ATOMIC STRUCTURE

Preface

Whenever we think about any real world problem, materials come into picture which is comprised of small prime units or atoms.

When the whole issue is concerned with chemistry we have to discuss chemical or compound which are made from elements. Each element has its own identity due to a specific structure of its atom. Therefore it is essential to know about structure of atoms and different theories related with it.

After accomplishing this chapter you should be able to -

(i) Know about diligent research work which reaches to real structure of an atom.

(ii) Explain electromagnetic radiations, various types of spectrums and their properties.

(iii) Understand picture of a revolving electron around its nucleus and calculations of its energy, velocity and radius of revolution.

(iv) Familiar with few more advanced atomic model viz. wave mechanical model which leads to the concept of orbital and ofcourse, their shape too.

(v) Explain quantum numbers and their significance.

(vi) Know various principles related to electronic configuration and its representation for all elements.

This book consists of theoretical & practical explanations of all the concepts involved in the chapter. Each article is followed by a ladder of illustration. At the end of the theory part, there are miscellaneous solved examples which involve the application of multiple concepts of this chapter.

Students are advised to go through all these solved examples in order to develop better understanding of the chapter and to have better grasping level in the class.

Total number of Questions in **Atomic Structure** are :

In chapter Examples.	....	26
Solved Examples	....	27
Total no. of questions	....	**53**

ATOMIC STRUCTURE

1. INTRODUCTION

(a) The word atom was first introduced by Ostwald (1803 - 1807) in scientific world.

(b) According to him matter is ultimately made up of extremely small indivisible particles called atoms.

(c) It takes part in chemical reactions.

(d) Atom is neither created nor destroyed

2. DALTON'S ATOMIC THEORY

Dalton proposed the atomic theory on the basis of the law of conservation of mass and law of definite proportions. He also proposed the law of multiple proportion as a logical consequence of this theory. The salient features of this theory are

(a) Each element is composed by extremely small particles called atoms.

(b) Atoms of a particular element are all alike but differ with the atoms of other elements.

(c) Atom of each element is an ultimate particle, and has a characteristic mass but is structureless.

(d) Atom is indestructible i.e. it can neither be destroyed nor created by simple chemical reactions.

(e) Atom of an element takes part in chemical reaction to form molecule.

(f) In a given compound, the relative number and kind of atom are same.

(g) Atoms of different elements combine in fixed ratio of small whole numbers to form compound atoms (now called molecules).

2.1 Merits and Demerits of Dalton's theory :

2.1.1 Merits :

(a) Dalton's theory explains the law of conservation of mass and some other laws of chemical combination.

(b) Atoms of elements take part in chemical reaction is true till today.

2.1.2 Demerits :

(a) There is no mention of atomic weights of elements.

(b) He could not explain that why do atoms of same element combined with each other.

(c) The law of definite proportion fails if different isotopes are used.

3. FUNDAMENTAL PARTICLES

3.1 Properties of electron

(a) Electron was discovered by Sir J.J. Thomson

(b) The charge on the electron is 1.6×10^{-19} coulomb/gm (Millikan)

(c) The molar mass of electron is 5.48×10^{-4} gm/mole

(d) The mass of electron in motion is expressed as

$$m' = \frac{m}{\left(1 - \frac{v^2}{c^2}\right)^{1/2}}$$

where m' = mass of the electron in motion

m = rest mass, v = velocity of the electron

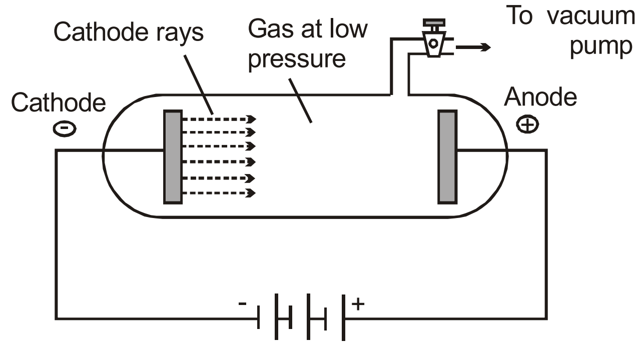
c = velocity of light

(e) In 1897, J.J. Thomson determined the e/m value (charge/mass) of the electron by studying the deflections of cathode rays in electric and magnetic fields. The value of e/m has been found to be -1.7588×10^8 coulomb

(f) The first precise measurement of the charge on the electron was made by Robert A. Millikan. in 1909 by oil drop experiment. Its value was found to be - 1.6022×10^{-19} coulomb.

(g) The mass of electron can be calculated from the value of e/m and the value of e which is 9.1096×10^{-31} Kg.

3.1.1 Cathode rays



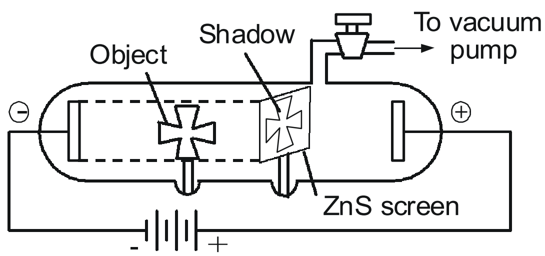
(a) The electron was discovered as a result of the studies of the passage of electricity through gases at extremely low pressures known as discharge tube experiments.

(b) When a high voltage of the order of 10,000 volts or more was impressed across the electrodes, some sort of invisible rays moved from the negative electrode to the positive electrodes these rays are called as cathode rays

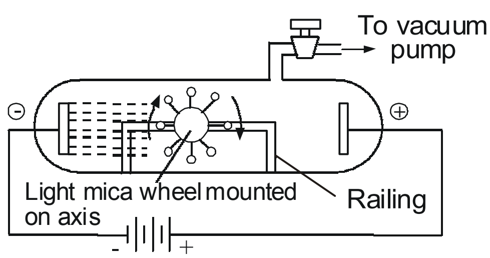
(c) Cathode rays have the following properties.

(i) Path of travelling is straight from the cathode with a very high velocity

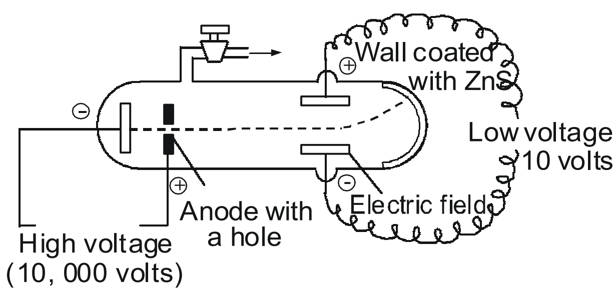
As it produces shadow of an object placed in its path



(ii) Cathode rays produce mechanical effects. If a small pedal wheel is placed between the electrodes, it rotates. This indicates that the cathode rays consist of material part



(iii) When electric and magnetic fields are applied to the cathode rays in the discharge tube, the rays are deflected thus establishing that they consist of charged particles.



(iv) Cathode rays produce X-rays when they strike against hard metals like tungsten, copper etc.

(v) When the cathode rays are allowed to strike a thin metal foil, it gets heated up. Thus the cathode rays possess heating effect.

(vi) They produce a green glow when strike the glass wall beyond the anode. Light is emitted when they strike the zinc sulphide screen.

(vii) Cathode rays penetrate. Through thin sheets of aluminium and other metals.

(viii) They affect the photographic plates

(ix) The ratio of charge to mass i.e. charge/mass is same for all the cathode rays irrespective of the gas used in the tube.

3.2 Properties of proton

(a) Proton was discovered by Goldestein

(b) Proton carries a charge of +1.602 x 10^{-19} coulomb, i.e., one unit positive charge.

(c) Mass of proton is 1.672 x 10^{-27} kg or 1.0072 amu

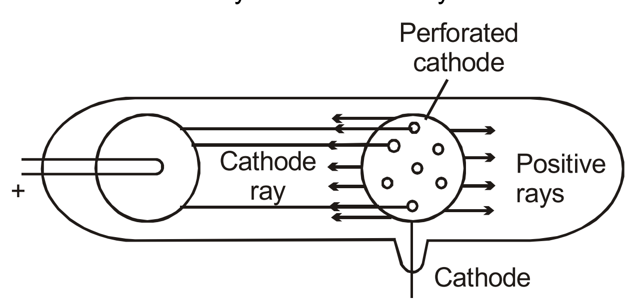
(d) A proton is defined as a sub-atomic particle which has a mass nearly 1 amu and a charge of +1 unit

3.2.1 Positive Rays-Discovery of Proton

(a) The existence of positively charged particles in an atom was shown by E. Goldstein in 1886

(b) He repeated the same discharge tube experiments by using a perforated cathode.

(c) It was observed that when a high potential difference was applied between the electrodes, not only cathode rays were produced but also a new type of rays were produced simultaneously from anode moving towards cathode and passed through the holes or canal of the cathode. These termed as canal ray or cathode ray



(d) Characteristics of Anode Rays are as follows.

(i) These rays travel in straight lines and cast shadow of the object placed in their path.

(ii) The anode rays are deflected by the magnetic and electric fields like cathode rays but direction is different that mean these rays are positively charged.

(iii) These rays have kinetic energy and produces heating effect also.

(iv) The e/m ratio of for these rays is smaller than that of electrons

(v) Unlike cathode rays, their e/m value is dependent upon the nature of the gas taken in the tube.

(vi) These rays produce flashes of light on Zn-S screen

(vii) These rays can pass through thin metal foils

(viii) They are capable to produce ionisation in gases

(ix) They can produce physical and chemical changes.

3.3 Properties of neutron

(a) This was discovered 20 years after the structure of atom was elucidated by Rutherford.

(b) It has been found that for all atoms except hydrogen atomic mass is more than the atomic number. Thus Rutherford (1920) suggested that in an atom, there must be present at least a third type of fundamental particle.

(c) It should be electrically neutral and posses mass nearly equal to that of proton. He proposed its name as neutron.

(d) Chadwick (1932), bombarded beryllium with a stream of α-particles and observed electrically and magnetically neutral radiations.

(e) There were neutral particles which was called neutron. Nuclear reaction is as follows

$$_4Be^9 + {}_2He^4 \longrightarrow {}_6C^{12} + {}_0n^1$$

(f) A neutron is a subatomic particle which has a mass 1.675×10^{-24}g, approximately 1 amu, or nearly equal to the mass of proton on hydrogen atom and carrying no electrical charge.

Examples based on Fundamental Particles

Ex.1 For cathode rays' the value of e/m -

(A) Is independent of the nature of the cathode and the gas filled in the discharge tube

(B) Is constant

(C) Is -1.7588×10^8 coulombs/g

(D) All of the above are correct

Ans.(D)

Sol. Cathode rays consists of electrons which are fundamental particles of matter.

Ex.2 Which has highest e/m ratio –

(A) He^{2+} (B) H^+

(C) He^{1+} (D) H **Ans.(B)**

Sol. Mass of H^+ is minimum

Ex.3 Arrange the following particles in increasing order of values of e/m ratio : Electron (e), proton (p), neutron (n) and α-particle (α) -

(A) n, p, e, α (B) n, α, p, e

(C) n, p, α, e (D) e, p, n, α **Ans.(2)**

Sol.

	Electron	Proton	Neutron	α-particle
e	1 unit	1 unit	zero	2 units
m	1/1837 unit	1 unit	1 unit	4-units
e/m	1837	1	zero	1/2

Ex.4 Mass of neutron is .... times the mass of electron -

(A) 1840 (B) 1480

(C) 2000 (D) None **Ans.(1)**

Sol. Mass of neutron = 1.675×10^{-27} kg, mass of electron = 9.108×10^{-31} kg.

4. NON FUNDAMENTAL PARTICLES

4.1 Positron :

(a) It is also called positive electron and symbolised as $_1e^0$ or e^+.

(b) It was discovered by **ANDERSON** in 1932.

(c) It is the positive counterpart of electron.

(d) Mass of positron is same as electron m = 9.1×10^{-28} g.

(e) Charge of positron is same but opposite signed as electron e = -1.6×10^{-19} C.

(f) It is very unstable and combines with electron producing γ rays.

4.2 Neutrino and Antineutrino :

These are particles of approximately zero masses and zero charge.

4.3 Antiproton :

(a) It was discovered by ***Seagre***.

(b) Mass of this particle is equal to 1.673×10^{-24} g.

(c) Charge of Antiproton is -1.6×10^{-19} C.

4.4 Meson (π) :

(a) It was discovered by **Yukawa** in 1935.

(b) It may possess 3 types of charges.

(c) On the Basis of charge, the meson is of three types, π-meson, μ-meson and neutral meson (π^o).

(d) π-mesons are called pions.

(e) It tells about the stability of nucleus.

(f) The mass of this particle is 200 times of electron i.e. It is heavier than electron but lighter than proton.

5. THOMSON'S MODEL

It states the arrangement of electrons and protons in an atom. The main principles are

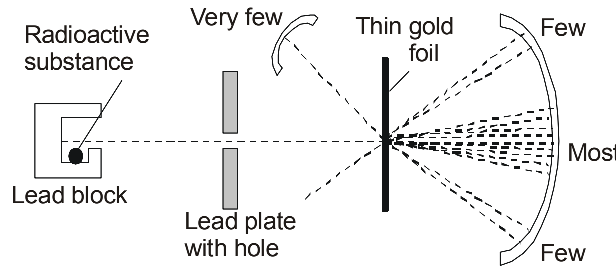
(a) After discovery of electron and proton attempts were made to find out their arrangement in an atom. The first simple model was proposed by J.J. Thomson known as Thomson's atomic model.

