SRIMAAN

UG-TRB

PHYSICS

UNIT-V- ATOMIC AND NUCLEAR PHYSICS

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INTRODUCTION

The eventual constituent of matter is called atom which was considered indivisible for a longtime. Later on, the experiments performed by J . J Thomson, Wilson, Millikan and others showed beyond doubt the existence of negatively charged particles, called the electrons within the atom. Since atom as a whole is neutral, it must have an equal positive charge. J.J. Thomson suggested atomic model that has a large sphere of positive charge with negative electrons embedded in it. But this model was not accepted. After this, Rutherford formulated an atomic model according to which an atom consists of a central positively charged nucleus surrounded by planetary electrons. The number of orbital electrons is equal to the number of protons in the nucleus. In this unit, we shall deal with various atomic models like Thomson model, Rutherford model, Bohr's model, Sommerfeld's model.

ELECTRONIC SPECIFIC CHARGE

The ratio of electronic charge (e) to the electronic mass (m) is known as electronic specific charge.

Thus, electronic specific charge = $\frac{e}{m}$

$$= \frac{1.6\times10^{-19}\ \text{Coulomb}}{9\times10^{-31}\ \text{Kg}} = 1.77\times10^{11}\ \text{Coulomb/ Kg}$$

This value is constant.

ATOMIC MODELS

Any worthwhile model of the atom must be able to account for the following established properties of the atom-

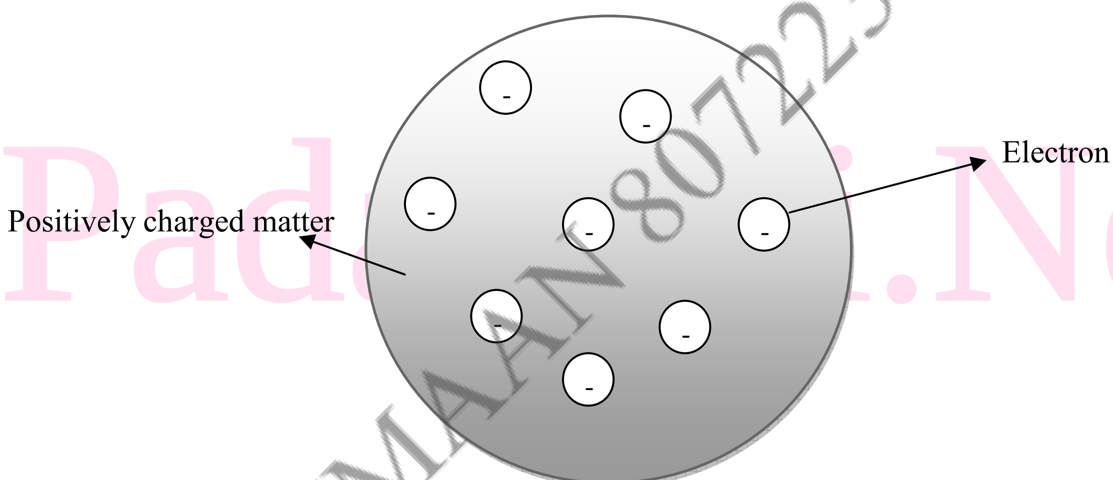
(i) The atom is electrically neutral
(ii) The atom of different elements is stable.
(iii) Every atom, under proper conditions, emits a characteristic spectrum.
(iv) Atoms of different elements have different chemical properties.

(v) Atoms are arranged in the periodic table and show 'periodicity' in their properties.

We now discuss some of the models of the atom that can account for the most of the above mentioned expectation.

THOMSON'S ATOMIC MODEL

In 1898, Sir J.J. Thomson, for the first time, suggested a model for the atom which is known as 'Thomson's model'. According to this model, an atom is a positively charged sphere of radius 10^{-10} meter, in which mass and positive charge of the atom are uniformly distributed. The electrons are embedded here and there within this sphere. The number of electrons is such that their negative charge is equal to the positive charge of the atom. Thomson's atomic model is



Thomson's atomic model

This model is equivalent to water-melon where seeds are embedded here and there (equivalent to electrons) and the rest part (reddish part) is equivalent to the positively charged matter. Hence this atomic model is also called 'water-melon model' or 'plum pudding model'. Various phenomenon like thermionic emission, photo-electric emission and ionization, were explained on the basis of this model.

