13)$$

where $K = \frac{1}{4\pi\varepsilon_0}\frac{2Ze^2}{m\upsilon_0^2}$ and θ is called scattering angle. Equation (9.13) implies that when impact parameter increases, the scattering angle decreases. Smaller the impact parameter, larger will be the deflection of alpha particles.

Drawbacks of Rutherford model

Rutherford atom model helps in the calculation of the diameter of the nucleus and also the size of the atom but has the following limitations:

(a) This model fails to explain the distribution of electrons around the nucleus and also the stability of the atom.

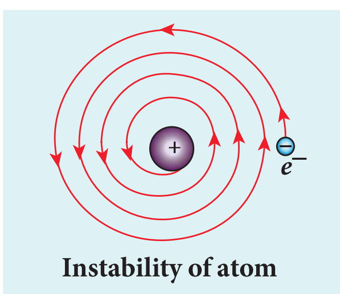


Figure 9.13 Spiral in motion of an electron around the nucleus

According to classical electrodynamics, any accelerated charge should emit electromagnetic radiations continuously. Due to emission of radiations, the charge loses its energy. Hence, it can no longer sustain the circular motion. The radius of the orbit, therefore, becomes smaller and smaller (undergoes spiral motion) as shown in Figure 9.13 and finally the electron should fall into the nucleus and the atoms should disintegrate. But this does not happen. Hence, Rutherford model could not account for the stability of atoms.

(b) According to this model, emission of radiation must be continuous and must give continuous emission spectrum but experimentally we observe only line (discrete) emission spectrum for atoms.

9.3.3 Bohr atom model

In order to overcome the limitations of the Rutherford atom model in explaining the stability and also the line spectrum observed for a hydrogen atom (Figure 9.14), Niels Bohr made modifications in Rutherford atom model. He is the first person to give better theoretical model of the structure of an atom to explain the line spectrum of hydrogen atom. The following are the assumptions (postulates) made by Bohr.

Figure 9.14 The line spectrum of hydrogen

Postulates of Bohr atom model:

(a) The electron in an atom moves around nucleus in circular orbits under the influence of Coulomb electrostatic force of attraction. This Coulomb force gives necessary centripetal force for the electron to undergo circular motion.

(b) Electrons in an atom revolve around the nucleus only in certain discrete orbits called stationary orbits and electron in such orbits do not radiate electromagnetic energy. Only those discrete orbits allowed are stable orbits.

The angular momentum of the electron in these stationary orbits are quantized – that is, it can be written as an integer or integral multiple of $\frac{h}{2\pi}$ called as reduced Planck's constant – that is, $\hbar$ (read it as h-bar) and the integer n is called as principal quantum number.

$$l = n\hbar \qquad \text{where } \hbar = \frac{h}{2\pi}$$

This condition is known as angular momentum quantization condition.

According to quantum mechanics, particles like electrons have dual nature (Refer unit 7, volume 2 of +2 physics text book). The standing wave pattern of the de Broglie wave associated with orbiting electron in a stable orbit is shown in Figure 9.15.

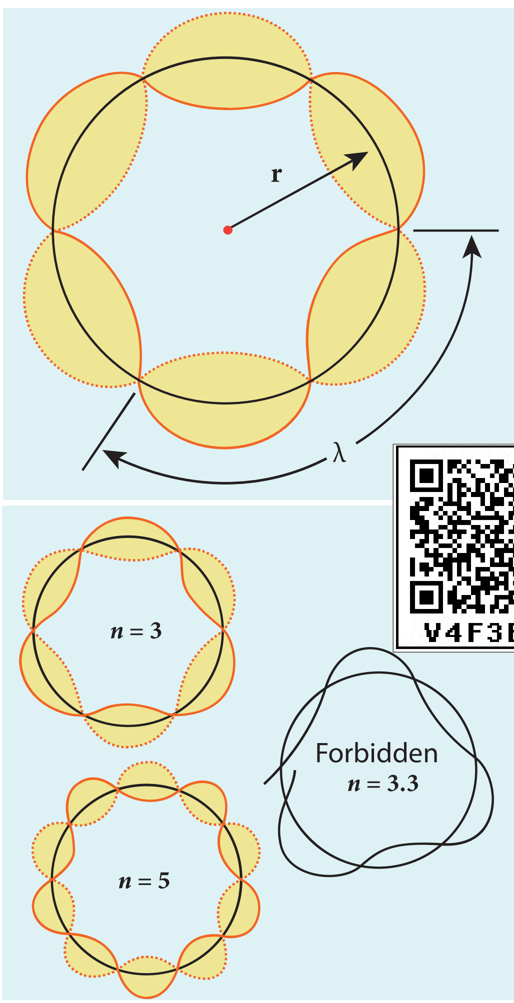


Figure 9.15 Standing wave pattern for electron in a stable orbit

The circumference of an electron's orbit of radius r must be an integral multiple of de Broglie wavelength – that is,

$$2\pi r = n\lambda \qquad (9.14)$$

where n = 1,2,3,......

But the de Broglie wavelength (λ) associated with an electron of mass m moving with velocity υ is $\lambda = \frac{h}{m\upsilon}$ where h is called Planck's constant. Thus from equation (9.14),

$$2\pi r = n\left(\frac{h}{m\upsilon}\right)$$

$$m\upsilon r = n\frac{h}{2\pi}$$

For any particle of mass m undergoing circular motion with radius r and velocity υ, the magnitude of angular momentum l is given by

$$l = r(m\upsilon)$$

$$m\upsilon r = l = n\hbar$$

(c) Energy of the electron in orbits is not continuous but only discrete. This is called the quantization of energy. An electron can jump from one orbit to another orbit by absorbing or emitting a photon whose energy is equal to the difference in energy (ΔE) between the two orbital levels (Figure 9.16)

$$\Delta E = E_{final} - E_{initial} = h\nu = h\frac{c}{\lambda}$$

where c is the speed of light and λ is the wavelength and ν is the frequency of the radiation emitted. Thus, the frequency of the radiation emitted is related only to change in atomic energy levels and it does not depend on frequency of orbital motion of the electron.

