29

STRUCTURE OF ATOM

29.1 INTRODUCTION

In our earlier lessons we have studied about different properties of matter like mechanical, thermal, electrical and magnetic properties. In order to study the microlevel details of such properties of the matter, it becomes imperative to know the inner structure of the atom which is primarily responsible for such properties and other effects.

Therefore to understand the structure of atoms, we have taken the different atomic models for discussion in this lesson. Starting with the Rutherford's scattering experiment and its shortcomings, this lesson includes Bohr's theory of atomic model to explain the first electronic structure of the smallest and lightest atom i.e. hydrogen .

In this lesson you will also study about the energy of the electrons in different orbits specially in the case of hydrogen atom. A brief introduction of the quantum numbers and Pauli's exclusion principle will also be discussed in this lesson.

29.2 OBJECTIVES

After studying this lesson, you should be able to :

- *explain Rutherford's scattering experiment and state the Ruther ford's atomic model;*
- *state the short comings of Rutherford's model of an atom;*
- *state the postulates of Bohr's model of the atom in mathematical form;*
- *compute the radius and velocity of an electron in the first orbit of hydrogen atom;*
- *derive the expression for the energy of the electron in the first orbit of a hydrogen atom;*
- *represent the energy of the electrons in an energy level diagram;*
- - *list the four quantum numbers and Pauli's exclusion principle;*

29.3 RUTHERFORD'S ATOMIC MODEL

As you might have studied earlier, the first man who talked about the atom was John Dalton, an English chemist. He described atom as the smallest indivisible particle of elements which have all the properties of element and takes part in all chemical reactions. Later in 1907, J.J Thomson while studying discharge of electricity through gases discovered that all matter contains negatively charged particle called electron. Because the atom as a whole is neutral, it must carry equal and opposite charges. On the basis of this he suggested ***plum pudding model*** of atom *according to which the atom is spherical in shape and it consisted of a uniformly charged positive ball in which a suitable number of electrons are embedded at different places to make it neutral.* By this time the phenomena of radioactivity was discovered by Henry Becquerel. It was found that the radio active atom emitted energetic radiations consisting of α -particle, β -particles and γ-radiations. While performing an experiment with α-particles in 1899, Lord Ernest Rutherford, a New Zealand born Physicist observed large angle scattering of α particles from the foils of Gold, Which could not be explained by the plum-pudding model of the atom. Further it led him to the discovery of the nucleus also.

The first correct description of the distribution of positive and negative charges within the atom was presented by Lord Ernest Rutherford (1871- 1937). His model was based on experiments of the scattering of α-particles (alpha particles) from thin metal foils. Let us study about Rutherford's Scattering Experiment

29.3.1 Rutherford's Scattering Experiment

The apparatus (experimental arrangement) used by Rutherford for carrying out scattering experiment was very simple and is shown in Fig. 29.1. A lead diaphragm P allows a thin pencil of α–particles emitted by the polonium source at S to pass througt it. This beam falls on target T, which is an extremely thin film of gold . The fluorescent screen (zinc sulphide) produced O tiny flashes of light or ***scintillation*** when α -particles strike it. Scintillation can be observed through a low power microscope M. The whole apparatus was placed in an evacuated chamber to avoid collisions of α– particles with the air molecules.

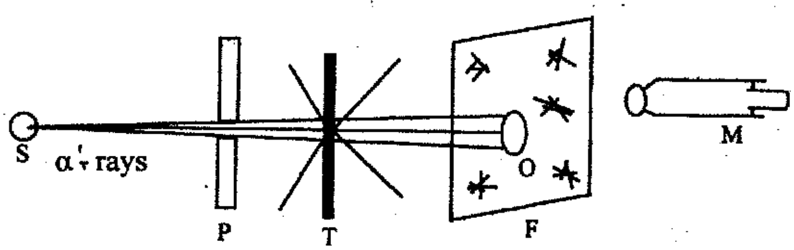

Fig.29.1: *Rutherford' Scattering of α – Particles*

Rutherford was curious to know the angle of deflection suffered by the α–particles on passing through the foil. He believed that this kind of study might lead to some important information regarding the structure of atom. He observed that most of the scattered α–particles passed straight through the target foil T or suffered only small deflections as expected. A few scattered α–particles were, however, observed to have been deflected through large angles (of 90°) and some of them were even deflected backwards, towards the source (making an angle of almost 180°). Scattering of α -particles by a small angle was expected but this large angle scattering, particularly back along the direction from which the particles came, was very surprising.