Drawbacks of Thomson's Atomic Model

Thomson's atomic model explained various phenomenon however this model has some drawbacks-

(i) In order to explain the emission of light from atom, it was assumed that whenever an atom receives energy from outside, the electrons in it start vibrating and radiate electromagnetic radiation (light waves) of the frequency of their vibration. But, according to this explanation, there should be only one line in the spectrum of hydrogen atom (which has a single electron), whereas, in fact, it has many lines. Thus, this model was unable to explain this point.
(ii) This model could not predict any reason for α-particle scattering.

These were failures of Thomson's atomic model. Due to these drawbacks, Thomson's atomic model could not be accepted. This led to the idea of another atomic model, known as the nuclear atom model.

RUTHERFORD'S EXPERIMENT OF α-PARTICLE SCATTERING

α- particles are doubly charged helium (He) atoms that have lost both of their electrons. Thus, the mass of an α-particle is four times the mass of a hydrogen atom and a positive charge two times the charge of a proton.

If a beam of α-rays is made to fall upon a photographic plate in vacuum, the shadow image formed there has sharp and clear edges. But, if air or some other gas is introduced or a screen having thin foil of metal is placed in the path of α-rays, the image formed is not sharp but becomes diffused. This spreading out of the stream of α-particles on passing through thin layers of matter is known as scattering.

The schematic diagram of Rutherford's experiment on scattering of α-particles is shown in following figure 2. α-particles from radioactive source are incident on a thin gold foil of thickness ~ 10^{-7} m, beyond which is placed a screen coated with zinc sulphide. A-particle on striking the screen caused tiny flashes which could be viewed and counted. The following observations are noted from this experiment-

(i) Most of the α-particles passing straight through the gold foil indicating thereby that the atoms are hollow.
(ii) Some of the α-particles are scattered through small angles but a few of them were deviated through a large angle and occasionally a particle went back along the directions from which it comes i.e. deviates through about 180^0. This large scale scattering is called 'anomalous scattering'.

Thomson's atomic model could not explain these observations. Rutherford, therefore, was forced to picture an atom in which the entire positive charge and almost whole of the atomic mass are concentrated in a tiny central core called the nucleus about which the electrons are grouped in some sort of configuration.

Assumptions of Rutherford's Theory of α-Particle Scattering

Rutherford developed his theory of α-particle scattering on the basis of a number of assumptions. These assumptions are mentioned below-

(i) The first and foremost assumption, that was the basis of his entire theory, was the concept of the nucleus. Rutherford assumed that the entire positive charge of the atom was concentrated in a very tiny, very massive part at its centre. The negatively charged electrons were present in a relatively much vaster volume around this nucleus and the major part of the atom was thus just empty space or vacuum.

(ii) α-particle as well as the nucleus

Theory of α-Particle Scattering Experiment

Let us explain α-particle scattering. Rutherford assumed that when an α-particle approaches the positive nucleus, Coulomb's repulsive force is given by-

$$F = \frac{1}{4\pi\varepsilon_0}\frac{(+Ze)(+2e)}{r^2}$$

Here, +Ze and +2e are the charges on nucleus and α-particle respectively and r is the distance between the nucleus and approaching α-particle..

This force comes into play and hence the α-particles are deflected from their path. In this expression ε_0 is the permittivity of free space and $\frac{1}{4\pi\varepsilon_0} = 9 \times 10^9$ Newton-m^2/Coulomb^2

Rutherford also derived a mathematical formula describing the scattering of α-particles by thin foil. For the purpose of calculations, following assumptions were made-

(i) The nucleus and the particle, both act as point charges and masses.

(ii) The nucleus is considered to be so heavy that its motion during the scattering process may be ignored.

(iii) The scattering is caused by repulsive electrostatic force between the nucleus and α- particle which obeys Coulomb's inverse square law.