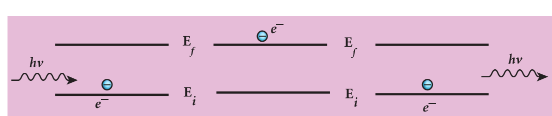


Figure 9.16 Absorption and emission of radiation

EXAMPLE 9.1

The radius of the 5th orbit of hydrogen atom is 13.25 Å. Calculate the de broglie wavelength of the electron orbitting in the 5th orbit.

Solution:

$2\pi r = n\lambda$

$2 \times 3.14 \times 13.25\text{Å} = 5 \times \lambda$

$\therefore\ \lambda = 16.64\text{Å}$

EXAMPLE 9.2

Find the (i) angular momentum (ii) velocity of the electron revolving in the 5th orbit of hydrogen atom.

($h = 6.6 \times 10^{-34}$ Js, $m = 9.1 \times 10^{-31}$ kg)

Solution

(i) Angular momentum is given by

$$l = n\hbar = \frac{nh}{2\pi}$$

$$= \frac{5\times6.6\times10^{-34}}{2\times3.14} = 5.25\times10^{-34}\ \text{kgm}^2\text{s}^{-1}$$

(ii) Velocity is given by

$$\text{Velocity } \upsilon = \frac{l}{mr}$$

$$= \frac{(5.25\times10^{-34}\ \text{kgm}^2\text{s}^{-1})}{(9.1\times10^{-31}\ \text{kg})(13.25\times10^{-10}\ \text{m})}$$

$$\upsilon = 4.4\times10^{5}\ \text{ms}^{-1}$$

Radius of the orbit of the electron and velocity of the electron

Consider an atom which contains the nucleus at rest and an electron revolving around the nucleus in a circular orbit of radius r_n as shown in Figure 9.17. Nucleus is made up of protons and neutrons. Since proton is positively charged and neutron is electrically neutral, the charge of a nucleus is entirely due to the charge of protons.

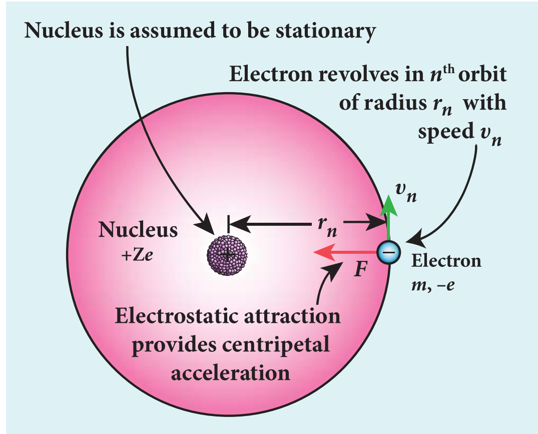


Figure 9.17 Electron revolving around the nucleus

Let Z be the atomic number of the atom, then $+Ze$ is the charge of the nucleus. Let $-e$ be the charge of the electron. From Coulomb's law, the force of attraction between the nucleus and the electron is

$$\vec{F}_{coloumb} = \frac{1}{4\pi\varepsilon_0}\frac{(+Ze)(-e)}{r_n^2}\hat{r}$$

$$= -\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r_n^2}\hat{r}$$

This force provides necessary centripetal force

$$\vec{F}_{centripetal} = \frac{m\upsilon_n^2}{r_n}\hat{r}$$

where m be the mass of the electron that moves with a velocity υ_n in a circular orbit. Therefore,

$$\left|\vec{F}_{coloumb}\right| = \left|\vec{F}_{centripetal}\right|$$

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

Multiplied and divided by 'm'

$$r_n = \frac{4\pi\varepsilon_0(m\upsilon_n r_n)^2}{Zme^2} \quad (9.15)$$

From Bohr's assumption, the angular momentum quantization condition, $m\upsilon_n r_n = l_n = n\hbar$,

$$\therefore r_n = \frac{4\pi\varepsilon_0 (m\upsilon_n r_n)^2}{Zme^2}$$

$$r_n = \frac{4\pi\varepsilon_0 (n\hbar)^2}{Zme^2} = \frac{4\pi\varepsilon_0 n^2\hbar^2}{Zme^2}$$

$$r_n = \left(\frac{\varepsilon_0 h^2}{\pi me^2}\right)\frac{n^2}{Z} \qquad (\because \hbar = \frac{h}{2\pi}) \quad (9.16)$$

where $n \in \mathbb{N}$. Since, ε_0, h, e and π are constants. Therefore, the radius of the orbit becomes

$$r_n = a_0 \frac{n^2}{Z}$$

where $a_0 = \frac{\varepsilon_0 h^2}{\pi me^2} = 0.529$ Å. This is known as Bohr radius which is the smallest radius of the orbit in hydrogen atom. Bohr radius is also used as unit of length called Bohr. 1 Bohr = 0.53 Å. For hydrogen atom (Z = 1), the radius of n^{th} orbit is

$$r_n = a_0 n^2$$

For $n = 1$ (first orbit or ground state),

$$r_1 = a_0 = 0.529 \text{ Å}$$

For n = 2 (second orbit or first excited state),

$$r_2 = 4a_0 = 2.116 \text{ Å}$$

For n = 3 (third orbit or second excited state),

$$r_3 = 9a_0 = 4.761 \text{ Å}$$

and so on.