(b) He proposed that the positive charge is spread over a sphere of the size of the atom (i.e. 10^{-8} cm radius) in which electrons are embedded to make the atom as whole neutral.

(c) This model could not explain the experimental results of Rutherfords α-particle scattering, therefore it was rejected.

6. RUTHERFORD'S MODEL

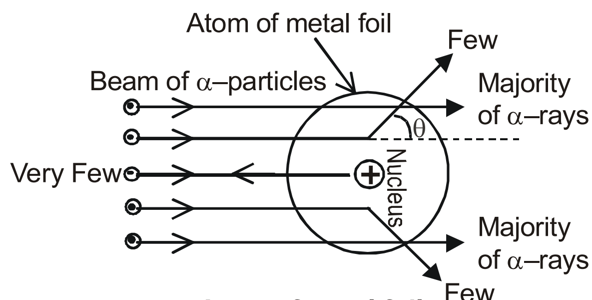
Rutherford carried out experiment on the bombardment of atoms by high speed positively charged α - particles emitted from radium and gave the following observations, which was based on his experiment.



(a) Most of the α - particles (nearly 99%) continued with their straight path.

(b) Some of the α - particles passed very close to the centre of the atom and deflected by small angles.

(c) Very few particles thrown back (180º) .



Atom of metal foil

6.1 Main features :

(a) Most of the α - particles were continued their straight path that means most of the space of the atom is empty.

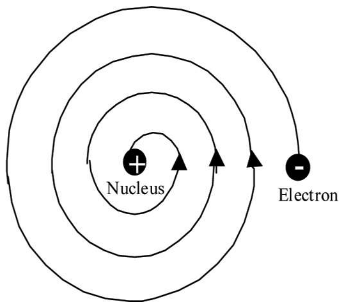
(b) The centre of an atom has a positively charged body called **nucleus** which repel positively charged α - particles and thus explained the scattering phenomenon.

(c) Whole mass of an atom is concentrated in its nucleus and very few throw back means the size of the nucleus is very small 10^{-13} cm. It showed that the nucleus is 10^{-5} times small in size as compared to the total size of atom.

(d) The size and volume of the nucleus is very small as compared to the total size and volume of atom.

(e) As atomic number increases, the angle of deflection (θ) increases.

6.2 Drawbacks of Rutherford's model :



(a) According to classical electromagnetic theory, when an electron moves around the nucleus under the influence of the attractive force, the electron loses its energy continuously and move closer and closer to the nucleus in a spiral path, the ultimate result will be that it will fall into the nucleus but it can't be possible because an atom is quite stable.

(b) If an electron loses energy continuously, the observed spectrum should be continuous but the actual observed spectrum consist of discontinuous well defined lines of definite frequencies.

Examples based on Rutherford's Experiment

Ex.5 Rutherford's scattering experiment is related to reveal structure of –

(A) Nucleus (B) Atom

(C) Electron (D) Neutron

Ans . (B)

Sol. To reveal structure of atom

Ex.6 When the atoms of gold sheet are bombarded with a beam of α-particles, only a few α-particles get deflected whereas most of them go straight undeflected. This is because –

(A) The force of attraction on the α-particles by the oppositely charged electron is not sufficient

(B) The nucleus occupies much smaller volume as compared to the volume of atom

(C) The force of repulsion on fast moving α-particles is very small

(D) The neutrons in the nucleus do not have any effect on α-particles.

Sol. It was the logical conclusion of his experiment.

7. MOSELEY'S EXPERIMENT

7.1 Atomic number (Z) :

The number of positive charge carried by the nucleus of an atom is termed as atomic no. (Z) or the number of protons in an atom of an element is equal to its atomic number. Since an atom is electrically neutral it contains an equal number of extra nuclear electrons. Thus –

Atomic No. = Number of unit positive charge in nucleus = Number of protons

= Number of electrons.

7.2 Mass number or Neucleon number (A) :

The mass number being the sum of the number of protons and neutrons in the nucleus, which is always a whole number.

$$A = P + n$$

or

$$A = Z + n$$

where :

A = Mass number

P = Number of protons

n = Number of neutrons

Z = Atomic number

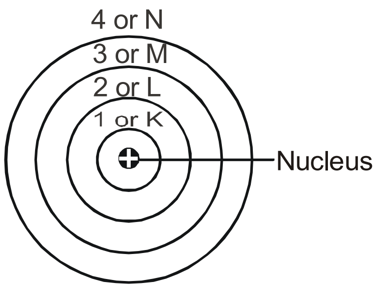
On the another side of that statement since mass of a proton or a neutron is not a whole number (on atomic weight scale), atomic weight is not necessarily a whole number.

For example : The isotopes of oxygen having mass number 17 and 18, have atomic weights equal to 17.00045 and 18.0037 respectively.

8. BOHR'S ATOMIC MODEL

Bohr developed atomic model for hydrogen and hydrogen like one electron species on the basis of Planck's quntum theory.

8.1 The important postulates of Bohr model of an atom :



(a) Electron revolves around the nucleus in a fixed circular orbit of definite energy.

(b) Electron revolves only in those orbits whose angular momentum (mvr) is an integral multiple of the factor h/2π (where 'h' is Planck's constant)

$$mvr = n\frac{h}{2\pi}$$

where :

m = mass of the electron

v = velocity of the electron

n = number of orbit in which electron revolves

i.e. n = 1, 2, 3 ....

r = radius of the orbit.

(c) As long as the electron occupy a definite energy level, it does not radiate out energy i.e. it does not lose or gain energy.

(d) The energy is emitted or absorbed only when the electron jumps from one energy level to another. If energy is supplied to an electron, It may jump higher energy level to the lower by the emission of energy. This higher energy level called excited state. Similarly in the reverse process it may absorb energy and jump from lower to higher energy level.

This amount of energy emitted or absorbed is given by the difference of the energies of the two energy levels concerned.

n = 2 ———— E_2 $\quad \Delta E = E_2 - E_1$

$\uparrow \qquad \downarrow$

n = 1 ———— E_1 $\quad$ Absorbed energy

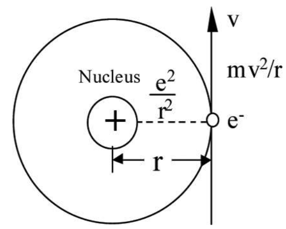
n = 2 ———— E_2 $\quad \Delta E = E_1 - E_2$

$\downarrow \qquad \downarrow$

n = 1 ———— E_1 $\quad$ Emitted energy

8.2 Mathematical term of Bohr's Postulates :

8.2.1 Calculation of the radius of the Bohr's orbit:



Suppose that an electron having mass 'm' and charge 'e' revolving around the nucleus of charge 'Ze' (Z is atomic number & e = charge) with a tangential / linear velocity of 'v'. Further consider that 'r' is the radius of the orbit in which electron is revolving.

According to **Coulomb's law**, the electrostatic force of attraction (F) between the moving electron and nucleus is -

$$F = \frac{KZe^2}{r^2}$$

Where : K = constant = $\frac{1}{4\pi\varepsilon_0}$ = 9×10^9 Nm^2/C^2

and the centrifugal force $\quad F = \frac{mv^2}{r}$

For the stable orbit of an electron both the forces are balanced,

i. e. at equilibrium $\quad \frac{mv^2}{r} = \frac{KZe^2}{r^2}$

then $\quad V^2 = \frac{KZe^2}{mr} \quad$ ...(1)

From the postulate of Bohr,

$$mvr = \frac{nh}{2\pi}$$

$$v = \frac{nh}{2\pi mr}$$

$$v^2 = \frac{n^2h^2}{4\pi^2m^2r^2} \quad \text{...(2)}$$

From equation (1) and (2) :

on solving, we will get

$$r = \frac{n^2h^2}{4\pi^2mKZe^2}$$

In C.G.S. $\boxed{\text{unit} K = 1,}$

$\therefore \quad \boxed{r = \frac{n^2h^2}{4\pi^2 Ze^2}}$

where ; h = 6.62×10^{-27} erg. sec.

m = 9.1×10^{-28} g

e = 1.6×10^{-19} C.

on putting the value or e, h, m

then $\boxed{r = 0.529 \times \frac{n^2}{Z} \text{Å}}$

$$\boxed{\begin{array}{l} r \propto n^2 \\ r \propto \frac{1}{Z} \end{array}}$$

Orbital frequency $f = \frac{v}{2\pi r}$

8.2.2 Calculation of velocity of an electron in Bohr's orbit :

Velocity of the revolving electron in n^{th} orbit is given by –

$$mvr = \frac{nh}{2\pi} \qquad v = \frac{nh}{2\pi mr}$$

To keep the value of 'r' on the equation (1)

then $$v = \frac{nh \times 4\pi^2 mZe^2}{2\pi mn^2h^2}$$

$$\boxed{v = \frac{2\pi Ze^2}{nh}}$$

on purtting the values of π, e^- and h

$$\boxed{v = 2.188 \times 10^8 \times \frac{Z}{n} \text{cm/sec}}$$

$$\boxed{\begin{array}{l} v \propto Z \\ v \propto \frac{1}{n} \end{array}}$$

8.2.3 Calculation of energy of an electron :

The total energy of an electron revolving in a particular orbit is -

T. E. = K. E. + P. E.

Where ;

P. E. = Potential energy,

K.E. = Kinetic energy,

T.E.= Total energy

The K.E. of an electron $= \frac{1}{2}mv^2$

and the P.E. of an electron $= -\frac{KZe^2}{r}$

Hence, $$T.E. = \frac{1}{2}mv^2 - \frac{KZe^2}{r}$$

We know that, $$\frac{mv^2}{r} = \frac{KZe^2}{r^2}$$

or $$mv^2 = \frac{KZe^2}{r}$$

substituting the value of mv^2 in the above equation : -

$$T.E. = \frac{KZe^2}{2r} - \frac{KZe^2}{r} = -\frac{KZe^2}{2r}$$

So, $$T.E. = -\frac{KZe^2}{2r}$$

In C.G.S. unit K = 1

$$\therefore \quad T.E. = -\frac{Ze^2}{2r}$$

Substituting the value of 'r' in the equation of T.E. .

Then,

$$E = -\frac{Ze^2}{2r} \times \frac{4\pi^2 Ze^2 m}{n^2h^2} = -\frac{2\pi^2 Z^2 e^4 m}{n^2h^2}$$

Thus, the total energy of an electron in n^{th} orbit is given by

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

Note : – The P.E. at the infinite = 0

The K.E. at the infinite = 0

8.2.4 Relation between P. E., K. E. & T. E. :

$$P.E. = -\frac{Ze^2}{r}, \quad K.E. = \frac{1}{2}\frac{Ze^2}{r},$$

$$T.E. = -\frac{1}{2}\frac{Ze^2}{r}$$

So, $$\frac{T.E.}{P.E.} = \frac{-\frac{1}{2}\frac{Ze^2}{r}}{-\frac{Ze^2}{r}} = \frac{1}{2}$$ Then

T. E. = ½ P. E. ...(1)

$$\frac{T.E.}{K.E.} = \frac{-\frac{1}{2}\frac{Ze^2}{r}}{\frac{1}{2}\frac{Ze^2}{r}}$$

Then T . E. = – K. E. ...(2)

$$\boxed{T.E. = \frac{P.E.}{2} = -K.E.} \quad \ldots(3)$$

(a) $T.E. = -13.6 \times \frac{Z^2}{n^2}$ eV / atom

(b) $T.E. = -21.8 \times 10^{-19} \times \frac{Z^2}{n^2}$ J / atom

(c) $T.E. = -21.8 \times 10^{-12} \times \frac{Z^2}{n^2}$ erg / atom

(d) $T.E. = -313.6 \times \frac{Z^2}{n^2}$ Kcal / mole

8.2.5 Conclusions from equation of energy :

(a) The negative sign of energy indicates that there is attraction between the negatively charged electron and positively charged nucleus.

(b) All the quantities of R.H.S. in the energy equation are constant for an element except 'n' which is an integer such as 1, 2, 3 etc. i. e. the energy of an electron is constant as long as the value of 'n' is kept constant.

(c) The energy of an electron is directly proportional to the square of 'n'.