The calculations show that the number of α-particles received per square cm. of a screen placed at an angle θ to the direction of incident particles as a distance R from the scatter, i.e. thin metal foil is given as-

$$N(\theta) = K' \frac{N_0}{sin^4(\frac{\theta}{2})} = K' N_0 \,\text{cosec}^4 (\theta/2) \;\log \left(\frac{1}{sin^4\frac{\theta}{2}}\right)$$

A graph is plotted between log $\left(\frac{1}{sin^4\frac{\theta}{2}}\right)$ and log N(θ) which came to be a straight line as was expected from Rutherford scattering formula indicating that-

$$N(\theta) \propto \frac{1}{sin^4\frac{\theta}{2}}$$

The experimental data of Rutherford's scattering of α-particles also led to an estimate of the size of the nucleus. The velocity of those α-particles which meet the nucleus in head on collision i.e. scattered through 180^0, becomes zero at the distance of closet approach to the nucleus.

Let us consider an α-particle travelling initially with a constant velocity v towards the nucleus at a very large distance from it. As the α-particle approaches the nucleus, its velocity goes on decreasing due to electrical repulsion. As a result, the kinetic energy of α-particle goes on decreasing while its potential energy increases. At the distance of closest approach (r_0), the α-particle is stopped momentarily and thus whole of its initial kinetic energy is changed into potential energy.

If (+2e) is the charge and m the mass of α-particle, Z the atomic number and (+Ze) the charge on the nucleus, then-

Initial kinetic energy of α-particle, $K = (1/2)\, mv^2$

The potential energy of α-particle when it is at the distance of closest approach r_0 from the nucleus,

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

Now,

$$K = U$$

$$(1/2)\, mv^2 = \frac{1}{4\pi\varepsilon_0}\frac{(+Ze)(+2e)}{r_0}$$

or

$$r_0 = \frac{1}{4\pi\varepsilon_0}\frac{4Ze^2}{mv^2}$$

Conclusions from Rutherford's Scattering Formula

The Rutherford scattering formula tells us that –

(i) For a given foil and for α-particles of a given incident energy, the number N(θ) is directly proportional to $\frac{1}{sin^4\frac{\theta}{2}}$ or $cosec^4(\theta/2)$

(ii) For α-particles, detected at a particular angle of scattering θ, the number N(θ) is-
(a) directly proportional to the number of foil atoms per unit volume.
(b) directly proportional to the square of the atomic number of the foil atoms.
(c) directly proportional to the thickness of the foil of a given element and
(d) inversely proportional to the square of the kinetic energy of the incident alpha particles.

RUTHERFORD'S ATOMIC MODEL

In 1911, Rutherford presented a model of the atom, called 'Rutherford's atomic model'. According to this model, the mass of the atom (leaving the mass of its electrons) and its entire positive charge are concentrated at the centre of the atom in a nucleus of radius ~ 10^{-15} meter. Electrons have no place inside the nucleus. Around the nucleus, electrons are distributed in a hollow sphere of radius ~ 10^{-10} meter. The dimensions of the radius and of the electron are negligibly small as compared to the overall size of the atom so that most of the volume occupied by an atom is actually an empty space. In this way, the discovery of the nucleus of the atom is due to Rutherford. The total negative charge of the electrons is equal to the positive charge of the nucleus.

Rutherford assumed that the electrons in the atom are not stationary but are revolving around the nucleus in different orbits and the necessary centripetal force is provided by the electrostatic force of attraction between the electron and the nucleus.

Shortcomings of Rutherford's Atomic Model

Rutherford's atomic model was supported by the periodic table of the elements. However, this model suffers from two main defects-

(i) Rutherford proposed that electrons revolve at a high speed in circular orbits around the positively charged nucleus. When a charged particle i. e. electron revolves around positively charge nucleus, it needs to be accelerated so as to keep it moving in circular orbits. However, according to electromagnetic theory, whenever a charged particle such as an electron is accelerated around another charged center (nucleus) which are under force of attraction, there will be continuous radiation of energy. This loss of energy would slow down the speed of the electron. This would reduce the radius of the electron–orbit. Eventually the electron would fall into the nucleus. The result would be that the atom would collapse. But this does not happen. Thus Rutherford's atom could not explain the stability of the atom. Failure of Rutherford's model i.e. reduction of radius of orbit is shown below.