Explanations & Conclusion : The large angle scattering of α–particles indicated the presence of some hard core inside the atoms. In order to produce large angle scattering, a sufficiently strong electrostatic repulsion between the positive charge of the bombarding particles and the positive charge of bombarded target atoms was necessary. Such a strong field could only be provided by all the positive charge concentrated at a point. Thus, he concluded that,

- ***All the positive charges and most of the mass of the atom must be concentrated in a very small central region inside the atom. This region was named by Rutherford as nucleus.***

  He, also estimated that the size of the nucleus was of the order of 10^{-15}m i.e. 10^{-5} times the size of the atom.
- ***The nucleus is surrounded by a cloud of electrons whose total negative charge is equal to the total positive charge of the nucleus so that the atom as a whole is electrically neutral.***

Possible paths of α particles scattered by an atom are shown in Fig.29.2

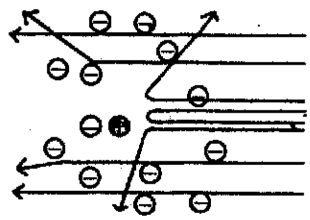

Fig 29.2: *Rutherford's experiment of scattering of α -particles.*

An approaching α particle (carrying positive charge) faces a Coulombian repulsive force due to the positive charge on the nucleus. An α-particle passing through the atom at a large distance from the nucleus experiences small repulsive force so it passes almost undeflected as indicated by α-particle 1 in Fig. 29.2. But another particle passing close to the nucleus experiences a strong repulsive force and so it deflects through a large angle as indicated by α-particle 2 in the same figure. It is very rare that an α-particle aims just at the nucleus. In this rare case the repulsive force will slow down the α-particle which will finally stop and then will be repelled back along the direction of its approach as shown by the α -particle 4 in the figure.

29.2.2 Drawbacks of Rutherford's Model

The atomic model proposed by Rutherford offered some difficulties. Some of the consequences of Rutherford's atomic model was in contradiction with the experimental observation as discussed here.

(i) Stability of the Atom: According to Rutherford's model, electrons in the atom revolve around the nucleus in various orbits (like the planets of our solar system revolving around the sun). If the electrons were stationary they would at once be attracted towards the nucleus and fall into it. It was thought that it was the centrifugal force which prevented the electrons from falling into the nucleus.

This model was also found to be unsatisfactory. In accordance with electron-magnetic theory, all accelerated electrons emit electromagnetic radiations continuously. This is because electrons revolving in orbits around the nucleus are continuously accelerated, they will emit radiation and lose energy.

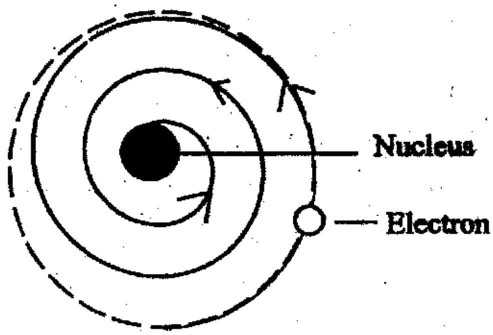

Fig. 29.3 Orbital motion of the electron.

Thus, an electron in the Rutherford's model must orbit along a spiral path of decreasing radius till it collapses into the nucleus as shown in Fig. 29.3.

From Fig 29.3 we see that the orbital motion of the planetary electrons is short lived (about 10^{-10} s) and is not stable. This is in contradiction with experimental results.