8.3 Calculation of Rydberg Constant

Suppose that an electron transist from first energy level to second energy level. Then, the change of energy is given by

$$\Delta E = E_2 - E_1$$

$$hv = E_2 - E_1$$

$$hv = \left[\frac{-2\pi^2 mZ^2 e^4}{n_2^2 h^2}\right] - \left[\frac{-2\pi^2 mZ^2 e^4}{n_1^2 h^2}\right]$$

$$hv = \frac{2\pi^2 mZ^2 e^4}{n_1^2 h^2} - \frac{2\pi^2 mZ^2 e^4}{n_2^2 h^2} \qquad \because \ v = \frac{c}{\lambda}$$

$$\frac{hc}{\lambda} = \frac{2\pi^2 mZ^2 e^4}{h^2} \times \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$R_H = \frac{2\pi^2 me^4}{ch^3} \Rightarrow \text{Rydberg constant}$$

Then, $\bar{\mu} = \frac{1}{\lambda} = R_H Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$ where

$m = 9.1 \times 10^{-28}$ gram

by the theoretical value of R_H = 109737 / cm

$e = 4.8 \times 10^{-10}$ e.s.u.

by the practical value of R_H = 109677 / cm

$c = 3 \times 10^{10}$ cm/sec

by the calculative value of R_H = 109700 / cm

$h = 6.625 \times 10^{-27}$ erg-sec

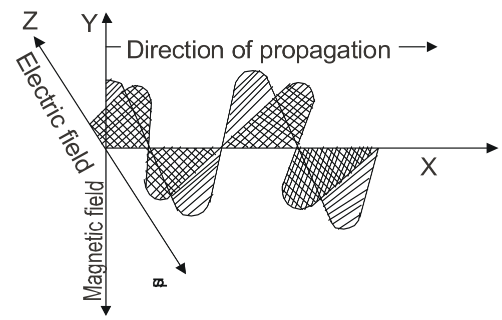
Rydberg constang for other atom $R = R_H \times Z^2$

9. ELECTROMAGNETIC RADIATIONS

Light and other forms of radiant energy propagate without any medium in the space in the form of waves. These waves can be produced by a charged body moving in a magnetic field or a magnet in an electric field.

e.g. α - rays, γ - rays, Cosmic rays, Ordinary light rays etc.

9.1 Characteristics of electromagnetic radiations:

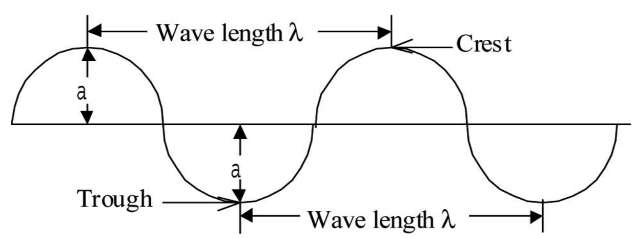


(a) All electromagnetic waves move or travel with the same velocity equal to that of light.

(b) They do not require any medium to propagate.

(c) These consist of electric and magnetic field that oscillate in the direction perpendicular to each other and to the direction in which the wave is propagate.

9.2 Some Important characteristics of electromagnetic waves :



(a) **Frequency (ν) :** It is defined as the no. of waves which pass through a given point in per sec. It's unit is expressed by cycle per second (cps) or Hertz (Hz).

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

Note : A cycle is said to be completed when a wave consisting of crest and trough passes through a point.

(b) Wavelength (λ) : The distance between two adjacent crest or troughs of the wave as shown in the fig. It is denoted by lambda (λ) a greek letter and unit is Angstrom (Å) or nanometer (nm).

1 Å = 10^{-10} m or 10^{-8} cm

1 nm = 10^{-9} m or 10^{-7} cm

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

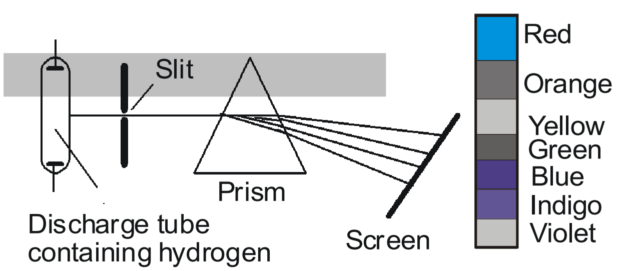
(c) Wave No. ($\bar{\nu}$) : It is defined as the number of wave per cm and it is equal to the inverse of wavelength. Its unit is cm^{-1}.

$$\bar{\nu} = \frac{1}{\lambda} \qquad \nu = \frac{c}{\lambda} = c\bar{\nu}$$

(d) Amplitude (a) : It denotes the height of the crest or depth of the trough of a wave. It determines the intensity of brightness of radiation.

(e) Velocity (v) : The distance traveled per sec by a wave called velocity of a wave. It is expressed by the unit of m/sec. or cm/sec.

10. SOLAR SPECTRUM



Line spectrum

(a) When sunlight is passed through a prism, It absorbs wavelength range of black colour radiation and other splits into a series of colour bands known as emission spectrum and black colour band which is known as absorption spectrum.

(b) The splitting of light into seven colours is called emission **Spectrum**.

(c) The characteristic range of wavelength of electromagnetic radiation situated in an increasing or decreasing order called electromagnetic spectrum.

Name	λ in Å	Origin
Radio waves	3×10^{9} to 3×10^{14}	by the Alternating current of high frequency
Microwaves	3×10^{6} to 3×10^{9}	by the generator of high quality
I.R.	$7600\times3\times10^{6}$	from the heated things
Visiblewave	3800 to 7600	
U.V. wave	150 to 3800	from the sun rays
X-rays	0.1 to 150	to put a metallic barrior in path of moving electron
γ–rays	0.01 to 0.1	by radio active disintegration
Cosmic rays	0 to 0.01	from the outer most part of sun

— λ decreases ⟶

— ν Increases ⟶

(d) Band spectrum is originated by molecules and linear spectrum is originated by atoms.

10.1 Difference between Emission and absorption spectra :

Emission spectrum	Absorption spectrum
1. It is obtained when radiation emitted by the excited substance which is analysed in a spectroscope	It is obtained when white light is passed through the substance either gases or in the form of solution.
2. This type of spectrum consist of bright coloured lines separated by dark spaces.	It is consist of dark lines on a colour back ground.

11. ATOMIC SPECTRA OR LINE SPECTRA

Atomic spectra is line spectra. So atomic spectrum is also called line spectrum. It is of two types

11.1 Emission spectrum :

A substance gets excited on heating at a very high temperature or by giving energy and radiations are emitted. These radiations when analysed with the help of spectroscope, spectral lines are obtained. A substance may be excited as follows -

(a) By heating at a higher temperature.

(b) By passing electric current at a very low pressure in a discharge tube filled with gas.

(c) By passing electric current into metallic filament.

Emission spectra is of two types -

(i) Continuous spectrum

(ii) Line spectrum

(i) **Continuous spectrum :** When sunlight is passed through a prism, it gets dispersed into continuous bands of different colours. If the light of an incandescent object is resolved through prism or spectroscope, it also gives continuous spectrum of colours.

(ii) **Line spectrum :** If the radiations obtained by the excitation of a substance are analysed with the help of a spectroscope a series of thin bright lines of specific colours are obtained. There is dark space in between two consecutive lines. This type of spectrum is called line spectrum or atomic spectrum. For example on heating sodium chloride or any other salt of sodium in Bunsen flame bright yellow light is emitted. The emitted light when viewed through a spectroscope two isolated yellow lines separated by dark space are obtained. The wave lengths of these lines are 5890Å and 5896Å.

If an electron from nth excited state comes to various energy states, the maximum spectral lines obtained will be -

$$= \frac{n(n-1)}{2}$$

11.2 Absorption spectrum

When the white light of an incandescent substance is passed through any other substance, this substance absorbs the radiations of certain wavelength from the white light. On analysing the transmitted light we obtain a spectrum in which dark lines of specific wave lengths are observed. These lines constitute the absorption spectrum. The wave length of the dark lines correspond to the wavelength of light absorbed.

12. HYDROGEN SPECTRUM

(a) Hydrogen spectrum is an example of line emission spectrum or atomic emission spectrum.

(b) When an electric discharge is passed through hydrogen gas at low pressure, a bluish light is emitted.

(c) This light shows discontinuous line spectrum of several isolated sharp lines through prism.

(d) All thes lines of H-spectrum have following six series

Spectral series	Wavelength region
Lyman	U.V.
Balmer	Visible
Paschen	IR
Brackett	IR
Pfund	IR
Humphrey	Far I.R.

These spectral series were named by the name of scientist who discovered them

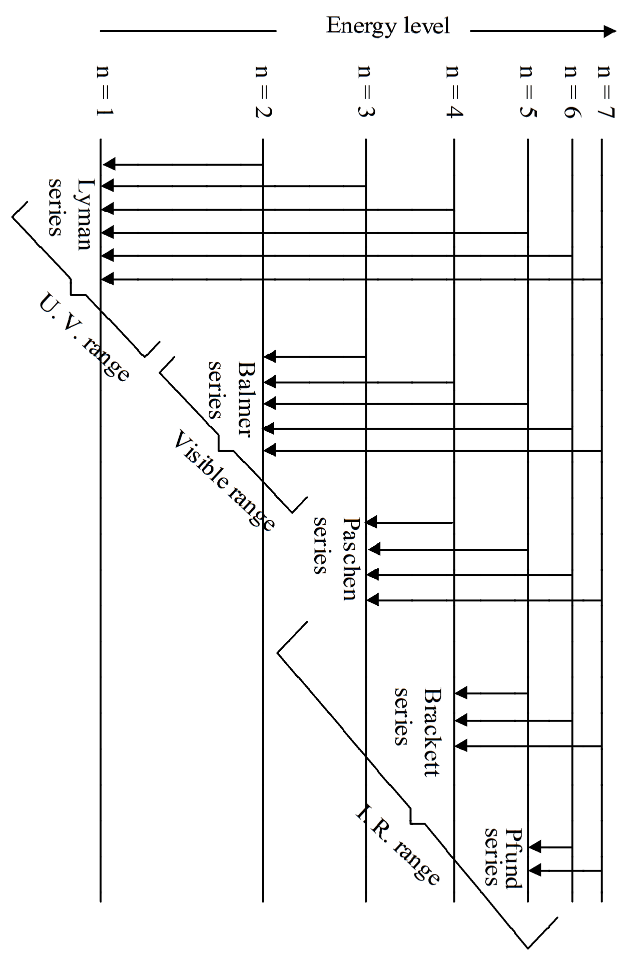
(e) To evaluate wavelength of various H-lines Ritz introduced the following expression

$$\bar{\nu} = \frac{1}{\lambda} = \frac{\nu}{C} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

Where R is a universal constant known as Rydberg's constant its value is 109, 678cm^{-1}.

(f) Although H - atom consists only one electron yet it's spectra consist of many spectral lines.

$$\frac{1}{\lambda} = \bar{\nu} = R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$



12.1 Lyman Series

(a) It is a first series of spectral series of H.

(b) It was found out in ultraviolet region in 1898 by ***Lyman***.

(c) It's value of n_1 = 1 and n_2 = 2, 3, 4 where 'n_1' is ground state and 'n_2' is called excited state of electron present in a H - atom.

(d) If the electron goes to n_1 = 1

to n_2 = 2 — first Lyman series

If the electron goes to n_1 = 1

to n_2 = 3 — Second Lyman series

If the electron goes to n_1 = 1

to n_2 = 4 — third Lyman series ----- so on.

(e) $\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{n_2^{\,2}}\right]$ where n_2 > 1 always.

(f) The wavelength of marginal line = $\frac{n_1^{\,2}}{R_H}$ for all series. So, for lyman series = $\frac{1}{R_H}$

12.2 Balmer series :

(a) It is the second series of H-spectral series.

(b) It was found out in 1892 in visible region by **Balmer**.

(c) Balmer series was found out before all series. Because it was found to be in visible region.

(d) It's value of n_1 = 2 and n_2 = 3, 4, 5 .... where n_1 is ground state and n_2 is excited state.

(e) If the electron goes to n_1 = 2

to n_2 = 3 —First Balmer series

If the electron goes to n_1 = 2

to n_2 = 4 — Second Balmer series

If the electron goes to n_1 = 2

to n_2 = 5 — third Balmer series .... so on

(f) The wavelength of marginal line of Balmer series = $\frac{n_1^{\,2}}{R_H} = \frac{2^2}{R_H} = \frac{4}{R_H}$

(g) $\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n_2^{\,2}}\right)$ where n_2 > 2 always

12.3 Paschen series :

(a) It is the third series of H - spectrum.

(b) It was found out in infra red region by ***Paschen***.

(c) It's value of n_1 = 3 and n_2 = 4, 5, 6 .... where n_1 is ground state and n_2 is excited state.

(d) If the electron goes to n_1 = 3

to n_2 = 4 — First paschen series

If the electron goes to n_1 = 3

to n_2 = 5 — second paschen series

If the electron goes to n_1 = 3

to n_2 = 6 — third paschen series ------ so on.

(e) The wavelength of marginal line of paschen series = $\frac{n_1^{\,2}}{R_H} = \frac{3^2}{R_H} = \frac{9}{R_H}$.

(f) $\frac{1}{\lambda} = R_H\left[\frac{1}{3^2} - \frac{1}{n_2^{\,2}}\right]$ where n_2 > 3 always.

12.4 Brackett series :

(a) It is fourth series of H - spectrum.

(b) It was found out in infra red region by **Brackett**.

(c) It's value of n_1 = 4 and n_2 = 5, 6, 7 .... where n_1 is ground state and n_2 is excited state.

(d) If the electron goes to n_1 = 4

to n_2 = 5 — first brackett series

If the electron goes to n_1 = 4

to n_2 = 6 — second brackett series

If the electron goes to n_1 = 4

to n_2 = 7 — third brackett series ------ so on.