Thus the radius of the orbit from centre increases with n, that is, $r_n \propto n^2$ as shown in Figure 9.18.

Further, Bohr's angular momentum quantization condition leads to

$$\frac{m\upsilon_n a_0 n^2}{Z} = n\frac{h}{2\pi} \quad \left[\therefore r_n = a_0\frac{n^2}{Z}\right]$$

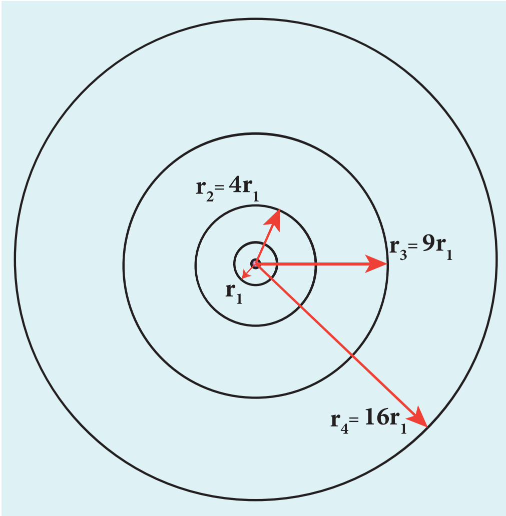


Figure 9.18 Variation of radius of the orbit with principal quantum number

$$\upsilon_n = \frac{h}{2\pi m a_0}\frac{Z}{n}$$

in atomic physics $\upsilon_n \propto \frac{1}{n}$

Note that the velocity of electron decreases as the principal quantum number (orbit number) increases as shown in Figure 9.19. This curve is the rectangular hyperbola. This implies that the velocity of electron in ground state is maximum when compared to that in excited states.

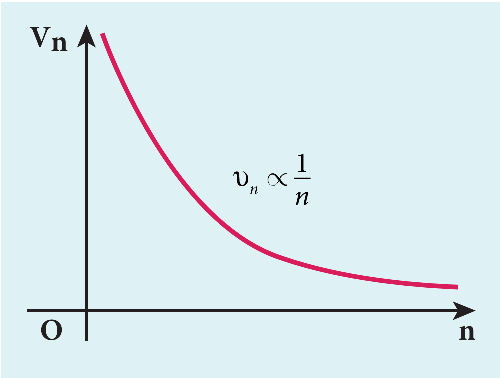


Figure 9.19 Variation of velocity of the electron in the orbit with principal quantum number

The energy of an electron in the n^{th} orbit

Since the electrostatic force is a conservative force, the potential energy for the n^{th} orbit is

$$U_n = \frac{1}{4\pi\varepsilon_0}\frac{(+Ze)(-e)}{r_n} = -\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r_n}$$

$$= -\frac{1}{4\varepsilon_0^2}\frac{Z^2me^4}{h^2n^2} \quad \left(\because r_n = \frac{\varepsilon_0 h^2}{\pi me^2}\frac{n^2}{Z}\right)$$

The kinetic energy of the electron in n^{th} orbit is

$$KE_n = \frac{1}{2}m\upsilon_n^2 = \frac{me^4}{8\varepsilon_0^2h^2}\frac{Z^2}{n^2}$$

This implies that $U_n = -2\ KE_n$. Total energy of the electron in in the n^{th} orbit is

$$E_n = KE_n + U_n = KE_n - 2KE_n = -KE_n$$

$$E_n = -\frac{me^4}{8\varepsilon_0^2h^2}\frac{Z^2}{n^2}$$

For hydrogen atom (Z = 1),

$$E_n = -\frac{me^4}{8\varepsilon_0^2h^2}\frac{1}{n^2}\ \text{joule} \qquad (9.17)$$

where n stands for principal quantum number. The negative sign in equation (9.17) indicates that the electron is bound to the nucleus.

Substituting the values of mass and charge of an electron (m and e), permittivity of free space ε_0 and Planck's constant h and expressing energy in terms of electron(+(eV)), we get

$$E_n = -13.6\frac{1}{n^2}\,eV$$

For the first orbit (ground state), the total energy of electron is $E_1 = -13.6\ eV$.
For the second orbit (first excited state), the total energy of electron is $E_2 = -3.4\ eV$.
For the third orbit (second excited state), the total energy of electron is $E_3 = -1.51\ eV$ and so on.

Notice that the energy of the first excited state is greater than that of the ground state, second excited state is greater than that of the first excited state and so on. Thus, the orbit which is closest to the nucleus (r_1) has lowest energy (minimum energy what it is compared with other orbits). So, it is often called ground state energy (lowest energy state). The ground state energy of hydrogen (–13.6 eV) is used as a unit of energy called Rydberg (1 Rydberg = –13.6 eV).

The negative value of this energy is because of the way the zero of the potential energy is defined. When the electron is taken away to an infinite distance (very far distance) from nucleus, both the potential energy and kinetic energy terms vanish and hence the total energy also vanishes.

The energy level diagram along with the shape of the orbits for increasing values of n are shown in Figure 9.20. It shows that the energies of the excited states come closer and closer together when the principal quantum number n takes higher values.

EXAMPLE 9.3

(a) Show that the ratio of velocity of an electron in the first Bohr orbit to the speed of light c is a dimensionless number.

(b) Compute the velocity of electrons in ground state, first excited state and second excited state in Bohr atom model for hydrogen atom.