(ii) Frequency of Electromagnetic Radiation: According to Rutherford's model electrons can rotate in any orbit of any radius and thus must emit electromagnetic radiations of all frequencies. This is contradictory to the experimental results. Experimentally, it was found that atoms emit radiations of certain well defined frequencies only and not of all frequencies.

Thus, this model faced some difficulties which could not be explained and thus it requirec an improvement.

Now it is time to assess your progress and therefore try to solve the following questions.

INTEXT QUESTIONS 29.1

1. Choose the correct answers:

(a) *In Rutherford's scattering experiment, the thin film target was bombarded with.*
(i) β particle (11) γ rays (iii) α particle

(b) *The nucleus is surrounded by a cloud of*
(i) Electrons (ii) Protons (iii) α particles.

(c) *The large angle scattering of particles indicated the presence of*
(i) Some positively charged hard core inside the atoms (ii) some porous core inside the atoms (iii) Negatively charged core.

2. What was the draw back of Thomson's plum pudding model of atom?

29.4 BOHR'S MODEL OF THE ATOM

A satisfactory model of an atom was suggested by Neils Bohr in the year 1913. Rutherford's experiment on the scattering of α-particles led Bohr to the conclusion that the atom consists of a positively charged nucleus at its center with electrons revolving around it in circular orbits. Moreover, Bohr applied the quantum theory of radiation as developed by Plank and Einstein to the Rutherford's model to overcome the difficulties faced by this model. Bohr dealt mainly with the simplest of atoms i.e. the hydrogen atom with atomic number $Z = 1$

29.4.1 Bohr's Theory of Hydrogen Atom

It is also referred to as the planetary model of the atom. It consists of a single electron revolving around a nucleus, which is positively charged . It is called the nucleus which consist of a positively charged particls called proton and. is about 2000 times heavier than an electron. The charge of a proton is exactly equal to that of the electrons. Bohr's theory is based on the following postulates :

Postulates of Bohr's Model of the Atom

Bohr formulated the following postulates to account for the stability of the hydrogen atom and for the series of spectral lines emitted by it.

(a) The atom consists of a positively charged nucleus at its centre. If Z be the atomic number, the charge on the nucleus is equal to Ze, where e is the magnitude of electronic charge.

The negatively charged electrons revolve only in a few permitted circular orbits around the nucleus. While revolving in these circular orbits, an electron does not radiate energy. These nonradiating orbits are called stationary orbits.

(b) The permissible orbits in which electron can revolve, are those for which angular momentum, $I\omega$ of the electron is an integral multiple of $h/2\pi$, where h is Plank's constant. Thus for these orbits.

$$\text{angular momentum } (I\omega) = m\,va = \frac{nh}{2\pi} \qquad (29.1)$$

Where m and v are the mass and speed of the electron, a is the radius of the permitted orbit and n is a positive integer (n = 1,2,3..............) called the ***principal quantum number.***

(c) ***To explain the spectral lines, Bohr postulated that emission of energy takes place when an electron jumps from a higher to a lower orbit***, while absorption of energy takes place when an electron jumps from lower to higher orbit. Thus if E_i and E_f are the energies associated with two permitted orbits, the frequency ν of the radiation emitted or absorbed is given by the relation.

$$E_i - E_f = h\nu \qquad (29.2)$$

This relation is also known as ***Bohr frequency condition.***

29.4.2 Radii of the Permitted Orbit

Consider Figure 29.4. Let the electron (mass m) in the hydrogen atom revolve around the nucleus in its nth orbit.

Let a_n and v_n be the radius and the speed of the electron in its nth orbit around the proton of mass m.

The radius a_n is measured from the center of the nucleus.

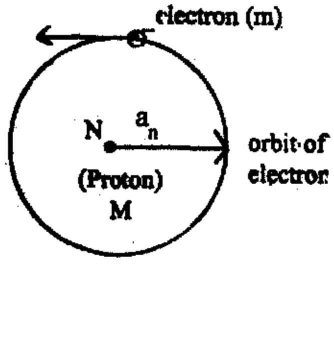

Fig. 29.4: A model of the atom

To keep the electron moving in the nth orbit with velocity v_n, the centripetal force must be provided by the electrostatic force of attraction between the electron and the nucleus.