(e) The wavelength of marginal line of brackett series = $\frac{n_1^{\,2}}{R_H} = \frac{4^2}{R_H} = \frac{16}{R_H}$

(f) $\frac{1}{\lambda} = R_H\left[\frac{1}{4^2} - \frac{1}{n_2^{\,2}}\right]$ Where n_2 > 4 always.

12.5 Pfund series :

(a) It is fifth series of H - spectrum.

(b) It was found out in infra red region by Pfund.

(c) It's value of $n_1 = 5$ and $n_2 = 6, 7, 8$ .... where n_1 is ground state and n_2 is excited state.

(d) If the electron goes to $n_1 = 5$

to $n_2 = 6$ — first Pfund series

If the electron goes to $n_1 = 5$

to $n_2 = 7$ — second Pfund series

If the electron goes to $n_1 = 5$

to $n_2 = 8$ — third Pfund series -----so on.

(e) The wavelength of marginal line of Pfund series $= \dfrac{n_1^2}{R_H} = \dfrac{5^2}{R_H} = \dfrac{25}{R_H}$

(f) $\dfrac{1}{\lambda} = R_H\left[\dfrac{1}{5^2} - \dfrac{1}{n_2^2}\right]$ where $n_2 > 5$ always.

12.6 Humfrey series :

(a) It is the sixth series of H - spectrum.

(b) It was found out in infra-red region by ***Humfrey***.

(c) It's value of $n_1 = 6$ and $n_2 = 7, 8, 9$ --------- where n_1 is ground state of electron and n_2 is excited state.

(d) If the electron goes to $n_1 = 6$

to $n_2 = 7$ — first Humfri series

If the electron goes to $n_1 = 6$

to $n_2 = 8$ — second Humfri series

If the electron goes to $n_1 = 6$

to $n_2 = 9$ — third Humfri series ... so on.

(e) The wavelength of marginal line of Humfri series $= \dfrac{n_1^2}{R_H} = \dfrac{6^2}{R_H} = \dfrac{36}{R_H}$

(f) $\dfrac{1}{\lambda} = R_H\left[\dfrac{1}{6^2} - \dfrac{1}{n_2^2}\right]$ where $n_2 > 6$.

13. CONCEPT OF QUANTIZATION

(a) E.M. wave theory successfully explains about reflection, refraction, diffraction, etc. but it fails to explain black body radiations and photo electric effect

(b) To explain all these things Max planck gave a new revolutionary theory in 1901, called as quantum theory of radiation.

(c) According to this theory, a hot body emits radiant energy not continuously but discontinuously in the form of small packets of energy called quantum.

(d) In case of light, the quantum of energy is often called photon.

(e) The amount of energy associated with a quantum radiation is proportional to the frequency of light

$E \propto \nu$ or $E = h\nu$

where the proportionality constant, h is a universal constant known as Planck's constant. Its value is 6.63×10^{-34} J-sec

(f) The total amount of energy emitted or absorbed by a body will be some whole number multiple of quantum, i.e.

$E = nh\nu$

where n is an integer such as 1, 2, 3 ......

Examples based on E.M. Radiation and Spectrum

Ex.7 The wavelengths of two photons are 2000Å and 4000Å respectively. What is the ratio of their energies-

(A) 1/4 (B) 4
(C) 1/2 (D) 2 **Ans . (D)**

Sol. $E_1 = h.\dfrac{c}{\lambda_1}$

$E_2 = h.\dfrac{c}{\lambda_2}$

$\dfrac{E_1}{E_2} = \dfrac{hc}{\lambda_1} \times \dfrac{\lambda_2}{hc} = \dfrac{\lambda_2}{\lambda_1} = \dfrac{4000}{2000} = 2$

Ex.8 There are three energy levels in an atom. How many spectral lines are possible in its emission spectra-

(A) One (B) Two
(C) Three (D) Four **Ans . (C)**

Sol. Number of spectral lines

$= \dfrac{n(n-1)}{2} = \dfrac{3(3-1)}{2} = 3$

Ex.9 Which of the following transitions will emit the photons of highest frequency in hydrogen atom -

(A) From n = 1 to n = 2
(B) From n = 2 to n = 1
(C) From n = 2 to n = 6
(D) From n = 6 to n = 2 **Ans . (B)**

Sol. The emission of photon is due to the transition of electrons from higher to lower energy levels. So the answer may be (2) or (4). From Planck's equation.

$\nu \propto E$

i.e. The frequency of emitted photon is directly proportional to the difference of energies of two energy levels.

Energy of n = 1 for H-atom

$E_1 = -13.6$ eV

Energy of n = 2 for H-atom

$E_2 = -\frac{13.6}{4}$ eV

Energy of n = 6 for H-atom

$E_6 = -\frac{13.6}{36}$ eV

So $E_2 - E_1 = 13.6 - \frac{13.6}{4} = 13.6 \times \frac{3}{4}$

$E_6 - E_2 = \frac{13.6}{4} - \frac{13.6}{36}$

$= 13.6\left(\frac{1}{4} - \frac{1}{36}\right) = 13.6 \times \frac{2}{9}$

$E_2 - E_1 > E_6 - E_2$

Ex.10 Which type of radiation is not emitted by the electronic structure of atoms –

(1) Ultraviolet light (2) X-rays
(3) Visible light (4) γ-rays **Ans . (4)**

Sol. γ-rays emission occurs due to radioactive change, a nuclear phenomenon.

14. FAILURES/LIMITATIONS OF BOHR'S THEORY

(a) He could not explain the line spectra of atoms containing more than one electron.

(b) He also could not explain the presence of multiple spectral lines.

(c) He was unable to explain the splitting of spectral lines in magnetic field (*Zeeman effect*) and in electric field (*Stark effect*).

(d) No one conclusion was given for the principle of quantisation of angular momentum.

(e) He was unable to explain the ***de-Broglie's*** concept of dual nature of matter.

(f) He could not explain ***Heisenberg's*** uncertainty principle.

15. SOMMERFELD'S CONCEPT

Extension of Bohr's theory

(a) Sommerfeld in 1915, introduced a new atomic model to explain fine spectrum of hydrogen atom

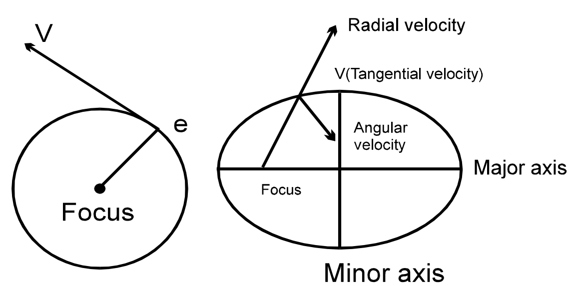
(b) He proposed that the moving electron might describe elliptical orbits in addition to circular, orbits and the nucleus is situated at one of the foci.

(c) During motion on a circle, only the angle of revolution changes while the distance from the nucleus remains the same but in elliptical motion both the angle of revolution and the distance of the electron from the nucleus change.

(d) The distance from the nucleus is termed as radius vector and the angle of revolution is known as azimuthal angle.

(e) The tangential velocity of the electron at a particular instant can be resolved into two components. One along the radius vector called radial velocity and the other perpendicular to the radius vector called transverse or angular velocity.

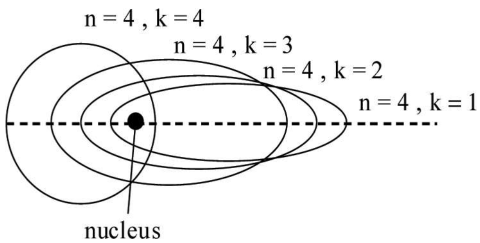
(f) These two velocities give rise to radial momentum and angular or azimuthal momentum.



(g) Sommerfeld proposed that both the momenta must be integral multiples,

radial momentum $= n_r \frac{h}{2\pi}$, Azimuthal momentum $= n_\phi \frac{h}{2\pi}$

The brief explanations of Sommerfeld's model are as follows



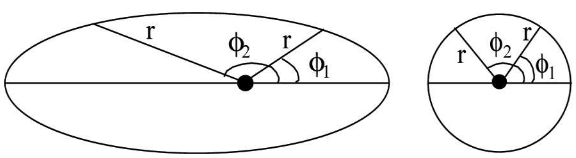
(i) Sommerfield model gives introduction of elliptical orbitals.

(ii) Bohr's circular orbit has considered to be a special case of an elliptical orbit, in which the length of major and minor axis is same.

(iii) Only one co-ordinate angle of revolution ϕ is variable in circular orbit but in elliptical, vector radius 'r' is also variable.

(iv) Introduction of azimuthal quantum number in addition to the principal quantum number.

Angular momentum of an electron moving around elliptical orbit is the sum of two vector terms as follows-



ϕ = change	ϕ = change
r = change	r = constant

$$P = P_r + P_\phi$$

$$= \frac{n_r h}{2\pi} + \frac{n_\phi h}{2\pi}$$

$n = n_r + n_\phi$

Where, n = Principal quantum number

n_r = Radial quantum number

= 1, 2, 3, ... ∞

n_ϕ = Azimuthal quantum number

= 1, 2, 3, ... n

(v) Introduction of sub shells or sub energy levels within principal energy levels. These were considered to had different energies.

(vi) Sub shell are termed as s, p, d, f (sharp, principal, diffused, fundamental).

(vii) These subshells were considered to had capacity of 2, 6, 10, 14 electrons respectively.

(viii) Energies of these subshells follow the order s < p < d < f.

(ix) The relation between principal (n) and azimuthal (ℓ) quantum number is -

$$\frac{n}{\ell} = \frac{\text{length of major axis}}{\text{length of minor axis}}$$

(x) The subshells (s, p, d, f) in principal energy level have very slight difference in their energies. The spectral lines corresponding to the transition of electron in their sublevels have a fine structure. This was major achievement of this extension.

(xi) First energy level (K or n = 1) of Bohr contains only one subshell (s). Second (L or n = 2) contains two subshell (s & p). Third (M or n = 3) contains three subshell (s, p and d). Fourth (N or n = 4) contains four subshell (s, p, d, f).

(According to Sommerfield, n^{th} shell of Bohr has n subshell in which one circular & (n – 1) elliptical subshells are present.)

The necessity of a modification of the Bohr theory stemmed from the spectral observation, under high resolution of the different hydrogen lines.

Extending Bohr's ideas sommerfeld suggested that

If a and b are semimajor and semiminor axes.

It follows $\frac{b}{a} = \frac{k}{k + n_r} = \frac{k}{n}$

Examples based on Sommerfeld's Model

Ex.11 To give designation to an orbital, we need -

(A) Principal and azimuthal quantum number

(B) Principal and magnetic quantum number

(C) Azimuthal and magnetic quantum number

(D) Principal, azimuthal and magnetic quantum numbers

Ans. (D)

Sol. The correct answer is **(D)**

Ex.12 The elliptical orbits of elctron in the atom were proposed by -

(A) Thomson (B) Bohr

(C) Sommerfeld (D) De Broglie

Ans. (C)

Sol. Follows sommerfeld concept.

16. WAVE MECHANICAL MODEL OF ATOM

16.1 Dual nature of electron

(a) Einstein had suggested that light can behave as a wave as well as like a particle i.e. it has dual character

(b) In 1924, de-Broglie proposed that an electron, behaves both as a material particle and as a wave.

(c) This proposed a new theory wave mechanical theory of matter. According to this theory, the electrons protons and even atom when in motion possess wave properties

(d) According to de-Broglie, the wavelength associated with a particle of mass m, moving with velocity v is given by the relation,

$$\lambda = \frac{h}{mv}$$

where h is Planck's constant.

(e) This can be derived as follows according to Planck's equation

$$E = h\nu = \frac{h \cdot c}{\lambda}$$

Energy of photon on the basis of Enstein's mass energy relationship

$$E = mc^2$$

Equating both $\frac{hc}{\lambda} = mc^2$ or $\lambda = \frac{h}{mc}$

Which is the same of de Broglie relation.

(f) This was experimentally verified by Davisson and Germer by observing diffraction effects with an electron beam.