Solution

(a) The velocity of an electron in n^{th} orbit is

$$\upsilon_n = \frac{h}{2\pi m a_0}\frac{Z}{n}$$

where $a_0 = \frac{\varepsilon_0 h^2}{\pi me^2}$ = Bohr radius. Substituting for a_0 in υ_n,

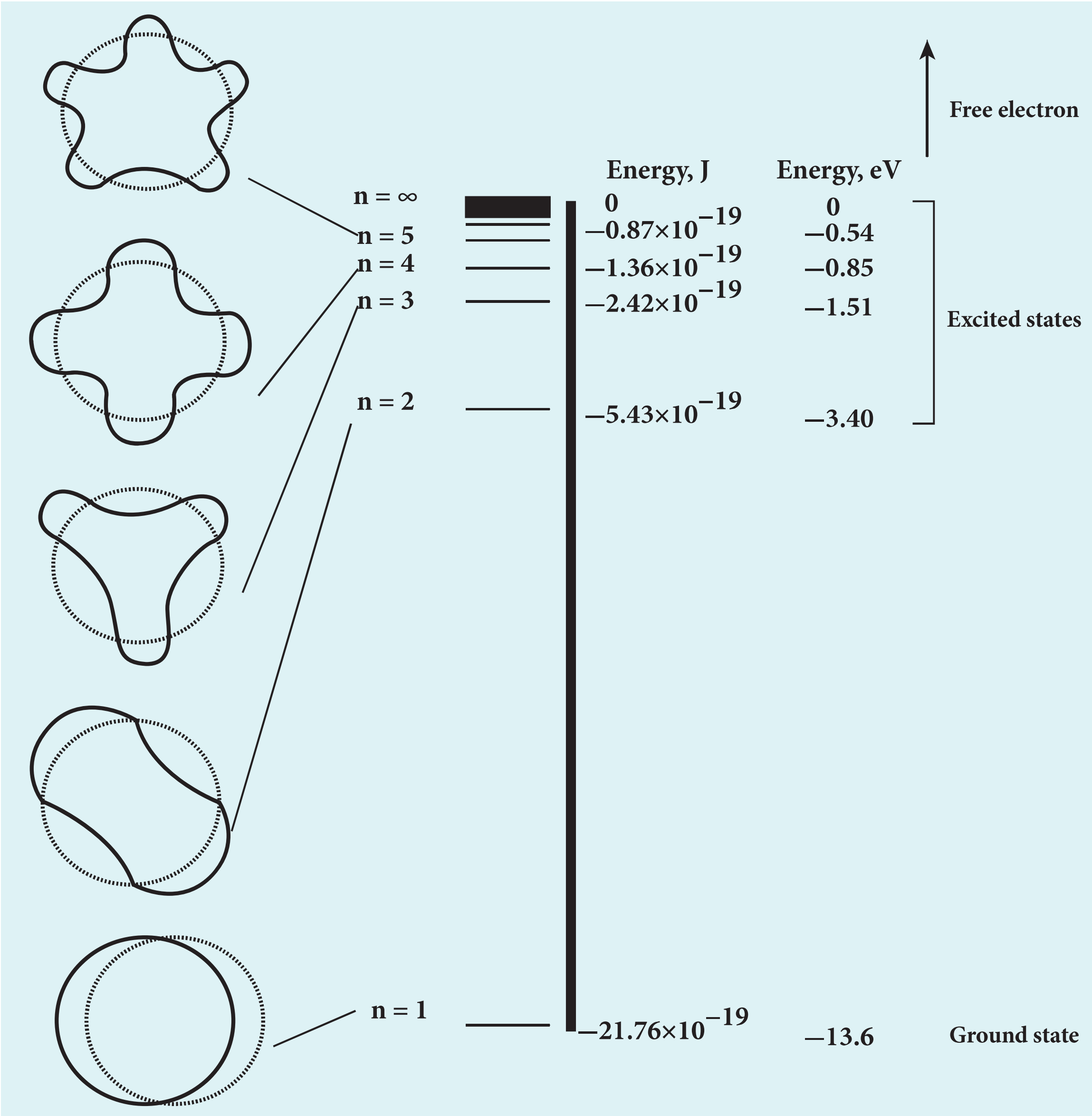


Figure 9.20 **Energy levels of a hydrogen atom**

$$\upsilon_n = \frac{e^2}{2\varepsilon_0 h}\frac{Z}{n} = c\left(\frac{e^2}{2\varepsilon_0 hc}\right)\frac{Z}{n} = \frac{\alpha c Z}{n}$$

where c is the speed of light in free space or vacuum and its value is $c = 3 \times 10^8$ m s^{-1}and α is called fine structure constant.

For a hydrogen atom, $Z = 1$ and for the first orbit, $n = 1$, the ratio of velocity of electron in first orbit to the speed of light in vacuum or free space is

$$\alpha = \frac{(1.6\times10^{-19}\,\mathrm{C})^2}{2\times(8.854\times10^{-12}\,\mathrm{C^2N^{-1}m^{-2}})} \times \frac{1}{(6.6\times10^{-34}\,\mathrm{Nms})\times(3\times10^8\,\mathrm{ms^{-1}})}$$

$$\frac{\upsilon_1}{c} = \alpha = \frac{e^2}{2\varepsilon_0 hc}$$

$\approx \frac{1}{136.9} = \frac{1}{137}$ which is a dimensionless number

$$\Rightarrow \alpha = \frac{1}{137}$$

(b) Using fine structure constant, the velocity of electron can be written as

$$\upsilon_n = \frac{\alpha c Z}{n}$$

For hydrogen atom ($Z = 1$) the velocity of electron in n^{th} orbit is

$$\upsilon_n = \frac{c}{137}\frac{1}{n} = (2.19 \times 10^6)\frac{1}{n}\ \text{ms}^{-1}$$