As we know the centripetal force is given by $f_c = \frac{mv_n^2}{a_n}$ (29.3)

The electrostatic force of attraction between the electron and the nucleus

$$f_e = \frac{kZe^2}{a_n^2} \quad (29.4)$$

where k is a constant given by

$$k = \frac{1}{4\pi\varepsilon_o} \quad (29.5)$$

Here, ε_o is the permittivity of free space = 8.85418×10^{12} Farad/m and Z the atomic number of the nucleus.

The value of constant k is 9×10^9 mF^{-1}

From equations (29.3) and (29.4) we have $f_c = f_e$

$$\text{i.e, } \frac{mv_n^2}{a_n} = \frac{kZe^2}{a_n^2}$$

$$a_n = \frac{kZe^2}{mv_n^2} \quad (29.6)$$

From Bohr's second postulate given earlier, we will recall that the angular momentum ($=mv_n a_n$) of the electron in the nth orbit is an integral multiple of $h/2\pi$

$$mv_n a_n = \frac{nh}{2\pi}$$

$$\text{Thus, } a_n = \frac{nh}{2\pi mv_n} \quad (29.7)$$

Equating eqns. (29.6) and (29.7), we get

$$\frac{kZe^2}{mv_n^2} = \frac{nh}{2\pi mv_n}$$

$$v_n = \frac{2\pi kZe^2}{nh} \quad (29.8)$$

Substituting this value of v_n in equation.(29.7) we get, radii of permitted orbit as,

$$\boxed{a_n = \frac{n^2 h^2}{4\pi^2 mkZe^2}} \qquad (29.9)$$

The above equation gives radii of the permitted orbits in which the electron revolves around the nucleus. n is called the principle quantum number. In case of Hydrogen atom $Z=1$

Substituting $Z=1$ and $n=1$ in eqn. (29.8), we get the speed of electron in the first orbit of the hydrogen atom as

$$\bar{v}_1 = \frac{2\pi ke^2}{h} \qquad (29.10)$$

On substituting the value of the constant h, e and k, the value of speed (v_1) comes out to be a value which is of the order of 1/137 of the velocity of light, and from eqn. (29.9), we get radius of the first orbit as

$$a_1 = \frac{h^2}{4\pi^2 mke^2} \qquad (29.11)$$

Substituting the values m, k, e and h in eqn (29.11), we get the radius of the first permitted orbit.

$$a_1 = 0.53\text{A}^\circ, \ (1\text{A}^\circ = 10^{-10}\text{ m})$$

Similarly for $n=2$

$$a_2 = (2^2)\, a_1 = 4\, a_1$$

and so on, $a_n = (n^2)\, a_1$

29.4.3 Energy of the Electron

Let us first calculate the energy of the electron in the nth orbit.

Let E_n be the total energy of the electron in the nth orbit.

It is equal to

E_n = Kinetic Energy + Potential Energy

or E_n = K.E. + P.E. (29.12)

The K.E. of the electron in nth orbit is given by

$$\text{K.E.} = \frac{1}{2}\, m\,(v_n)^2 \qquad (29.13)$$

From equn (29.6) we get.

$$\text{K.E.} = \frac{1}{2}\,\frac{kZe^2}{a_n} \qquad (29.14)$$

The potential energy of the electron is gives by

$$P.E. = -\frac{k Ze^2}{a_n} \qquad (29.15)$$

Therefore, the total energ , E_n is

$$E_n = \frac{1}{2}\frac{kZe^2}{a_n} - \frac{kZe^2}{a_n} \qquad (29.16)$$

$$E_n = -\frac{1}{2}\frac{kZe^2}{a_n} \qquad (29.17)$$

Substituting the value of a_n form eqn. (29.9), we get

$$\boxed{E_n = -\frac{2\pi^2 k^2 Z^2 e^4}{n^2 h^2} m} \qquad (29.18)$$

The negative sign of the total energy indicates that the electron is bound to the nucleus. For hydrogen atom the energy of the electron in the first orbit is obtained by putting $Z = 1$ and $n = 1$ in eqn (29.18) as