Let the electron is accelerated with a potential of V than the kinetic energy is

$$\frac{1}{2}mv^2 = eV$$

$$m^2v^2 = 2eVm$$

$$mv = \sqrt{2eVm} = p$$

$$\lambda = \frac{h}{\sqrt{2eVm}}$$

(g) If we associate Bohr's theory with De-Broglie Equation we find that the wavelength of an electron, moving in Bohr's orbit is related with its circumference through a whole number multiple

$$2\pi r = n\lambda$$

or $$\lambda = \frac{2\pi r}{n}$$

From de-Broglie equation

$$\lambda = \frac{h}{mv}$$

Thus, $\frac{h}{mv} = \frac{2\pi r}{n}$ or $mvr = \frac{nh}{2\pi}$

Examples based on Dual Nature of Electron

Ex.13 If the Planck's constant h = 6.6 x 10^{-34} Js, the de-Broglie wavelength of a particle having momentum of 3.3 x 10^{-24}kg m s^{-1} will be -

(A) 0.002 Å (B) 0.02 Å

(C) 0.2 Å (D) 2Å **Ans.(D)**

Sol. $\lambda = \frac{h}{mv}$

Ex.14 K.E. of the electron is 4.55 x 10^{-25} J. Its de Broglie wave length is -

(A) 4700 Å (B) 8300Å

(C) 7200Å (D) 7400Å

Ans.(C)

Sol. $\lambda = \frac{h}{\sqrt{2mKE}}$

Ex.15 For particles having same kinetic energy, the de Broglie wavelength is -

(A) Directly proportional to its velocity

(B) Inversely proportional to its velocity

(C) Independent of velocity and mass

(D) Unpredictable. **Ans.(A)**

Sol. $\lambda = \frac{h}{mv}$ KE = $\frac{1}{2}$ mv^2

$$mv = \frac{2KE}{v}$$

$$\therefore \lambda = \frac{h}{\frac{2KE}{v}}$$

$$\lambda = h\left(\frac{v}{2KE}\right)$$

Ex.16 Velocity of helium atom at 300K is 2.40 × 10^2 meter per sec. What is its wave length. (mass number of helium is 4) -

(A) 0.416 nm (B) 0.83 nm

(C) 803 Å (D) 8000Å **Ans.(A)**

Sol. $\lambda = \frac{h}{mv}$

mass of helium = $\frac{4.0 \times 10^{-3}}{6.023 \times 10^{23}}$ kg. and

h = 6.62 x 10^{-34}

$$\lambda = 6.62 \times 10^{-34} \times \frac{6.023 \times 10^{23}}{4.0 \times 10^{-3}} \times \frac{1}{2.4 \times 10^2}$$

= 0.416 x 10^{-9} meter

λ = 0.416 nm

16.2 Heisenberg's uncertainty principle

(a) While treating e^- as a wave it is not possible to ascertain simultaneously the exact position and velocity of the e^- more precisely at a given instant since the wave is extending throughout a region of space

(b) As the photons of longer wavelengths are less energetic, hence they have less momentum and cannot be located exactly

(c) In 1927, Werner Heisenberg presented a principle known as Heisenberg's uncertainty principle

(d) According to this principle it is impossible to measure simultaneously the exact position and exact momentum of a body as small as an electron.

(e) If uncertainty of measurement of position is Δx uncertainty of momentum is Δp or $m\Delta v$. then according to Heisenberg

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$$

or $$\Delta x \cdot m\Delta v \geq \frac{h}{4\pi}$$

where h is planck's constant

(f) For other canonical conjugates of motion the equation for Heisenberg's uncertainty principle may be given as

$$\Delta E\ \Delta t \geq \frac{h}{4\pi} \text{ (for energy and time)}$$

Examples based on Heisenberg's Uncertainty Principle

Ex.17 If uncertainty in position and momentum are equal, the uncertainty in velocity is -

(A) $\sqrt{h/2\pi}$ (B) $\frac{1}{2m}\sqrt{h/\pi}$

(C) $\sqrt{h/\pi}$ (D) None **Ans.(B)**

Sol. $\Delta x \cdot m\Delta v = \frac{h}{4\pi}$ or $(m\Delta v)^2 = \frac{h}{4\pi}$

$m\Delta v = \sqrt{\frac{h}{4\pi}}$ or

$$\Delta v = \sqrt{\frac{h}{4\pi m^2}} = \frac{1}{2m}\sqrt{\frac{h}{\pi}}$$

Ex.18 The uncertainty in position and velocity of a particle are 10^{-10} m and 5.27×10^{-24} ms^{-1} respectively. Calculate the mass of the particle is ($h = 6.625 \times 10^{-34}$ J-s) –

(A) 0.099 kg (B) 0.99 g

(C) 0.92 kg (D) None **Ans.(A)**

Sol. According to Heisenberg's uncertainty principle,

$$\Delta x \cdot m\ \Delta v = \frac{h}{4\pi}$$

or $$m = \frac{h}{4\pi\,\Delta x.\ \Delta v}$$

$$= \frac{6.625 \times 10^{-34}}{4 \times 3.143 \times 10^{-10} \times 5.27 \times 10^{-24}}$$

$= 0.099$ kg.

16.3 Schrodinger wave equation

(a) Wave mechanical model of an atom was developed on the basis of dual nature of electron by Erwin schrodinger in 1926.

(b) In this model electron is described as a three-dimensional wave in the electric field of a positively charged nucleus.

(c) This approach is also called probability approach. The probability of finding an electron at any point around the nucleus can be calculated with the help of schrodinger wave equation. The schrodinger wave equation is

$$\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

where x, y and z are three space coordinates

m is the mass of electron

h is planck's constant

E is total energy

V is potential energy of e^-

ψ is the amplitude of wave also called wave function

(d) As we know that the variable quantity in a wave is the amplitude. Similarly we take wave function ψ in case of de-broglie waves

(e) In most of the cases, this wave function is a complex quantity and hence cannot be measured experimentally

(f) By performing proper mathematical operations on the wave function (ψ) information regarding position, momentum kinetic and potential energy etc. of the particle can be obtained.

(g) The most important property of ψ is that it gives a measure of the probability of finding the electron at a given position around the nucleus.

(h) The quantity ψ^2 gives the probability of finding an electron in a unit volume and is called probability density. This definition of probability is in agreement with the uncertainty principle as one cannot talk about the precise position of subatomic particles.

Examples based on Wave Mechanical Model of Atom

Ex.19 The wave-mechanical model of atom is based upon -

(A) De Broglie concept of dual character of matter
(B) Heisenberg's uncertainty principle
(C) Schrodinger wave equation
(D) All the above three **Ans. (D)**

Sol. Wave mechanical model of atom is based upon all the above.

Ex.20 Electron density in an orbital is correctly described by -

(A) ψ^2 (B) ψ
(C) $|\psi^2|\psi$ (D) None **Ans. (A)**

Sol. The correct representation is described by ψ^2.

17. QUANTUM NUMBERS

(a) The measurement scale by which the orbitals are distinguished, can be represented by sets of numbers called as quantum number.

(b) It is a very important number to specify and display to complete information about size, shape and orientation of the orbital. These are principle, azimuthal and magnetic quantum number, which follows directly from solution of schrodinger wave equation.

(c) Except of these quantum numbers, one additional quantum number designated as spin quantum number, which specify the spin of electron in an orbital.

(d) Each orbital in an atom is specified by a set of three quantum numbers and each electron is designated by a set of four quantum numbers.

These quantum numbers are as follows :

17.1 Principal quantum number (n) :

(a) It was proposed by Bohr and denoted by 'n'.

(b) It determines the average distance between electron and nucleus, means it is denoted the size of atom.

(c) It determine the energy of the electron in an orbit where electron is present.

(d) The maximum number of an electron in an orbit represented by this quantum number as $2n^2$.

(e) It gives the information of orbit K, L, M, N, ...

(f) The value of energy increases with the increasing value of n.

(g) It represents the major energy shell from which the electron belongs.

(h) An orbital momentum of any orbit = $\frac{nh}{2\pi}$

17.2 Azimuthal quantum number or angular quantum number (ℓ) -

(a) It was proposed by sommerfeld and denoted by 'ℓ'.

(b) It determines the number of subshells or sublevels to which the electron belongs.

(c) It tells about the shape of subshells.

(d) It also expresses the energies of subshells s < p < d < f (Increasing energy).

(e) The value of ℓ = (n – 1) always where 'n' is the number of principle shell.

(f)

Value of ℓ =	0	1	2	3 ----(n–1)
Name of subshell =	s	p	d	f
Shape of subshell =	spherical	Dumbbell	Double dumbbell	Complex dumbbell

(g) It represent the orbital angular momentum, which is equal to $\frac{h}{2\pi}\sqrt{\ell(\ell+1)}$.

(h) The number of electrons in subshell = $2(2\ell + 1)$.

(i) For a given value of 'n' the total value of 'ℓ' is always equal to the value of 'n' .

(j) The energy of any electron is depend on the value of n & ℓ because total energy = (n + ℓ). The electron enters in that sub orbit whose (n + ℓ) value or the value of energy is less.

17.3 Magnetic quantum number (m) :

(a) It was proposed by **Linde** and denoted by 'm'.

(b) It gives the number of permitted orientation of subshells.

(c) The value of m varies from – ℓ to + ℓ through zero.

(d) It tells about the splitting of spectral lines in the magnetic field i.e. this quantum number proved the Zeeman effect.

(e) For a given value of 'n' the total value of 'm' is equal to n^2.

(f) For a given value of 'ℓ' the total value of 'm' is equal to $(2\ell + 1)$.

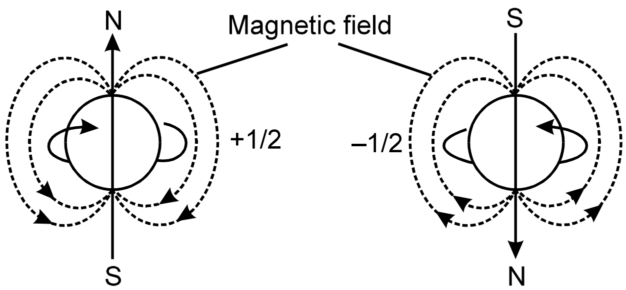
(g) *Degenerate orbitals* - Orbitals having the same energy are known as degenerate orbitals.

e.g. for P subshell P_x P_y P_z

(h) The number of degenerate orbitals of s subshell = 0.

17.4 Spin quantum number (s) :

(a) It was proposed by ***Goldshmidt*** & ***Uhlenbeck*** and denoted by the symbol of 'S'.



(b) The value of 's' is + ½ & – ½, Which is signified the spin or rotation or direction of electron on it's axis during the movement.

(c) The spin may be clockwise & anticlockwise.

(d) It represents the value of spin angular momentum is equal to $\frac{h}{2\pi}\sqrt{s(s+1)}$.

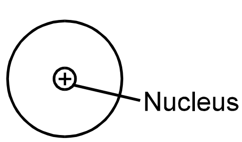
(e) Maximum spin of an atom = ½ × number of unpaired electron.

18. SHAPE OF ORBITALS

Orbital : Orbital is the three dimensional region around the nucleus where there is a maximum tendency of finding an electron of definite energy

18.1 Shape of orbitals on the basis of quantum number

18.1.1 Shape of 's' orbital :

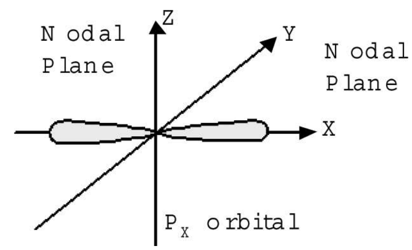


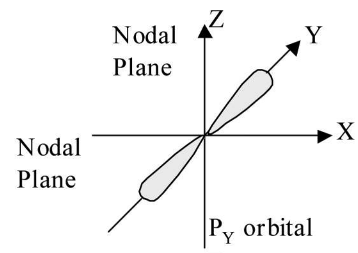
(a) For 's' orbital ℓ = 0 & m = 0 so 's' orbital have only one unidirectional orientation i.e. the probability of finding the electron is same in all directions.

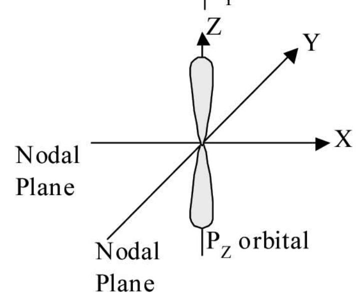
(b) The size and energy of 's' orbital with increasing 'n' will be 1s < 2s < 3s < 4s.

(c) It does not consist any directional property.

18.1.2 Shape of 'p' orbitals:







(a) For 'p' orbital ℓ = 1 & m = +1, 0, – 1 means there are three 'p' orbitals, which is symbolised as p_x, p_y, p_z.

(b) Shape of 'p' orbital is dumbbell in which the two lobes on opposite side separated by the nodal plane.

(c) p-orbital has directional properties.

18.1.3 Shape of d-orbital:

(a) For the 'd' orbital ℓ = 2 then the values of 'm' are –2, –1, 0, +1, +2. It shows that the 'd' orbitals has five orbitals as d_{xy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$, d_{z^2} .

(b) Each 'd' orbital identical in shape, size and energy.

(c) The 'd' orbital is bidumb-belled.

(d) It has directional properties.

Example based on Quantum Numbers and Shape of Orbitals

Ex.21 The maximum number of atomic orbitals associated with a principal quantum number 5 is –

(A) 9 (B) 12
(C) 16 (D) 25 **Ans.(D)**

Sol. The number of orbitals in a principle shell is $n^2 = 5^2 = 25$.

Ex.22 Beryllium's fourth electron will have the four quantum numbers –:

	n	ℓ	m	s	
(A)	1	0	0	1/2	
(B)	1	1	1	1/2	
(C)	2	0	0	–1/2	
(D)	2	1	0	+1/2	**Ans.(C)**

Sol. It is $2s^1$

19. ELECTRONIC CONFIGURATION PRINCIPLES

The distribution of electrons in different orbitals is known as electronic configuration of the atoms.

Filling up of orbitals in the ground state of atom is governed by the following rules :

19.1 Aufbau Principle :

(a) It is a German ward, meaning 'building up'

(b) According to this principle, "In the ground state, the atomic orbitals are filled in order of increasing energies". i.e. in the ground state the electrons occupy the lowest orbitals available to them.