For the first orbit (ground state), the velocity of electron is

$$\upsilon_1 = 2.19 \times 10^6\,\text{ms}^{-1}$$

For the second orbit (first excited state), the velocity of electron is

$$\upsilon_2 = 1.095 \times 10^6\,\text{ms}^{-1}$$

For the third orbit (second excited state), the velocity of electron is

$$\upsilon_3 = 0.73 \times 10^6\,\text{ms}^{-1}$$

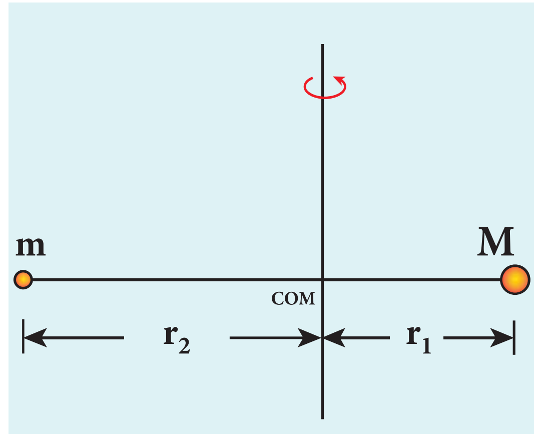
Here, $v_1 > v_2 > v_3$

EXAMPLE 9.4

The Bohr atom model is derived with the assumption that the nucleus of the atom is stationary and only electrons revolve around the nucleus. Suppose the nucleus is also in motion, then calculate the energy of this new system.

Solution

Let the mass of the electron be m and mass of the nucleus be M. Since there is no external force acting on the system, the centre of mass of hydrogen atom remains at rest. Hence, both nucleus and electron move about the centre of mass as shown in figure.



Let V be the velocity of the nuclear motion and v be the velocity of electron motion. Since the total linear momentum of the system is zero,

$$-m\upsilon + M\upsilon = 0 \text{ or}$$

$$MV = m\upsilon = p$$

$$\vec{p}_e + \vec{p}_n = \vec{0} \text{ or}$$

$$|\vec{p}_e| = |\vec{p}_n| = p$$

Hence, the kinetic energy of the system is

$$KE = \frac{p_n^2}{2M} + \frac{p_e^2}{2m} = \frac{p^2}{2}\left(\frac{1}{M} + \frac{1}{m}\right)$$

Let $\frac{1}{M} + \frac{1}{m} = \frac{1}{\mu_m}$. Here the reduced mass is, $\mu_m = \frac{mM}{M+m}$

Therefore, the kinetic energy of the system now is $KE = \frac{p^2}{2\mu_m}$

Since the potential energy of the system is same, the total energy of the hydrogen can be expressed by replacing mass by reduced mass, which is

$$E_n = -\frac{\mu_m e^4}{8\varepsilon_0^2 h^2}\frac{1}{n^2}$$

Since the nucleus is very heavy compared to the electron, the reduced mass is closer to the mass of the electron.

Note

In 1931, H.C. Urey and co-workers noticed that in the shorter wavelength region of the hydrogen spectrum lines, faint companion lines are observed. From the isotope displacement effect (isotope shift), the isotope of the same element can produce slightly different spectral lines. The presence of these faint lines confirmed the existence of isotopes of hydrogen atom (which is named as Deuterium).

On calculating wavelength or wave number difference between the faint and bright spectral lines, atomic mass of deuterium is measured to be twice that of atomic mass of hydrogen atom. Bohr atom model could not explain this isotopic shift. Thus by considering nuclear motion (although the movement of the nucleus is much smaller) into account in the Bohr atom model, the wave number or wavelength difference between the lines produces by the hydrogen atom and deuterium is theoretically calculated which perfectly agreed with the spectroscopic measured values.

The difference between hydrogen atom and deuterium is in the number of neutron. Hydrogen atom contains an electron and a proton, whereas deuterium has an electron, a proton and a neutron.

Excitation energy and excitation potential

The energy required to excite an electron from lower energy state to any higher energy state is known as excitation energy.

The excitation energy for an electron from ground state (n = 1) to first excited state (n = 2) is called first excitation energy.

For hydrogen atom, it is

$$E_I = E_2 - E_1 = -3.4\ eV - (-13.6\ eV) = 10.2\ eV$$

Similarly, the excitation energy for an electron from ground state (n = 1) to second excited state (n = 3) is called second excitation energy, which is

$$E_{II} = E_3 - E_1 = -1.51\ eV - (-13.6\ eV) = 12.1\ eV$$

and so on.

Excitation potential is defined as excitation energy per unit charge.

First excitation potential for hydrogen atom is,

For hydrogen atom, the ground state ionization energy is

$$E_I = eV_I \Rightarrow V_I = \frac{1}{e}E_I = 10.2\ \text{volt}$$

Second excitation potential is,

$$E_{II} = eV_{II} \Rightarrow V_{II} = \frac{1}{e}E_{II} = 12.1\ \text{volt}$$

and so on.

Ionization energy and ionization potential

An atom is said to be ionized when an electron is completely removed from the atom – that is, it reaches the state with energy $E_{n\to\infty}$. **The minimum energy required to remove an electron from an atom in the ground state is known as binding energy or ionization energy**.