$$E_1 = -\frac{2\pi^2 k^2 e^4 m}{h^2} m \qquad (29.19)$$

This is the lowest energy level and is called the ***ground state of atom.*** On substituting the values of the constants, in the equation 29.19 , the value of E_1 comes out to be 13.6 eV. Therefore, $E_n = -13.6\, n^2$ eV. Energy levels E_2, E_3, corresponding to quantum numbers 2,3.......... are called ***excited states.*** For $n = \propto E = 0$ It means that the electron is no longer bound to the nucleus and becomes free.

We all know that when the excited electron jumps from a higher orbit to lower orbit it emits radiation (photon)

Let the electron jumps from the orbit n_i to n_f and let E_i and E_f be the corresponding energies of the electrons then

$$h\nu = E_i - E_f \qquad (29.20)$$

Where ν is the frequency of the photon or radiation emitted, we can write it as,

$$\nu = \frac{E_i - E_f}{h}$$

Using eqn (29.18), we get

$$\nu = \frac{2\pi^2 k^2 Z^2 e^4 m}{h^2}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \qquad (29.21)$$

It is customary to write the result in terms of a wave number . A *wave number*($\bar{\nu}$) is *defined as the number of waves per meter,*

$$\bar{\nu} = \nu/c,$$

Here c is the velocity of light.

Also we know that $c = \nu\lambda$, λ being the wavelength of the radiation

$$\therefore \quad \overline{\nu} = \frac{1}{\lambda} = \frac{\nu}{c}$$

for eqn. (29.21)

$$\overline{\nu} = \frac{2\pi^2 k^2 Z^2 e^4 m}{ch^3}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \qquad (29.22)$$

For hydrogen atom Z = 1

$$\therefore \quad \overline{\nu} = \frac{2\pi^2 k^2 e^4 m}{ch^3}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \qquad (29.23)$$

$$\text{or} \qquad \overline{\nu} = R_H\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \qquad (29.24)$$

Here R_H is a constant called ***Rydberg constant*** and is equal to

$$R_H = \frac{2\pi^2 k^2 e^4 m}{ch^3} = 1.097 \times 10^7\ \text{m}^{-1}$$

29.4.4 Spectrum of Hydrogen

The spectrum of hydrogen as obtained from eqn. (29.24) consists of several series of lines as shown in Fig 29.5 They are as given below,

(I) Layman Series: Bohr showed that if the electron jumps from energy levels 2,3,4 (ie $n_i = n_2, n_3, n_4$..........) to energy level 1 i e $n_f = 1$, or the ground state, a series of spectral lines are obtained and are called *Lyman Series*, These lie in the *ultra violet region*.

(ii) Balmer Series : If the electron jumps from higher orbits or energy levels n_i = 3,4,5,......... to energy level 2 (n_f = 2) a series of spectral lines are obtained and are known as ***Balmer Series****. These lie in the ultraviolet and visible region.*

(iii) Paschen Series: When electron jumps from higher energy levels n_i = 4,5, 6 .... to n_f = 3. the series of spectral lines are obtained are called ***Paschen Series***. These lie in the *infrared region.*

(iv) Brakett Series: This is obtained when electron jumps from n_i = 5,6,7.......... energy level to energy level 4(i.e n_f = 4). These lie in the *infrared region.*

(v) Pfund Series : In this case n_i = 6,7.......... and n_f = 5. These lie in the *far infrared region.*

(iv) Humphree Series: Here n_i = 7,8 .... and n_f = 6. These lie in *far infrared region.*