(c) The sequence of filling of e^- we have already discussed in previous article

(d) In fact the energy of an orbital is determined by the quantum number n and ℓ with the help of (n + ℓ) rule or Bohr Bury rule

(e) According to this rule

(i) Lower the value of n + l, lower is the energy of the orbital and such an orbital will be filled up first

(ii) When two orbitals have same value of (n + ℓ) the orbital having lower value of "n" has lower energy and such an orbital will be filled up first.

19.2 Pauli's Exclusion Principle :

(a) According to this principle, " No two electrons in an atom can have all the four quantum numbers n, ℓ, m and s identical.

(b) In an atom, any two electrons may have three quantum numbers identical but fourth quantum number must be different.

(c) Since this principle excludes certain possible combinations of quantum numbers for any two electrons in an atom, it was given the name exclusion principle,

Its results are as follows

(i) The maximum capacity of a main energy shell is equal to $2n^2$ electron

(ii) The maximum capacity of a subshell is equal to $2(2\ell + 1)$ electrons

(iii) Number of sub-shells in a main energy shell is equal to the value of n

(iv) Number of orbitals in a main energy shell is equal to n^2

(v) one orbital cannot have more than two electrons

(d) According to this principle an orbital can accomod at the most two electrons with their spins opposite to each other.

(e) It means that an orbital can have 0, 1, or 2 electron

(f) If an orbital has two electrons they must be of opposite spin

correct Incorrect

19.3 Hund's Rule of Maximum Multiplicity :

(a) This rule governs the filling up of degenerate orbitals of the same sub-shell

(b) Accordint to this rule "Electron filling will not take place in orbitals of same energy unitill all the available orbital of a given subshell contain one electron each with parallel spin."

(c) This implies that electron pairing begins with fourth, sixth and eighth electron in p, d and f orbitals of the same sub-shell respectively.

(d) The reason behind this rule is related to repulsion between identical charged electron present in the same orbital

(e) They can minimise the repulsive forces between them serves by occupying different orbitals.

(f) Moreover, according to this principle, the e^- entering the different orbitals of subshell have parallel spins. This keeps them farther apart and lowers the energy through electron exchange or resonance.

(g) The term maximum multiplicity means that the total spin of unpaired e^- is maximum in case of correct filling of orbitals as per this rule.

19.4 (n + ℓ) Rule

This rule states that electrons are filled in orbitals according to their n + ℓ values. Electrons are filled in increasing order of their (n + ℓ) values. When (n + ℓ) is same for sub energy levels, the electrons first occupy the sublevels with lowest "n" value.

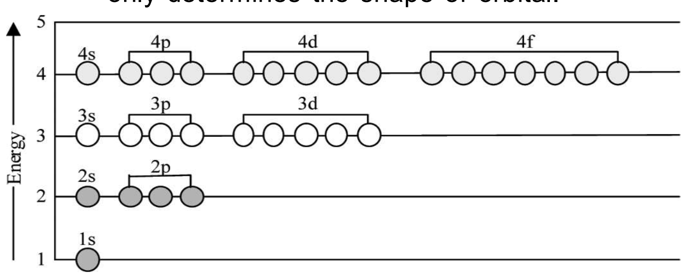
Thus, order of filling up of orbitals is as follows:

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d

20. ENERGY LEVEL DIAGRAM

(a) The representation of relative energy levels of various atomic orbital is made in the terms of energy level diagrams.

(b) **One electron system :** In this system electron is in $1s^2$ level and all orbital of same principal quantum number have same energy, which is independent of (ℓ). In this system ℓ only determines the shape of orbital.



(c) **Multiple electron system :** The energy levels of such system not only depend upon the nuclear charge but also upon the another electron present in them -

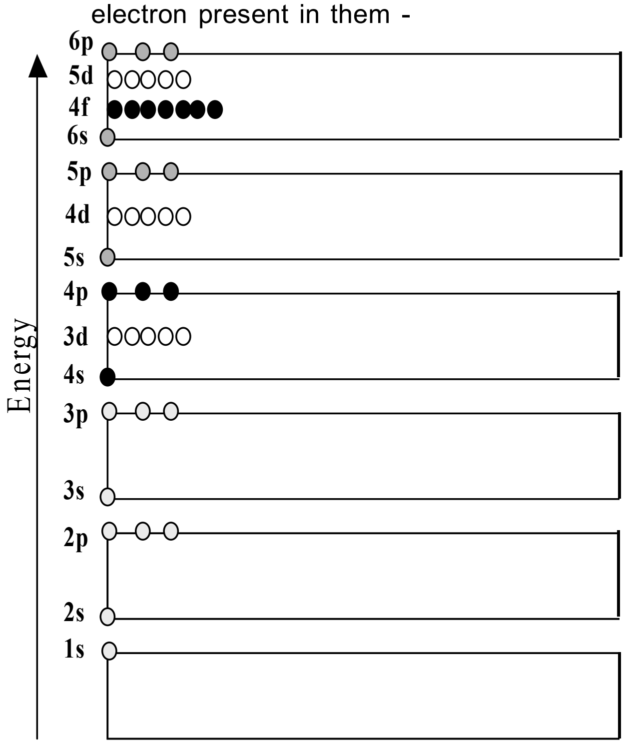


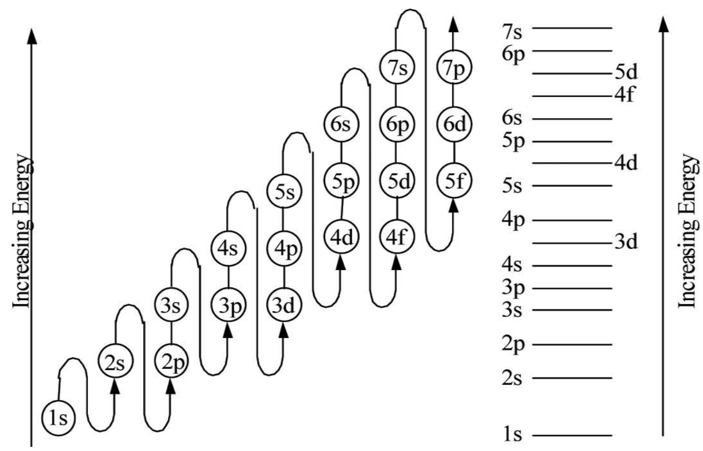
Diagram of multielectron atoms reveals the following points

(a) As the distance of the shell increases from the nucleus, the energy level increases. For example energy level of 2 > 1.

(b) The different sub shells have different energy levels who have possess definite energy levels. For a definite shell, the subshell having higher value of ℓ possesses higher energy level. For example in 4[th] shell.

Energy level order

4f	>	4d	>	4p	>	4s
ℓ = 3		ℓ = 2		ℓ = 1		ℓ = 0



(c) The relative energy of sub shells of different energy shell can be explained in terms of the (n + ℓ) rule.

(i) The subshell with lower values of (n + ℓ) possess lower energy level.

For 3d n = 3 ℓ = 2

$\therefore$ n + ℓ = 5

For 4s n = 4 ℓ = 0

n + ℓ = 4

(ii) If the value of (n + ℓ) for two orbitals is same, one with lower values of 'n' possess lower energy level.

21. EXTRA STABILITY OF HALF FILLED AND COMPLETELY FILLED ORBITALS :

Half-filled and completely filled sub-shells have extra stability due to the following reasons.

21.1 Symmetry of orbitals :

(a) It is a well known fact that symmetry leads to stability.

(b) Thus, if the shift of an electron from one orbital to another orbital differing slightly in energy results in the symmetrical electronic configuration. it becomes more stable

(c) For example p^3, d^5, f^7 confiogurations are more stable than their near ones

21.2 Exchange Energy

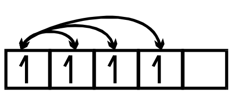
(a) The e^- in various subshells can exchange their positions, since e^- in the same subshell have equal energies.

(b) The energy is released during the exchange process with in the same subshell.

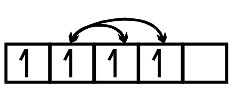
(c) In case of half filled and completely filled orbitals, the exchange energy is maximum and is greater then the loss of orbital energy due to the transfer of electron from a higher to a lower sublevel e.g. from 4s to 3d orbitals in case of Cu and Cr

(d) The greater the number of possible exchanges between the electrons of parallel spins present in the degenerate orbitals, the higher would be the amount of energy released and more will be the stability

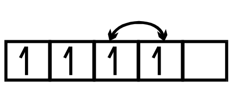
(e) Let us count the number of exchange that are possible in d^4 and d^5 configuration among electrons with parallel spins :

d^4 (1) 

3 exchanges by 1st e^-

(2) 

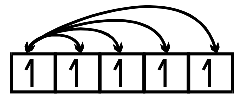
2 Exhanges by 2nd e^-

(3) 

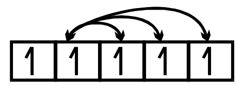
only 1 exchange by 3rd e^-

Total number of possible exchanges

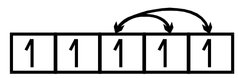
= 3 + 2 + 1 = 6

d^5 (1) 

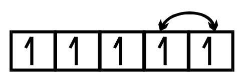
4 exchanges by 1st e^-

(2) 

3 exchanges by 2nd e^-

(3) 

2 exchanges by 3rd e^-

(D) 

1 exchange by 4th e^-

Total number of possible exchanges

= 4 + 3 + 2 + 1 = 10

22. ELECTRONIC CONFIGURATION OF ELEMENTS

Element	At.No.	1s	2s	2p	3s	3p	3d	4s	4p	4d	4f	5s	5p	6d	5f
H	1	1													
He	2	2													
Li	3	2	1												
Be	4	2	2												
B	5	2	2	1											
C	6	2	2	2											
N	7	2	2	3											
O	8	2	2	4											
F	9	2	2	5											
Ne	10	2	2	6											
Na	11	2	2	6	1										
Mg	12	2	2	6	2										
Al	13	2	2	6	2	1									
Si	14	2	2	6	2	2									
P	15	2	2	6	2	3									
S	16	2	2	6	2	4									
Cl	17	2	2	6	2	5									
Ar	18	2	2	6	2	6									
K	19	2	2	6	2	6		1							
Ca	20	2	2	6	2	6		2							
Sc	21	2	2	6	2	6	1	2							
Ti	22	2	2	6	2	6	2	2							
V	23	2	2	6	2	6	3	2							
*Cr	24	2	2	6	2	6	5	1							
Mn	25	2	2	6	2	6	5	2							
Fe	26	2	2	6	2	6	6	2							
Co	27	2	2	6	2	6	7	2							
Ni	28	2	2	6	2	6	8	2							
*Cu	29	2	2	6	2	6	10	1							
Zn	30	2	2	6	2	6	10	2							
Ga	31	2	2	6	2	6	10	2	1						
Ge	32	2	2	6	2	6	10	2	2						
As	33	2	2	6	2	6	10	2	3						
Se	34	2	2	6	2	6	10	2	4						
Br	35	2	2	6	2	6	10	2	5						
Kr	36	2	2	6	2	6	10	2	6						
Rb	37	2	2	6	2	6	10	2	6			1			
Sr	38	2	2	6	2	6	10	2	6			2			
Y	39	2	2	6	2	6	10	2	6	1		2			
Zr	40	2	2	6	2	6	10	2	6	2		2			
*Nb	41	2	2	6	2	6	10	2	6	4		1			
*Mo	42	2	2	6	2	6	10	2	6	5		1			
Tc	43	2	2	6	2	6	10	2	6	5		2			
*Ru	44	2	2	6	2	6	10	2	6	7		1			
*Rh	45	2	2	6	2	6	10	2	6	8		1			
*Pd	46	2	2	6	2	6	10	2	6	10					
*Ag	47	2	2	6	2	6	10	2	6	10		1			
Cd	48	2	2	6	2	6	10	2	6	10		2			
In	49	2	2	6	2	6	10	2	6	10		2	1		
Sn	50	2	2	6	2	6	10	2	6	10		2	2		
Sb	51	2	2	6	2	6	10	2	6	10		2	3		
Te	52	2	2	6	2	6	10	2	6	10		2	4		
I	53	2	2	6	2	6	10	2	6	10		2	5		
Xe	54	2	2	6	2	6	10	2	6	10		2	6		