For hydrogen atom, the ground state ionization energy is,

$$E_{\text{ionizaation}} = E_\infty - E_1 = 0 - (-13.6 eV)$$
$$= 13.6 eV$$

When an electron is in n^{th} state of an atom, the energy required to remove an electron from that state – that is, the corresponding ionization energy is

$$E_{\text{ionization}} = E_\infty - E_n = 0 - \left(-\frac{13.6}{n^2}Z^2\ \text{eV}\right)$$
$$= \frac{13.6}{n^2}Z^2\ \text{eV}$$

At normal room temperature, the electron in a hydrogen atom (Z=1) spends most of

Table 9.1

Physical quantity	Ground state	First excited state	Second excited state
Radius ($r_n \propto n^2$)	0.529 Å	2.116 Å	4.761 Å
Velocity ($v_n \propto n^{-1}$)	2.19×10^6 m s^{-1}	1.095×10^6 m s^{-1}	0.73×10^6 m s^{-1}
Total Energy ($E_n \propto n^{-2}$)	-13.6 *e*V	-3.4 *e*V	-1.51 *e*V

its time in the ground state. **The amount of energy required to remove an electron from the ground state of an atom *(E = 0* for *n→∞)* is known as first ionization energy (13.6 *e*V).** Then, the hydrogen atom is said to be in ionized state or simply called as hydrogen ion, denoted by H^+. If we supply more energy than the ionization energy, the excess energy appear as the kinetic energy of the free electron.

Ionization potential is defined as ionization energy per unit charge.

$$V_{\text{ionization}} = \frac{1}{e} E_{\text{ionization}} = \frac{13.6}{n^2} Z^2 \, V$$

Thus, for a hydrogen atom (Z=1), the ionization potential is

$$V = \frac{13.6}{n^2} \text{ volt}$$

The radius, velocity and total energy in ground state, first excited state and second excited state are given in Table 9.1.

EXAMPLE 9.5

Suppose the energy of an electron in hydrogen–like atom is given as $E_n = -\frac{54.4}{n^2}$ eV where $n \in \mathbb{N}$. Calculate the following:

(a) Sketch the energy levels for this atom and compute its atomic number.

(b) If the atom is in ground state, compute its first excitation potential and also its ionization potential.

(c) When a photon with energy 42 *e*V and another photon with energy 56 *e*V are made to collide with this atom, does this atom absorb these photons?

(d) Determine the radius of its first Bohr orbit.

(e) Calculate the kinetic and potential energies of electron in the ground state.

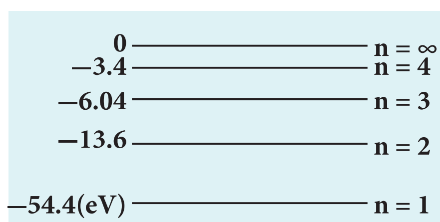
Solutions

(a) Given that

$$E_n = -\frac{54.4}{n^2} \, e\text{V}$$

For $n = 1$, the ground state energy $E_1 = -54.4$ *e*V and for $n = 2$, $E_2 = -13.6$ *e*V. Similarly, $E_3 = -6.04$ *e*V, $E_4 = -3.4$ *e*V and so on.

For large value of principal quantum number – that is, $n = \infty$, we get $E_\infty = 0$ *e*V.



(b) For a hydrogen-like atom, ground state energy is

$$E_1 = -\frac{13.6}{n^2} Z^2 \, e\text{V}$$

where Z is the atomic number. Hence, comparing this energy with given energy, we get, $-13.6 Z^2 = -54.4 \Rightarrow Z = \pm 2$. Since, atomic number cannot be negative number, $Z = 2$.

The first excitation energy is

$$E_I = E_2 - E_1 = -13.6\,eV - (-54.4\,eV)$$
$$= 40.8\,eV$$

Hence, the first excitation potential is

$$V_I = \frac{1}{e}E_I = \frac{(40.8\,eV)}{e}$$
$$= 40.8\,\text{volt}$$

The first ionization energy is

$$E_{\text{ionization}} = E_\infty - E_1 = 0 - (-54.4\,eV)$$
$$= 54.4\,eV$$

Hence, the first ionization potential is

$$V_{\text{ionization}} = \frac{1}{e}E_{\text{ionization}} = \frac{(54.4\,eV)}{e}$$
$$= 54.4\,\text{volt}$$

(c) Consider two photons to be A and B. Given that photon A with energy 42 *eV* and photon B with energy 51 *eV*

From Bohr assumption, difference in energy levels is equal to the energy photon absored, then atom will absorb energy, otherwise, not.

$$E_2 - E_1 = -13.6\,eV - (-54.4\,eV)$$
$$= 40.8\,eV \approx 41\,eV$$

Similarly,

$$E_3 - E_1 = -6.04\,eV - (-54.4\,eV)$$
$$= 48.36\,eV$$

$$E_4 - E_1 = -3.4\,eV - (-54.4\,eV)$$
$$= 51\,eV$$

$$E_3 - E_2 = -6.04\,eV - (-13.6\,eV)$$
$$= 7.56\,eV$$

and so on.

But note that $E_2 - E_1 \neq 42$ eV, $E_3 - E_1 \neq 42$ eV, $E_4 - E_1 \neq 42$ eV and $E_3 - E_2 \neq 42$ eV.

For all possibilities, no difference in energy is an integer multiple of photon energy. Hence, photon A is not absorbed by this atom. But for Photon B, $E_4 - E_1 = 51$ eV, which means, Photon B can be absorbed by this atom.