BOHR'S ATOMIC MODEL

There were some shortcomings of Bohr's atomic model. Rutherford's model could not explain the stability of the atom and line spectrum. In 1913, Neil Bohr removed these difficulties by the application of Planck's quantum theory. Let us discuss the postulates as proposed by Bohr. Bohr proposed the following postulates-

(i) Electrons can revolve only in those orbits in which their angular momentum is an integral multiple of $h/2\pi$, where 'h' is the Planck's constant ($h = 6.64 \times 10^{-34}$ J-sec).
If 'm' is the mass of the electron and it is revolving with velocity 'v' in an orbit of radius 'r' , then its angular momentum = mvr.
According to this postulate, we have-

$$mvr = n\,(h/2\pi)$$

where n = 1, 2, 3, ………. (integer). 'n' is called 'principle quantum number' of the orbit. Thus, according to Bohr's atomic model, electrons can revolve only in certain discrete orbits of definite radii, not in all. These are called 'stable orbits'.

(ii) While revolving in stable orbits, the electrons do not radiate energy inspite of their acceleration towards the centre of the orbit. Hence, the atom remains stable and is said to exist in a stationary state.

(iii) An atom radiates energy only when an electron jumps from a stationary orbit of higher energy to one of lower energy.

If the electron jumps from an initial orbit of energy E_i to a final orbit of energy E_f ($E_i > E_f$), a photon of frequency $\nu = (E_i - E_f)/h$ is emitted.
Based on these postulates, Bohr derived the formulae for the radii of the stationary orbits and the total energy of the electron in the orbit.
Let us derive Bohr formulae for the radius of nth orbit, velocity of electron in permitted orbits, the radius of nth orbit and the energy of the electron in the nth orbit.
Let us consider an atom whose nucleus has a positive charge Ze, where Z is the atomic number. Let an electron of charge (-e) and mass 'm' moves round the nucleus in an orbit of radius 'r' with velocity 'v'.

The electrostatics force of attraction between the nucleus and the electron is-

$$4\pi\varepsilon_0 \frac{(Ze)(e)}{r^2} \quad .$$

The centrifugal force on the electron is-

$$F_c = \frac{mv^2}{r} \quad .$$

For the stability of the atom, two forces should be equal i.e. $F_e = F_c$

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

or $$mv^2 = \frac{Ze^2}{4\pi\varepsilon_0 r} .$$

From the first postulate, we have-

$$mvr = n\frac{h}{\pi}; \qquad n = 1, 2, 3, \ldots\ldots\ldots\ldots$$

Squaring both sides, we have-

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

Dividing equation by equation , we get-

$$\frac{mv^2}{m^2v^2r^2} = \frac{\frac{Ze^2}{4\pi\varepsilon_0 r}}{n^2\left(\frac{h}{2\pi}\right)^2}$$

or $$r = \frac{n^2h^2\varepsilon_0}{\pi m Z e^2} .$$

In general, we can write-

$$r_n = \frac{n^2h^2\varepsilon_0}{\pi m Z e^2} .$$

This is the expression for the radius of nth orbit. It is clear that, $r_n \alpha\, n^2$.

Now from equation

$$v = \frac{nh}{\pi m r} .$$

Putting for 'r' in the above equation we get-

$$v = \frac{Ze^2}{h\varepsilon_0 n} .$$

This is the expression for the velocity of electron in permitted orbits.

Clearly, $$v\, \alpha \frac{1}{n}; \qquad n = 1, 2, 3, \ldots\ldots\ldots\ldots$$

In general, we can write-

$$v_n = \frac{Ze^2}{h\varepsilon_0 n} .$$

This is the velocity of electron in nth orbit. For n = 1, the velocity will be maximum. This shows that the velocity of electron is maximum in the lowest orbit (n = 1) and goes on decreasing in higher orbits.