Element	At.No.	K	L	M	4s	4p	4d	4f	5s	5p	5d	5f	6s	6p	6d	6f	7s
Cs	55	2	8	18	2	6	10		2	6			1				
Ba	56	2	8	18	2	6	10		2	6			2				
*La	57	2	8	18	2	6	10		2	6	1		2				
Ce	58	2	8	18	2	6	10	1	2	6	1		2				
Pr	59	2	8	18	2	6	10	3	2	6			2				
Nd	60	2	8	18	2	6	10	4	2	6			2				
Pm	61	2	8	18	2	6	10	5	2	6			2				
Sm	62	2	8	18	2	6	10	6	2	6			2				
Eu	63	2	8	18	2	6	10	7	2	6			2				
*Gd	64	2	8	18	2	6	10	7	2	6	1		2				
Tb	65	2	8	18	2	6	10	9	2	6			2				
Dy	66	2	8	18	2	6	10	10	2	6			2				
Ho	67	2	8	18	2	6	10	11	2	6			2				
Er	68	2	8	18	2	6	10	12	2	6			2				
Tm	69	2	8	18	2	6	10	13	2	6			2				
Yb	70	2	8	18	2	6	10	14	2	6			2				
Lu	71	2	8	18	2	6	10	14	2	6	1		2				
Hf	72	2	8	18	2	6	10	14	2	6	2		2				
Ta	73	2	8	18	2	6	10	14	2	6	3		2				
W	74	2	8	18	2	6	10	14	2	6	4		2				
Re	75	2	8	18	2	6	10	14	2	6	5		2				
Os	76	2	8	18	2	6	10	14	2	6	6		2				
Ir	77	2	8	18	2	6	10	14	2	6	7		2				
*Pt	78	2	8	18	2	6	10	14	2	6	9		1				
*Au	79	2	8	18	2	6	10	14	2	6	10		1				
Hg	80	2	8	18	2	6	10	14	2	6	10		2				
Ti	81	2	8	18	2	6	10	14	2	6	10		2	1			
Pb	82	2	8	18	2	6	10	14	2	6	10		2	2			
Bi	83	2	8	18	2	6	10	14	2	6	10		2	3			
Po	84	2	8	18	2	6	10	14	2	6	10		2	4			
At	85	2	8	18	2	6	10	14	2	6	10		2	5			
Rn	86	2	8	18	2	6	10	14	2	6	10		2	6			
Fr	87	2	8	18	2	6	10	14	2	6	10		2	6			1
Ra	88	2	8	18	2	6	10	14	2	6	10		2	6			2
*Ac	89	2	8	18	2	6	10	14	2	6	10		2	6	1		2
*Th	90	2	8	18	2	6	10	14	2	6	10	0	2	6	2		2
*Pa	91	2	8	18	2	6	10	14	2	6	10	2	2	6	1		2
*U	92	2	8	18	2	6	10	14	2	6	10	3	2	6	1		2
Np	93	2	8	18	2	6	10	14	2	6	10	4	2	6	1		2
Pu	94	2	8	18	2	6	10	14	2	6	10	6	2	6			2
Am	95	2	8	18	2	6	10	14	2	6	10	7	2	6			2
*Cm	96	2	8	18	2	6	10	14	2	6	10	7	2	6	1		2
*Bk	97	2	8	18	2	6	10	14	2	6	10	8	2	6	1		2
Cf	98	2	8	18	2	6	10	14	2	6	10	10	2	6			2
Fs	99	2	8	18	2	6	10	14	2	6	10	11	2	6			2
Fm	100	2	8	18	2	6	10	14	2	6	10	12	2	6			2
Md	101	2	8	18	2	6	10	14	2	6	10	13	2	6			2
No	102	2	8	18	2	6	10	14	2	6	10	14	2	6			2
*Lw	103	2	8	18	2	6	10	14	2	6	10	14	2	6	1		2
Ku	104	2	8	18	2	6	10	14	2	6	10	14	2	6	2		2
Ha	105	2	8	18	2	6	10	14	2	6	10	14	2	6	3		2

Examples based on e^- Configuration

Ex.23 For a given value of n (principal quantum number), the energy of different subshells can be arranged in the order of: -

(A) f > d > p > s (B) s > p < d > f
(C) f > p > d > s (D) s > f > p > d

Ans.(A)

Sol. It is the rule

Ex.24 Correct set of four quantum numbers for the outermost electron of rubidium (Z = 37) is -

(A) 5, 0, 0, 1/2 (B) 5, 1, 0, 1/2
(C) 5, 1, 1, 1/2 (D) 6, 0, 0, 1/2

Ans.(A)

Sol. Its configuration is $5s^1$

Ex.25 The order of increasing energies of the orbitals follows -

(A) 3s, 3p, 4s, 3d, 4p
(B) 3s, 3p, 3d, 4s, 4p
(C) 3s, 3p, 4s, 4p, 3d
(D) 3s, 3p, 3d, 4p, 4s **Ans.(A)**

Sol. Follow $(n + \ell)$ rule

Ex.26 The total spin resulting from a d^7 configuration is -

(A) 3/2 (B) 1/2
(C) 2 (D) 1. **Ans.(A)**

Sol. For d^7, three unpaired electrons, spin = 3 x $\frac{1}{2} = \frac{3}{2}$

23. RADIO ACTIVITY

(a) It is the property of the nucleus of the atom. Radio active element possesses unstable nucleus.

(b) It is discovered by Henery Becqural.

(c) The unstable nucleus gives $\alpha\ \left[{}_2He^4\right]$ or $\beta\ \left({}_{-1}e^{\circ}\right)$ particles and the product is nucleus of another element. During α – or β – emission the energy is emitted in the form of γ – radiation.

23.1 Nuclear forces :

(a) The force of attraction which binds protons & neutrons (nucleons) mutually is called as nuclear force.

(b) According to modern approach, nuclear force arises by the exchange of mesons between nuclear particles.

(c) Nuclear forces are stronger than repulsion forces between protons (having same charge).

(d) Due to exchange of mesons, exchange forces arise which binds nuclear particles to one another.

$$n + \pi^{\oplus} \longrightarrow p$$

$$p + \pi^{\ominus} \longrightarrow n$$

or

$$n \underset{\pi^{\ominus}}{\overset{\pi^{\oplus}}{\rightleftharpoons}} P$$

(e) Between two protons & two neutrons, neutral meson particles (Π^o) are exchanged.

$$P \underset{\pi^{\ominus}}{\overset{\pi^{\oplus}}{\rightleftharpoons}} P$$

$$n \underset{\pi^{\ominus}}{\overset{\pi^{\oplus}}{\rightleftharpoons}} n$$

23.2 Mass defect & binding energy of nucleus :

(a) The original mass of an atom is less than the total mass of original particles.

(b) When a stable nucleus is formed by protons & neutrons then a part of mass converts in energy & disappears so the mass of atomic nucleus is reduced as compared to total mass of nuclear particles by which it is formed.

(c) This loss of mass is called as mass defect.

Mass defect = (Total mass of nuclear particles)– (original mass of nucleus)

(d) When nuclear particles (protons & neutrons) form a stable atomic nucleus then mass defect takes place and energy emits which is called binding energy of nucleus.For e.g. in oxygen nucleus formation 127 Mev energy releases – which is called its binding energy.

23.3 Relation between mass defect & binding energy :

(a) Equivalent energy corresponding to mass defect is known as binding energy.

(b) If Δm mass defect occurs in the formation of a nucleus from nucleus particles then binding energy will be written as

$B_{En} = \Delta\ m \times 931$ Mev.

(c) Einstein equation is written as

E = mc^2 where E = Energy (arg.)

M = mol. wt. (gram)

C = velocity of light (cm/sec.)

According to the equation–

Equivalent energy of 1 amu mass is 931 Mev. So

$B_{En} = \Delta m \times 931$ Mev.

23.4 Binding energy per nucleon :

(a) Binding energy per nucleon is the measure of stability of the nucleus.

(b) Relative stability of isotopes and binding energy :

Consider the hypothetical nuclear reaction :

$${}_zX^A \xrightarrow{+\text{neutron}} {}_zX^{A_{1[+P\,MeV]}}$$

$${}_zX^A \xrightarrow{+\text{neutron}} {}_zX^{A_{2\,[-Q\,MeV]}}$$

(c) If the value of binding energy is positive then, The stability order is

Product nucleus > reactant nucleus

If the value of binding energy is negative then the stability order is

Product nucleus < reactant nucleus

23.5 Nuclear stability & the ratio of neutrons & protons :

(a) The stability of nucleus depends upon the ratio of neutrons (n) & protons (p)

(b) The nucleus in which $\frac{n}{p} = 1$ (approx), are very stable.

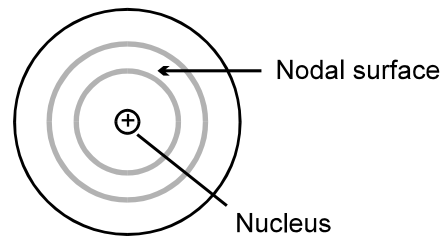
(c) When this ratio exceeds 1.5, then nucleus becomes unstable and radioactive.

(d) This ratio is approx. 1 (one) up to atomic number 1 to 20 & the atomic number above 83 the ratio is 1.5 to 1.6. So these atoms are radioactive.

for eg. in ${}^{40}Ca, \frac{n}{p} = \frac{20}{20} = 1$ so nucleus is stable.

in ${}^{235}_{92}U$, $\frac{n}{p} = \frac{235}{92} = 1.55$, so nucleus is unstable.

24. SOME IMPORTANT DEFINITIONS



(a) Nodal Surface :

The place find in between two 's' orbitals where the value of electron density is equal to zero called **Nodal surface**.

The number of Nodal surfaces in an atom = (n – 1), where 'n' is the number of total shell in an atom.

(b) Nodal Plane :

The place for 'p' and 'd' orbitals where the value of electron density is equal to zero called **Nodal Plane**.

For p_x = yz

p_y = xz

p_z = xy

For d_{xy} = yz, zx

d_{yz} = xy, xz

d_{x2-y2} = 0

d_{zx} = xy, yz

d_{z2} = 0

(c) Nodal Point :

The nucleus of an atom called **Nodal Point**.

(d) Isodiapheres :

The elements which have same value of (n – p) is called **Isodiapheres**.

eg.	${}_7N^{14}$	${}_8O^{16}$
Values of (n – p)	0	0

(e) Isotone :

Elements which contain same no. of neutron is called **Isotone**.

eg.	${}_{14}Si^{30}$	${}_{15}P^{31}$	${}_{16}S^{32}$
number of neutrons	16	16	

(f) Isotopes :

(i) First proposed by soddy.

(ii) The isotopes have same atomic number but different atomic weight.

(iii) They have same chemical properties because they have same atomic number.

(iv) They have different physical properties because they have different atomic masses.

eg.	${}_1H^1$	${}_1H^2$	${}_1H^3$
	Protonium	deuterium	Tritium
Z =	1	1	1
A =	1	2	3

(g) Isobar :

The two different atoms which have same atomic masses but different atomic number is called as **Isobar**.

eg.	$_{18}Ar^{40}$	$_{19}K^{40}$	$_{20}Ca^{40}$
Atomic mass	40	40	40
Atomic number	18	19	20

(h) Isomorphous :

The two different type of compound which contain same crystalline structure called **Isomorphous** and this property called **Isomorphism**.

eg.	$FeSO_4 \cdot 7H_2O$	$ZnSO_4 \cdot 7H_2O$
	Green vitriol	White vitriol
	Hepta hydrate Ferrous sulphate	Hepta hydrate Zinc Sulphate

(i) Isomers :

Species which have same molecular formula but different structural formula is called Isomer and this type of property is called Isomerism.

eg. $C_2H_6O \rightarrow C_2H_5OH$ & $CH_3 - O - CH_3$

(j) Isoelectronic :

Ion or atom or molecule or species which have the same number of electron is called Isoelectronic species.

eg.	$_{17}Cl^{-}$	$_{18}Ar$	$_{19}K^{+}$	$_{20}Ca^{+2}$
No. of electron	18	18	18	18

eg.	CN^{-}	CO
No. of electron	14	14

(k) Isosters :

Substance which have same number of electron and atoms called Isosters.

eg.	CO_2	N_2O
	22	22

(l) Kernel : Orbit which present after removing the outer most orbit of that atom is called kernel and electrons which is present that orbit called kernel electrons.

eg. Mg = $1s^2\ 2s^2\ 2p^6,\ 3s^2$

Total kernel electron = 2 + 2 + 6 = 10

(m) Core :

(i) The outer most shell of an any atom called **Core** and the number of electron present of that shell is called **Core electron**.

eg. Cl = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Core electron = 2 + 5 = 7

(ii) If the core is unstable for an atom then that atom shows variable valency.

(n) Photoelectric effect :

When a beam of light of high frequency is strike on a metal surface in vacuum condition, electrons are emitted from the metal surface. This phenomenon is called photoelectric effect and the emitted electron is called photoelectrons.

Total energy = $\frac{1}{2}mv^2 + \omega$

{$\frac{1}{2}mv^2$ = kinetic energy

ω = Threshold energy or work function}

(o) Threshold energy : The minimum energy required to emit an electron on the metal surface called threshold energy.

(p) The value of $\frac{e}{m}$ for n, p, α, & electron is equal to -

$\frac{e}{m}$ for n = 0

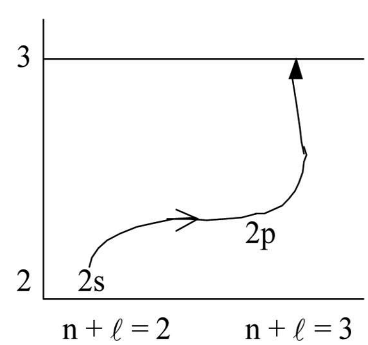
$$\frac{e}{m} \text{ for } \alpha = \frac{2 \times 1.6 \times 10^{-19}}{4 \times 1.67 \times 10^{-24}} = 4.8 \times 10^5$$

$$\frac{e}{m} \text{ for p} = \frac{1.6 \times 10^{-19}}{1.67 \times 10^{-24}} = 9.58 \times 10^4$$

$$\frac{e}{m} \text{ for } e^- = \frac{1.6 \times 10^{-19}}{9.1 \times 10^{-28}} = 1.76 \times 10^8$$

Note : When an electron is in the stationary state then the value of magnetic field for that electron is equal to zero.