(d) The radius of Bohr orbit is $r_n = \frac{a_0 \times n^2}{z}$

For $n = 1$, $z = 2$

$$r_1 = \frac{a_0}{2}$$
$$= \frac{0.529}{2}$$
$$= 0.265\ \text{Å}$$

(e) Since, total energy is equal to negative of kinetic energy in Bohr atom model, we get

$$KE_n = -E_n = -\left(-\frac{54.4}{n^2}eV\right)$$
$$= \frac{54.4}{n^2}eV$$

Since, Potential energy is negative of twice the kinetic energy,

$$U_n = -2KE_n = -2\left(\frac{54.4}{n^2}eV\right)$$
$$= -\frac{108.8}{n^2}eV$$

For a ground state, put n =1

Kinetic energy is KE_1 = 54.4 eV and Potential energy is U_1 = –108.8 eV

9.3.4 Atomic spectra

Materials in the solid, liquid and gaseous states emit electromagnetic radiations when they are heated up and these emitted radiations usually exhibit continuous spectrum. For example, when white light is examined through a spectrometer, electromagnetic radiations of all wavelengths are observed which is a continuous spectrum.

In early twentieth century, many scientists spent considerable time in understanding the characteristic radiations emitted by the atoms of individual elements exposed to a flame or

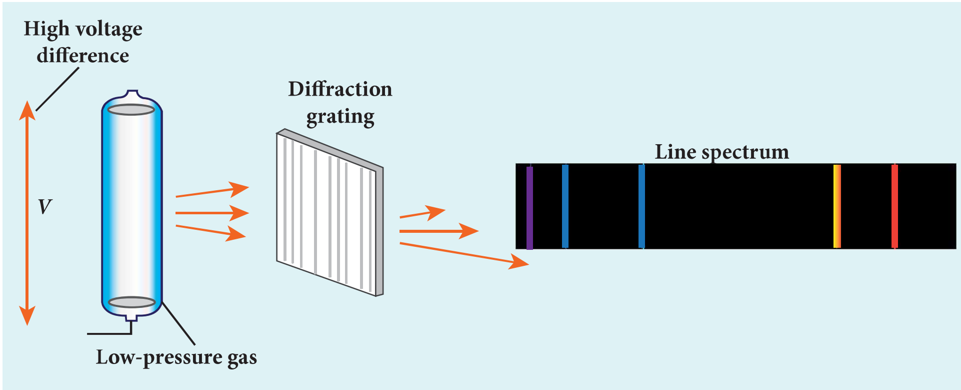


Figure 9.21 Spectrum of an atom

electrical discharge. When they were viewed or photographed, instead of a continuous spectrum, the radiation contains of a set of discrete lines, each with characteristic wavelength. In other words, the wavelengths of the radiation obtained are well defined and their positions and intensities are characteristic of the element as shown in Figure 9.21.

This implies that these spectra are unique to each element and can be used to identify the element of the gas (like finger print used to identify a person) – that is, it varies from one gas to another gas. This uniqueness of line spectra of elements made the scientists to determine the composition of stars, sun and also used to identify the unknown compounds.

Hydrogen spectrum

When the hydrogen gas enclosed in a tube is heated up, it emits electromagnetic radiations of certain sharply-defined characteristic wavelength (line spectrum), called hydrogen emission spectrum (Refer unit 5, volume 1 of +2 physics text book). The emission spectrum of hydrogen is shown in Figure 9.22(a).

When any gas is heated up, the thermal energy is supplied to excite the electrons. Similarly by all occurring light on the atoms, electrons can be excited. Once the electrons get sufficient energy as given by Bohr's postulate (c), it absorbs energy with particular wavelength (or frequency) and jumps from one stationary state (original state) to another state with those wavelengths (or frequencies) for the colours that are not observed are seen as dark lines in the absorption spectrum as shown in Figure 9.22 (b).

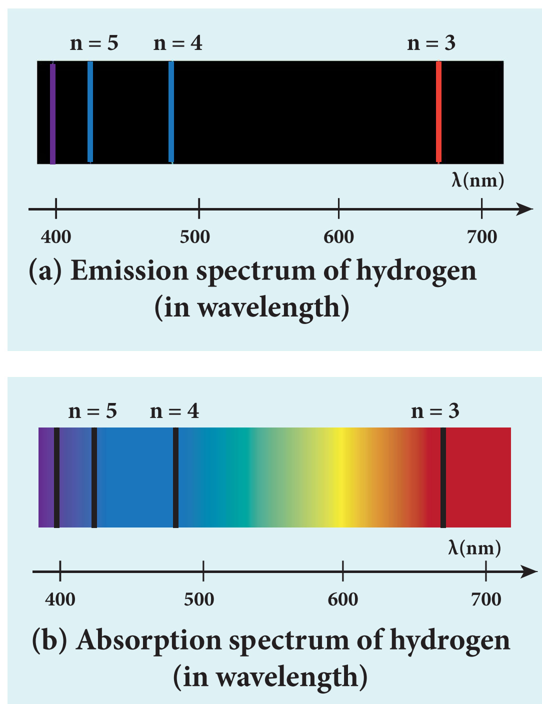


Figure 9.22 Hydrogen spectrum (a) emission (b) absorption

Since electrons in excited states have very small life time, these electrons jump back to ground state through spontaneous emission in a short duration of time (approximately 10^{-8} s) by emitting the radiation with same wavelength (or frequency) corresponding to the colours it absorbed (Figure 9.22 (a)). This is called emission spectroscopy.

The wavelengths of these lines can be calculated with great precision. Further, the emitted radiation contains wavelengths both lesser and greater than wavelengths of lines in the visible spectrum.