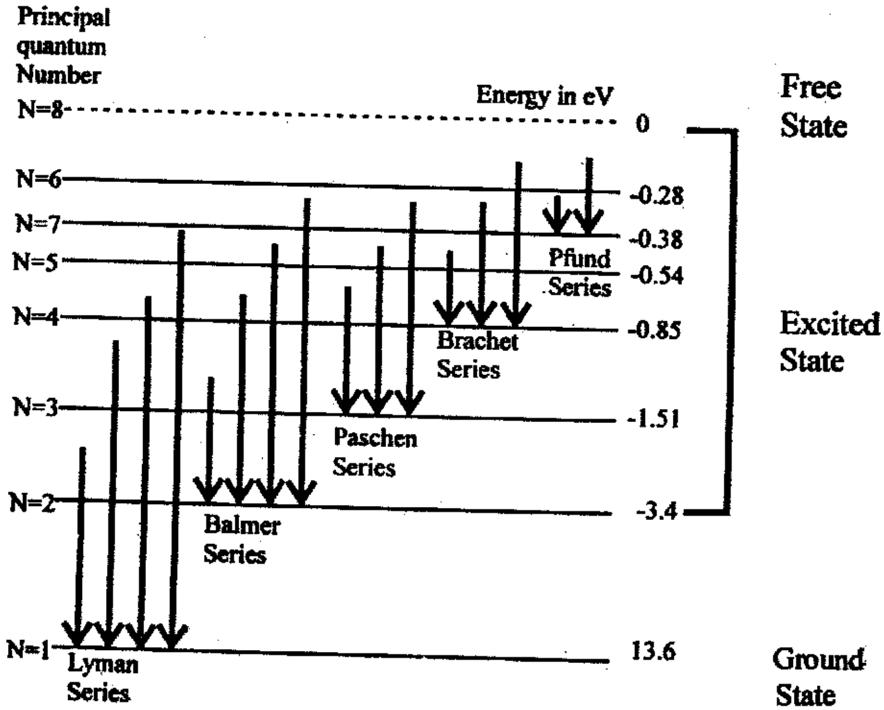

Fig. 295: Energy levels of hydrogen atom

The energy required to remove an electron from $n = 1$ ground state to $n = \infty$ orbit is called* ionization energy *or ionization potential for the hydrogen atom. The value of ionization potential for hydrogen atom in 13.6 V.

29.4.5 Limitations of Bohr's Theory

Though Bohr's theory had overcome several shortcomings of Rutherford's atomic model, it had some limitations. These were,

(i) The orbits were considered to be circular where as elliotical orbits were aiso possibl it means Bohr's theory was not general.

(ii) Some spectral lines of hydrogen were not single lines. They consisted of closely packed lines with slightly different frequencies. This theory could not account for such fine features of lines in the hydrogen atom.

Because of the limitations of the Bohr's model, efforts were made to find proper atomic model which could overcome all the drawbacks. At present the most satisfactory atomic model is wave mechanical atom model based on quantun theory. About which you will have an opportunity to study in higher classes.

29.5 ELECTRONIC CONFIGURATIONS IN ATOMS

Bohr and Stoner showed that all electrons in an atom do not occupy the same orbit. They gave an empirical rule according to which the maximum number of electrons in an orbit is $2n^2$, where n is Principal or total quantum number. Thus, the maximum number of electrons in different orbits are as follows:

First Orbit	$(n = 1)$	$= 2(1)^2 = 2$ electrons
Second orbit	$(n = 2)$	$= 2(2)^2 = 8$ electrons
Third orbit	$(n = 3)$	$= 2(3)^2 = 18$ electrons
Fourth orbit	$(n = 4)$	$= 2(4)^2 = 32$ electrons

The orbits corresponding to n = 1,2,3,4 .... etc. are referred as K; L; M; N............ shells, respectively.
The electrons in the outermost shells, not completely filled (i.e. having electrons less then $2n^2$ are called valence electrons.

29.5.1Quantum Numbers

Bohr considered the orbit of the electron as circular and introduced the ***principal quantum number n***, to determine the total energy of an electron in an atom.
Later Sommerfield extended Bohr's theory to solve the problem in which circular as well as elliptical orbits were permissible for electrons in an atom. He introduced a few quantum numbers as given below.