We can generalize this expression for hydrogen atom. For hydrogen atom, Z = 1, equation becomes-

$$r_n = \frac{n^2 h^2 \varepsilon_0}{\pi m e^2} \quad .$$

For first orbit, n = 1, therefore, the radius of first orbit,

$$r_1 = \frac{h^2 \varepsilon_0}{\pi m e^2}$$

$$= \frac{(6.64 \times 10^{-34}\ J\ sec)^2\ (8.85 \times 10^{-2} \frac{F}{m})}{3.14 \times (9.1 \times 10^{-31}\ kg)(1.6 \times 10^{-19} C)^2}$$

$$= 0.53 \times 10^{-10} \text{ metre}$$

$$= 0.53\ A^0 \quad .$$

This is called Bohr's radius. Its value is 0.53 A^0.

From equation we have-

$$r_n = n^2 \left(\frac{h^2 \varepsilon_0}{\pi m e^2} \right)$$

$$= n^2 r_1$$

or $$r_n = (0.53)\ n^2\ A^0 \quad .$$

This is the expression for the radius of nth orbit.

For second orbit, n = 2,

$$r_2 = (0.53) \times (2)^2 = 2.12\ A^0 \quad .$$

Similarly, the velocity of electron in nth orbit in hydrogen atom is given as-

$$v_n = -\frac{e^2}{\pi \varepsilon_0 n} \quad \text{[putting Z= 1 in equation]}$$

For first orbit, n = 1,

$$v_1 = -\frac{e^2}{\pi \varepsilon_0} = \frac{(1.6 \times 10^{-19} C)^2}{\times 3.14 \times (8.85 \times \frac{10^{-12} F}{m})} = 2.19 \times 10^6 \text{ m/sec}$$

$$\frac{v_1}{c} = \frac{2.19\times10^6\ m/sec}{3\times10^8\ m/sec} = \frac{1}{137}\quad .$$

The kinetic energy of the electron, $K = (1/2)\ mv^2 = (1/2)\ (\frac{Ze^2}{4\pi\varepsilon_0 r})$

$$= \frac{Ze^2}{8\pi\varepsilon_0 r}\quad .$$

The potential energy of the electron in an orbit of radius 'r' due to electrostatic attraction by the nucleus is given by-

$$U = \frac{1}{4\pi\varepsilon_0}\frac{(Ze)(-e)}{r}$$

$$= -\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r}\quad .$$

The energy E of an electroninan orbit is the sum of kinetic and potential energies. Therefore, the total energy of the electron-

$$E = K + U$$

$$= \frac{Ze^2}{8\pi\varepsilon_0 r} - \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r} = -\frac{Ze^2}{8\pi\varepsilon_0 r}\quad .$$

Putting for r from equation we get-

$$E = -\frac{mZ^2e^4}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right)\quad .$$

In general, $$E_n = -\frac{mZ^2e^4}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right)\quad .$$

This is the expression for the energy of the electron in the nth orbit.

Bohr's Interpretation of the Hydrogen Spectrum

If an electron jumps from an outer initial orbit n_2 of higher energy to an inner orbit n_1 of lower energy, the frequency of the radiation emitted is given by-

$$\upsilon = \frac{E_{n2} - E_{n1}}{h}\quad .$$

But $$E_{n1} = -\frac{mZ^2e^2}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right)$$

and $$E_{n2} = -\frac{mZ^2e^2}{8\varepsilon_0^2h^2}\left(\frac{1}{n_2^2}\right)$$

$$\upsilon = \frac{1}{h}\left[-\frac{mZ^2e^4}{8\varepsilon_0^2h^2}\left(\frac{1}{n_2^2}\right) + \frac{mZ^2e^4}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right)\right]$$

or
$$\upsilon = \frac{mZ^2e^4}{8\varepsilon_0^2h^3}\left(\frac{1}{n^2} - \frac{1}{n_2^2}\right) \quad .$$

The corresponding wavelength is given as-

$$\lambda = \frac{c}{\upsilon}$$

$$= \frac{c}{\frac{mZ^2e^4}{8\varepsilon_0^2h^3}\left(\frac{1}{n^2} - \frac{1}{n_2^2}\right)} \quad .$$

Therefore
$$\frac{1}{\lambda} = \frac{mZ^2e^4}{8\varepsilon_0^2h^3c}\left(\frac{1}{n^2} - \frac{1}{n_2^2}\right)$$

1/λ is called wave number i.e. the number of waves per unit length. It is denoted by $\bar{\upsilon}$.