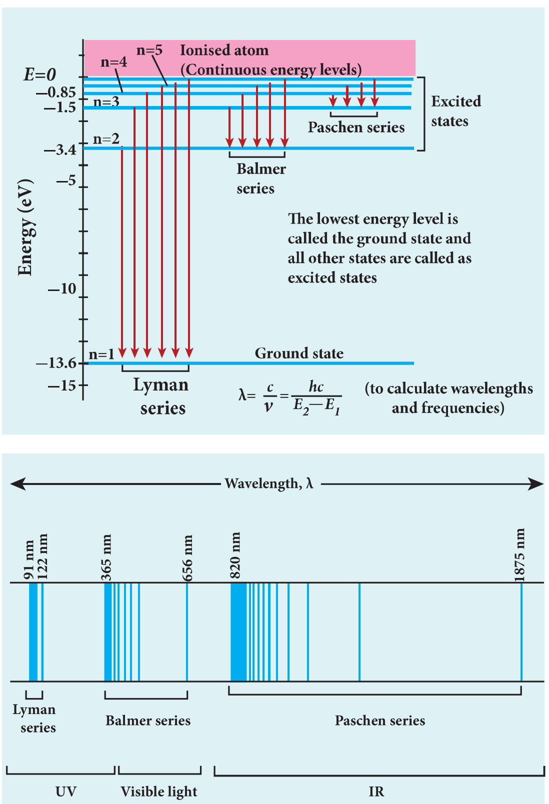


Figure 9.23 Spectral series – Lyman, Balmer, Paschen series

Notice that the spectral lines of hydrogen as shown in Figure 9.23 are grouped in separate series. In each series, the distance of separation between the consecutive wavelengths decreases from higher wavelength to the lower wavelength, and also wavelength in each series approach a limiting value known as the series limit. These series are named as Lyman series, Balmer series, Paschen series, Brackett series, Pfund series, etc. The wavelengths of these spectral lines perfectly agree with the wavelengths calculate using equation derived from Bohr atom model.

$$\frac{1}{\lambda} = R\left(\frac{1}{n^2} - \frac{1}{m^2}\right) = \bar{\nu} \qquad (9.18)$$

where $\bar{\nu}$ **is known as wave number which is inverse of wavelength,** R is known as Rydberg constant whose value is $1.09737 \times 10^7\ \text{m}^{-1}$ and m and n are positive integers such that $m > n$. The various spectral series are discussed below:

(a) Lyman series

For $n = 1$ and $m = 2,3,4.......$ in equation (9.18), the wave numbers or wavelength of spectral lines of Lyman series which lies in ultra-violet region,

$$\bar{\nu} = \frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{m^2}\right)$$

(b) Balmer series

For $n = 2$ and $m = 3,4,5.......$ in equation (9.18), the wave numbers or wavelength of spectral lines of Balmer series which lies in visible region,

$$\bar{\nu} = \frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{m^2}\right)$$

(c) Paschen series

Put $n = 3$ and $m = 4,5,6.......$ in equation (9.18). The wave number or wavelength of spectral lines of Paschen series which lies in infra-red region (near IR) is

$$\bar{\nu} = \frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{m^2}\right)$$

(d) Brackett series

For $n = 4$ and $m = 5,6,7........$ in equation (9.18), the wave numbers or wavelength of spectral lines of Brackett series which lies in infra-red region (middle IR),

$$\bar{v} = \frac{1}{\lambda} = R\left(\frac{1}{4^2} - \frac{1}{m^2}\right)$$

(e) Pfund series

For $n = 5$ and $m = 6,7,8........$ in equation (9.18), the wave numbers or wavelength of spectral lines of Pfund series which lies in infra-red region (far IR),

$$\bar{v} = \frac{1}{\lambda} = R\left(\frac{1}{5^2} - \frac{1}{m^2}\right)$$

Different spectral series are listed in Table 9.2.

Table 9.2

n	*m*	Series Name	Region
1	2,3,4.....	Lyman	Ultraviolet
2	3,4,5.....	Balmer	Visible
3	4,5,6.....	Paschen	Infrared
4	5,6,7......	Brackett	Infrared
5	6,7,8.....	Pfund	Infrared

Limitations of Bohr atom model

The following are the drawbacks of Bohr atom model

(a) Bohr atom model is valid only for hydrogen atom or hydrogen like-atoms but not for complex atoms.

(b) When the spectral lines are closely examined, individual lines of hydrogen spectrum are accompanied by a number of faint lines. This is called **fine structure**. This cannot be explained by Bohr atom model.

(c) Bohr atom model fails to explain the intensity variations in the spectral lines.

(d) The distribution of electrons in various levels cannot be completely explained by Bohr atom model.

9.4 NUCLEI

Introduction

In the previous section, we have discussed about various preliminary atom models, Rutherford's alpha particle scattering experiment and Bohr atom model. These played a vital role to understand the structure of the atom and the nucleus. In this section, the structure of the nuclei and their properties, classifications are discussed.

9.4.1 Composition of nucleus

Atoms have a nucleus surrounded by electrons. The nucleus contains protons and neutrons. The neutrons are electrically neutral ($q = 0$) and the protons have positive charge ($q = +e$) equal in magnitude to the charge of the electron ($q = -e$). **The number of protons in the nucleus is called the atomic number** and it is denoted by *Z*. The number of neutrons in the nucleus is called neutron number (*N*). **The total number of neutrons and protons in the nucleus is called the mass number** and it is denoted by *A*. Hence, $A = Z+N$.

The two constituents of nucleus namely neutrons and protons, are collectively called as nucleons. The mass of a proton is 1.6726×10^{-27} kg which is roughly 1836 times the mass of the electron. The mass of a neutron is slightly greater than the mass of the proton and it is equal to 1.6749×10^{-27} kg.

To specify the nucleus of any element, we use the following general notation

$$_Z^AX$$

where X is the chemical symbol of the element, A is the mass number and Z is the atomic number. For example, the