Azimuthal Quantum Number

Sommerfield retained the principal quantum number n. He introduced another quantum number, called the ***azimnthalh orbital quantum number***, ℓ. For a given value of n, ℓ can take any positive values from 0 to (n-ℓ) i.e. 0,1,2,3........... n- ℓ Thus for n=1 the allowed value of ℓ is only 0. For higher orbits,say for n= 5 it can assume the values 0,1,2,3,4, and so on. Quantum number of ℓ is associated with orbital angular momentum.

Magnetic Quantum Number

With the development of high resolution spectroscopy it was discovered that lines of the Balmer series, for example, were not single. It consisted of six lines, separate and distinct, but all of very nearly the same frequency. To explain this, another quantum number m_ℓ called that ***magnetic quantum number*** was introduced. For a given value of ℓ, it can take integral values from - ℓ to +ℓ i.e. values ℓ, (ℓ-1), 0,(ℓ -1),ℓ (i.e (2ℓ+1) possible values in all). For example for l=2, m_l has five values -2,-1,0,1,2,.

Spin Quantum Number

Finally to account for the fine structure of some of the spectral lines, it became necessary to imagine each electron as a sphere spinning about an axis, while moving in its orbit around the nucleus. There is angular momentum associated with a spinning particle and the magnitude of this is specified by another quantum number, s, called the ***spin*** of the particle. For an electron s = 1/2. Like an orbiting electron, a spinning electron is also associated with a magnetic moment. Yet another quantum number corresponding to the spin magnetic moment of the quantum number m_s, was introduced. For a given, s it can take the values s, s-1, s– 2,................s. For the electron m_s take only two values $+\frac{1}{2}$ or $-\frac{1}{2}$.

Thus, there are four quantum numbers n, ℓ, m_ℓ and m_s necessary to completely specify the state of an electron in an atom, (s is fixed and equal to $\frac{1}{2}$). To summarize, the different quantum numbers and the values they are permitted to take are given below:

Quantum number	Quantized values
(i) n (principle quantum number)	1,2,3,........n
(ii) (ℓ orbit quantum number)	0,1,2,3,4,.....(n-ℓ)
(iii) m_ℓ (magnetic quantum number)	-ℓ,-(ℓ-1), 0, (ℓ-1),ℓ
(iv) m_s (spin magnetic quantum number)	$-\frac{1}{2}, +\frac{1}{2}$

INTEXT QUESTIONS 29.2

1 *Who suggested the first most successful model of the atom ?*

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2. *In case of hydrogen atom what is the radius of the first permitted orbit (n=1) ?*

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3. *Let a_1 and a_2 be the radii of the first and second permitted orbits in the hydrogen atom, respectively. What is the value of a_2 interms of a_1 ?*

Choose the correct answer

4 *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

5 *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

29.6 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.
- The nucleus is surrounded by a cloud of electrons whose 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 Neils Bohr based on some postulates which are as follows:
  (i) The atom consists of a positively charge a nucleus at its centre;
  (ii) Electrons revolve only in few permitted orbits around the nucleus;
  (iii) Electrons do not radiate energy when revolving in permitted orbits. These orbits are called stationary orbits.
- Coulomb's law of electrostatic force and Newton's law of motion apply in the domain of the atom.
- Permissible orbits for electrons are those for which angular momentum $(I\omega) = nh/2\pi$
- Emission or absorption of energy takes place when electron jumps from a higher to a lower orbit or from a lower to a higher permitted orbits.
- The radii of the permitted orbits in which the electron is free to revolve around the nucleus are given by

$$a_n = \frac{n^2h^2}{4\pi^2 mkZe^2}$$

- In case of hydrogen atom, the radii of first permitted orbit is $a_1 = 0.53$ A'
- The energy of the electron in the *n*th orbit is given by

$$E_n = -\frac{2\pi^2 k^2 Z^2 e^4 m}{n^2h^2}$$

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

- Wave number ($\bar{\nu}$) is defined as the number of waves per meter. Wave number is related to energy levels as follow.

$$\bar{\nu} = R_H\left[\frac{1}{n_i^2} - \frac{1}{n_f^2}\right]$$

- Bohr's theory has some limitations. They are as fellow :

  (i) Orbits are considered to be circular, but they may be elliptical also.