The quantity $\frac{me^4}{8\varepsilon_0^2ch^3}$ is a constant called Rydberg constant (R).

Therefore, Rydberg constant R = $\frac{me^4}{8\varepsilon_0^2ch^3}$.

Therefore, equation becomes-

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n^2} - \frac{1}{n_2^2}\right) \quad .$$

For hydrogen or hydrogen- like atoms (He^+, Li^+, ……….), Z = 1, therefore equation becomes-

$$\frac{1}{\lambda} = R(1)^2\left(\frac{1}{n^2} - \frac{1}{n_2^2}\right)$$

or
$$\frac{1}{\lambda} = R\left(\frac{1}{n^2} - \frac{1}{n_2^2}\right) \quad .$$

Rydberg constant R = $\frac{me^4}{8\varepsilon_0^2ch^3} = \frac{(9.1\times10^{-31}\,kg)\times(1.6\times10^{-19}C)^4}{8\times(8.85\times10^{-12}F/m)^2\times(3\times10^8\,m/sec)\times(6.62\times10^{-34}\,J\,sec)^3}$

$= 1.097\times 10^7$ per m

The energy expression $E_n = -\frac{mZ^2e^2}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right)$ can be written in terms of Rydberg's constant (R).

Therefore
$$E_n = -Z^2\frac{Rhc}{n^2}$$

or
$$E_n = -\frac{Z^2\times(1.097\times10^7m^{-1})\times(6.62\times10^{-34}Jsec)\times(3\times10^8msec^{-1})}{n^2}$$

$$= -\frac{Z^2 \times (13.6 \times 1.6 \times 10^{-19})}{n^2} \text{ Joule}$$

$$E_n = -Z^2 \frac{13.6}{n^2} \text{eV}$$

For hydrogen, Z = 1

$$E_n = -\frac{Rhc}{n^2} \text{eV}$$

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

Spectral Series of Hydrogen Atom

In this section, we shall discuss five spectral series of hydrogen atom. These series are as follows-

(i) **Lyman series:** When an electron jumps from second, third, fourth,………..etc. orbits to the first orbit, the spectral lines are in the ultraviolet region. Here $n_1 = 1$ and $n_2 = 2, 3, 4, 5, \ldots$

$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right)$; n =2, 3, 4, 5, ……………

This is identified as Lyman series.

(ii) **Balmer series:** When an electron jumps from outer orbits to the second orbit, $n_1=2$ and $n_2 = 3, 4, 5, \ldots$ etc.

$\frac{1}{\lambda} = R\left(\frac{1}{2} - \frac{1}{n^2}\right)$; n = 3, 4, 5, ……………

This series is called Balmer series and lies in the visible region of the spectrum.

The first line in the series (n = 3) is called the H_α line; the second (n=4), the H_β line; the third (n=5) the H_γ line.

(iii) **Paschen series:** Paschen series in the infrared region are given by $n_1 = 3$ and $n_2 = 4, 5, 6, \ldots$ etc.

$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right)$; n = 4, 5, 6……………

(iv) **Brackett series:** If $n_1 = 4$ and $n_2 = 5, 6, 7, \ldots$ etc., we get the Brackett series.

$\frac{1}{\lambda} = R\left(\frac{1}{4^2} - \frac{1}{n^2}\right)$; n = 5, 6, 7,……………

This series lies in the very far infrared region of the hydrogen spectrum.

(v) **Pfund series:** If $n_1 = 5$ and $n_2 = 6, 7, 8, \ldots$ etc., we get the Pfund series.

$\frac{1}{\lambda} = R\left(\frac{1}{5^2} - \frac{1}{n^2}\right)$; n = 6, 7, 8,……………

This series also lies in the very far infrared region of hydrogen spectrum.

By putting n = ∞ in each one of the series, we get the wavelength of series limit, i.e. the last time in the series.

The Energy-Level Diagram

An energy level diagram is a sort of one-dimensional scale of energy along which each electron according to this energy state can be located. Let us represent the equation $E_n = -\frac{mZ^2e^4}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right)$ diagrammatically.

$$E_n = -\frac{mZ^2e^4}{8\varepsilon_0^2h^2}\left(\frac{1}{n^2}\right) = -\frac{(9.1\times10^{-31}\ kg)(1)^2(1.6\times10^{-19}\ C)^4}{8(8.85\times10^{-12}\ F/m)^2(6.62\times10^{-34}\ J\ sec)^2}\left(\frac{1}{n^2}\right) = -\frac{13.6}{n^2}\ \text{eV},\ n = 1, 2, 3, \ldots\ldots\ldots\ldots\ldots$$

The lowest energy level E_1 (for n = 1) is called the normal or ground state of the atom and the higher energy levels E_2, E_3, E_4,……… are called the excited states. As n increases, E_n increases. As n increases, the energy levels become crowded and tend to form continuum.

In the energy-level diagram, the discrete energy states are represented by horizontal lines and the electronic jumps between these states by vertical lines. The following figure shows schematically how spectral lines are related to atomic energy levels.

Shortcomings of Bohr's Atomic Model

Bohr's theory, although was very successful in explaining the spectrum of hydrogen atom and giving valuable information about atomic structure, has the following drawbacks-

(i) The fine structure i.e. individual line of hydrogen spectrum accompanied by a number of faint lines cannot be explained by Bohr's theory as such. The fine structure of spectral lines can only be explained when (a) the relativistic variation in the mass of the electron and (b) electron 'spin' are taken into account.

(ii) Bohr's theory fails to explain the variation in intensity of the spectral lines of an element. The intensity of the spectral lines can be explained by quantum mechanics.

(iii) Bohr's theory fails to explain the spectra of complex atoms. It is only applicable to one-electron atoms such as hydrogen, hydrogen isotopes, ionized helium, etc.

(iv) Bohr's theory fails to explain satisfactorily the distribution of electrons in atoms.

(v) The success of Bohr's theory in explaining the effect of magnetic field on spectral lines is only partial, i.e. it cannot explain the 'anomalous Zeeman effect'.

Example 1: An alpha particle with kinetic energy10 MeV is heading towards a stationary point-nucleus of atomic number 50. Estimate the distance of closest approach.

Solution: Given – kinetic energy K = 10 MeV, atomic number Z = 50

Using formula

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

or $$r_0 = \frac{1}{4\pi\varepsilon_0}\frac{(+Ze)(+2e)}{K} = 9\times10^9 \times \frac{(50\times1.6\times10^{-19})(2\times1.6\times10^{-19})}{10\times10^6\times1.6\times10^{-19}}$$

$$= 1.44\times10^{-14}\text{ m.}$$

Example 2: The number of particles scattered at 60^0 is 100 per minute in an alpha particle experiment, using gold foil. Find out the number of particles per minute scattered at 90^0 angle.

Solution: We know that- $$N(\theta)\ \alpha\ \frac{1}{\sin^4\frac{\theta}{2}}$$

or $$N_1/N_2 = \frac{\sin^4\left(\frac{\theta_2}{2}\right)}{\sin^4\left(\frac{\theta_1}{2}\right)}\ .$$

Given, $N_1 = 100$, $\theta_1 = 60^0$, $\theta_2 = 90^0$

Therefore, $$100/N_2 = \frac{\sin^4\left(\frac{90^0}{2}\right)}{\sin^4\left(\frac{60^0}{2}\right)}$$

or $$N_2 = 25\ .$$

Therefore, the number of particles per minute scattered at 90^0 angle is 25.

Example 3: How many revolutions does an electron in the first Bohr orbit of hydrogen atom make per second?

Solution: According to Bohr's postulate of quantization of angular momentum, we know

TO BE CONTINUED.....

$$mvr = n\,(h/2\pi)$$

=