  (ii) The theory could not account for the observed closely packed lines with slightly different frequencies in the spectrum of hydrogen.
- There are four quantum numbers namely
  (i) Principle quantum number, n
  (ii) Orbital quantum number , ℓ
  (iii) Magnetic quantum number, m_l
  (iv) Spin magnetic quantum number, m_s

29.7 TERMINAL QUESTIONS

1. In Rutherford's scattering experiment why do most of the α - particles passe straight through the target foil ?
2. Distinguish between Thomson's plum pudding model and Rutherfords planetary model of atom.
3. Why did Rutherford assume hat electrons revolve in circular orbits around the nucleus?
4. What is the ratio of the energies of the hydrogen atom in its first excited state to that in it second excited state?
5. What is the SI unit of Rydberg's constant ?
6. In Rutherford's α - particle scattering experiment, what observation led him to predict the existance of nucleus ?
7. The Rydberg constant for hydrogen in 1096700 m^{-1} calculate the short and long wave-length limits of Lyman series.
8. How many times does the electron of H-atom go round the first orbit in 1s?
9. Describe Rutherford's scattering experiment and discuss its conclusions.
10. Discuss the shortcomings of Rutherford's model.
11. Describe the postulates of Bohr's model of the atom.
12. Derive the expression of the energy of the electron in the nth orbit of hydrogen atom.
13. Choose the correct answers:

(a) Pauli's exclusion principle states that
  (i) No two electrons in an atom can have the same set of quantum numbers.
  (ii) Two electrons can have the same set of quantum numbers.
  (iii) All the electrons can have same set of quantum number .
  (iv) None of the above

(b) In case of hydrogen atom the radius of the electron in its nth orbit is proportional to
  (i) $1/e$ (ii) e (iii) $1/e^2$ (iv) e^2

(c) The total energy E_n of the electron in the nth orbit of hydrogen atom is proportional to.
  (i) $1/n^4$ (ii) $1/n^2$ (iii) $1/n^2$ (iv) $1/n$

(d) The energy required to remove an electron from n = 1 to n = in the case of hydrogen atom is
  (i) 13.6 eV (ii) 13.6 eV (iii) 13.6 MeV (iv) 13.6 keV

(e) For hydrogen atom when the electron jumps from higher energy level n = 5,6,7,....... etc to the energy level n = 4, a set of spectral lines are obtained. These are called
  (i) Balmer Series (ii) Bracket Series (iii) Paschen Series (iv) Lyman Series

14. Calculate the radius of the third and fourth permitted orbits of electron in the hydrogen atom.
15. Discuss the limitations of Bohr's theory of the atom.

16. The energy transition in H-atom accurs from $n = 3$ to $n = 2$ energy level. Given R $= 1.097 \times 10^{7}\ m^{-1}$.
    (i) What is the wavelength of the emitted radiations ?
    (ii) Will this radiation lie in the range of visible light ?
    (iii) To which spectral series does this transition belong ?
17. The ionisation potential of hydrogen is 13.6 volt. What is the energy of the atom in n = 2 state ?

CHECK YOUR ANSWERS

Intext Questions 29.1

1 a (iii), b (ii) c (i) d (i)

2 It could not explain the large angle scattering of particles from the gold foils as observed by Rutherford.

Intext Questions 29.2

1. Neils Bohr
2. 5.3×10^{-11}m
3. 4 a
4. (iv) (5) (i) (6) (ii)

Terminal Questions

7. $\lambda_s = 911.4$A°, $\lambda = 1215$ A
8. 6.57×10^{15} Hz
13. a (i), b (iii), c (iii), d (i), e (ii)
16. (i) 6563A°, (ii) Visible, (iii) Balmer series
17. -3.4 eV