(q) Promotion :



The transfer of electron between subshells in an orbit is called promotion. While the transfer of one energy level to another is called transition. After the completion of promotion the transition process is occurred.

eg. First promotion of an electron is 2s ($n + \ell = 2 + 0 = 2$) to 2p ($n + \ell = 2 + 1 = 3$) sub-shell and their transition to 2^{nd} orbit to 3^{rd} orbit or 2p to 3s.

25. SOME IMPORTANT POINTS

- The wave character is of no significance in case of large objects like cricket ball, a car, a train etc.
- The most important applications of de-Broglie concept is in the construction of electron microscope and the study of surface structure of solids by electron diffraction.
- Smaller the wavelength of the electron wave, more is the resolving power of the electron microscope
- Uncertainty in measurement is not due to lack of any experimental technique but due to nature of subatomic particle itself
- Shapes of orbitals are functional representation of mathematical solutions of Schrodinger equations. They do not represent any picture of electric charge or matter.

SOLVED EXAMPLE

Ex.1 Complete the following table -

Atom/ion	Atomic Number (Z)	Mass No.	Proton	Neutrons	Electrons
		(A)	(p)	(n)	(e)
Al^{3+}	13			14	
Cu	29	63			
Mg^{2+}	24				12
Sr		88	38		

Sol. (i) Atomic number (Z) = 13 = Number of protons
Number of electrons = 13 – 3 = **10**
Mass number = n + p = 14 + 13 = **27**

(ii) Atomic number = Number of protons
= Number of electrons = **29**
Mass number = n + p = **63**
since p = 29
$\therefore$ n = 63 – p = 63 – 29 = **34**

(iii) Number of protons = Z = **12** & Number of electrons = 12 – 2 = 10
Mass number = n + p = 24
$\therefore$ n = 24 – p = 24 – 12 = **12**

(iv) Number of electrons = Number of protons = Z = 38
Mass number = n + p = 88
$\therefore$ n = 88 – p = 88 – 38 = **50**

Ex.2 An oil drop has 6.39×10^{-19} C charge. Find out the number of electrons in this drop -

Sol. Charge on oil drop = 6.39×10^{-19} C
Now we know that
1.602×10^{-19} C is the charge on one 1 electron
$\therefore$ 6.39×10^{-19} C will be charge on =

$$\frac{6.39 \times 10^{-19}}{1.602 \times 10^{-19}} = \textbf{4 electrons}$$

Ex.3 Find out the number of wave made by a Bohr electron in one complete revolution in its 3rd orbit of hydrogen atom -

Sol. We know that

$$r_n = r_0 \times n^2$$

$\therefore$ $r_3 = 0.529 \times 10^{-8}\text{ cm} \times (3)^2$
($\because$ $r_0 = 0.529 \times 10^{-8}$ cm)
Also we know that

$$u_n = \frac{u_0}{n}$$

$$\therefore \quad u_3 = \frac{2.19 \times 10^8}{3}$$

($\because$ $u_0 = 2.19 \times 10^8$ cm sec^{-1})

No. of waves in one round = $\dfrac{2\pi r_3}{\lambda} = \dfrac{2\pi r_3}{h / mu_3}$

$$= \frac{2\pi r_3 \times u_3 \times m}{h}$$

Substituting the values of the different constants
No. of waves in one round

$$= \frac{2 \times 3.14 \times 0.529 \times 10^{-8} \times 9 \times 2.19 \times 10^8 \times 9.108 \times 10^{-28}}{3 \times 6.62 \times 10^{-27}}$$

= **3**

Ex.4 The ionization energy of He^+ is 19.6×10^{-18} J atom^{-1}. The energy of the first stationary state of Li^{+2} will be -
(A) 21.2×10^{-18} J/atom
(B) 44.10×10^{-18} J/atom
(C) 63.2×10^{-18} J/atom
(D) 84.2×10^{-18} J/atom **(Ans. B)**

Sol. E_1 for Li^{+2} = E_1 for H × Z^2_{Li} = E_1 for H × 9
E_1 for He^+ = E_1 for H × Z^2_{He} = E_1 for H × 4

or E_1 for Li^{+2} = $\dfrac{9}{4}$ E_1 for He^+

$= 19.6 \times 10^{-18} \times \dfrac{9}{4}$

= **44.10×10^{-18} J/atom**

Ex.5 The ionization energy of hydrogen atom is 13.6 eV. What will be the ionization energy of He^+ –

Sol. He^+ is a hydrogen like species i.e. the electron is ionised from first orbit.

$$\therefore \quad \text{Ionization energy of } He^+ = \frac{Z^2 E_H}{n^2}$$

$$= \frac{4 \times 13.6}{1^2} = \textbf{54.4 eV}$$

Ex.6 The ionization energy of H-aotm is 13.6 eV. The ionization energy of Li^{+2} ion will be -
(A) 13.6 eV (B) 27.2 eV
(C) 54.4 eV (D) 122.4 eV **(Ans D)**

Sol. E_1 for Li^{+2} = E_1 for H × Z^2 [for Li, Z = 3]
= 13.6 × 9
= **122.4 eV**

Ex.7 Which transition of the Hydrogen spectrum would have the same lenght as the Balmer transition, n = 4 to n = 2 of He^+ spectrum -
(A) $n_2 = 2$ to $n_1 = 1$ (B) $n_2 = 3$ to $n_1 = 1$
(C) $n_2 = 4$ to $n_1 = 2$ (D) $n_2 = 5$ to $n_1 = 3$
(Ans. A)

Sol. For He^+ ion, we have

$$\frac{1}{\lambda} = R_H Z^2\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$$

$$= R_H[2]^2\left[\frac{1}{2^2}-\frac{1}{4^2}\right]$$

$$= \frac{3}{4}R_H \qquad ...(A)$$

Now for H atom

$$\frac{1}{\lambda} = R_H\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right] \qquad ...(B)$$

Equating equs (A) and (B) we have

$$\frac{1}{n_1^2}-\frac{1}{n_2^2} = \frac{3}{4}$$

Obviously $n_1 = 1$ and $n_2 = 2$. Hence the transition n = 2 to n = 1 in hydrogen atom will have the same length as the transition n = 4 to n = 2 in He^+ species.

Ex.8 Given R = 1.0974×10^7 m^{-1} and h = 6.626×10^{-34} Js. The ionization energy of one mole of Li^{+2} ions will be as follows -

(A) 11240 KJ $mole^{-1}$ (B) 11180 KJ $mole^{-1}$

(C) 12350 KJ $mole^{-1}$ (D) 15240 KJ $mole^{-1}$

(Ans B)

Sol. The expression of Ionization energy is -

$\Delta E = RZ^2$ hc

For Li^{+2} ion, Z = 3, hence

ΔE = $(1.0974 \times 10^7\ m^{-1}) \times (9) \times (6.626 \times 10^{-34}$ J.S.$) \times (3 \times 10^8\ ms^{-1})$

= 1.964×10^{-17} J

For one mole of ions, we have

$\Delta E' = N_A . \Delta E$

= $(6.023 \times 10^{23}\ mol^{-1})(1.964 \times 10^{-17}$ J)

= 1.118×10^7 J mol^{-1}

= 11180 KJ mol^{-1}

Ex.9 Calculate the energy emitted when electron of 1.0 g atom of hydrogen undergo transition giving the spectral line of lowest energy in the visible region of its atomic spectrum -
(R_H = 1.1×10^7 m^{-1}, c = 3×10^8 ms^{-1}, h = 6.62×10^{-34} Js).

Sol. The spectral line lies in the visible regionn i.e., it corresponds to the Balmer series i.e. $n_2 = 2$ and hence n_1 = 3, 4, 5, etc.

For lowest energy $n_1 = 3$

Substituting the values in the following relation.

$$\frac{1}{\lambda} = R_H\left[\frac{1}{n_2^2}-\frac{1}{n_1^2}\right]$$

$$= 1.1 \times 10^7 \times \left[\frac{1}{4}-\frac{1}{9}\right]$$

$$= 1.1 \times 10^7 \times \frac{5}{36}$$

$$\lambda = \frac{36}{1.1 \times 10^7 \times 5}$$

$$= 6.55 \times 10^{-7}\ m$$

Now we know that

$$E = h\nu = h \times \frac{c}{\lambda}$$

$$= \frac{6.62 \times 10^{-34} \times 3 \times 10^8}{6.55 \times 10^{-7}} = 3.03 \times 10^{-19}\ J$$

$\therefore$ Energy corresponding to 1g atom of hydrogen
= $3.03 \times 10^{-19} \times 6.02 \times 10^{23}$
= 18.25×10^4 J = **182.5 KJ**

Ex.10 Estimate the difference in energy between 1st and 2nd Bohr orbit for a H atom. At what minimum atomic no., a transition from n = 2 to n = 1 energy level would result in the emission of X- ray with λ = 3.0×10^{-8} m. Which hydrogen spectrum like species does this atomic no. corresponds to -

Sol. E_1 for H = – 13.6 eV

$\therefore$ E_2 for H = (– 13.6/2^2) = – 13.6/4 = – 3.4 eV

$\therefore$ $E_2 - E_1$ = – 3.4 – (– 13.6) = **+ 10.2 eV**

Also for transition of H like atom ; λ = 3.0×10^{-8} m

$$\frac{1}{\lambda} = R_H . Z^2\left[\frac{1}{1^2}-\frac{1}{2^2}\right]$$

$$\frac{1}{3 \times 10^{-8}} = 1.09 \times 10^7 \times Z^2 \times \frac{3}{4}$$

$\therefore$ $Z^2 = 4$ and **Z = 2**

Ex.11 The shortest wave length in H spectrum of Lyman series when R_H = 109678 cm^{-1} is -

(A) 1215.67 Å (B) 911.7 Å

(C) 1002.7 Å (D) 1127.30 Å **(Ans B)**

Sol. For Lyman series $n_1 = 1$

For shortest 'l' of Lyman sereis the energy differnece in two levels showing transition should be maximum (i.e. $n_2 = \infty$).

$$\frac{1}{\lambda} = R_H\left[\frac{1}{1^2}-\frac{1}{\infty^2}\right]$$

= 109678

$\therefore$ $\lambda = 911.7 \times 10^{-8}$

= **911.7 Å**

Ex.12 The energy of an electron in the second and third Bohr orbits of the hydrogen atom is -5.42×10^{-12} ergs and -2.41×10^{-12} erg respectively. Calculate the wavelength of the emitted radiation when the electron drops from third to second orbit -

Sol. Here, $h = 6.62 \times 10^{-27}$ erg

$E_3 = -2.41 \times 10^{-12}$ erg

$E_2 = -5.42 \times 10^{-12}$ erg

$\Delta E = E_3 - E_2$

$= -2.41 \times 10^{-12} + 5.42 \times 10^{-12}$

Now we know that, $\Delta E = h\nu$

$$\nu = \frac{c}{\lambda} = \frac{\Delta E}{h} = \frac{3.01 \times 10^{-12}}{6.62 \times 10^{-27}}$$

$$\lambda = \frac{6.62 \times 10^{-27} \times 3 \times 10^{8}}{3.01 \times 10^{-12}}.$$

$\lambda = 6.6 \times 10^{-5}$ cm

Since, $1Å = 10^{-8}$ cm

$\lambda = 6.6 \times 10^{3}$ Å

Ex.13 Find the number of quanta of radiations of frequency 4.75×10^{13} sec^{-1}, required to melt 100 g of ice. The energy required to melt 1 g of ice is 350 J -

Sol. E $= nh\nu$

$= n \times 6.62 \times 10^{-34}$ J sec $\times 4.75 \times 10^{13}$ sec^{-1}

$= n \times 31.445 \times 10^{-21}$ J

Energy required to melt 100 g ice = 350 J × 100

= 35000 J

$n \times 31.445 \times 10^{-21} = 35000$

$$n = \frac{35000}{31.445 \times 10^{-21}} = \mathbf{1113 \times 10^{21}}$$

Ex.14 Calculate the number of photons emitted in 10 hours by a 60 W sodium lamp ($\lambda$of photon = 5893 Å) -

Sol. Energy emitted by sodium lamp in one sec.

= Watt. × sec = 60 × 1 J

Energy of photon emitted $= \frac{hc}{\lambda}$

$$= \frac{6.625 \times 10^{-34} \times 3 \times 10^{8}}{5893 \times 10^{-10}}$$

$= 3.37 \times 10^{-19}$ J

$\therefore$ No. of photons emitted per sec. $= \frac{60}{3.37 \times 10^{-19}}$

$\therefore$ No. of photons emitted in 10 hours

$= 17.8 \times 10^{19} \times 10 \times 60 \times 60$

$= \mathbf{6.41 \times 10^{24}}$

Ex.15 Calculate the wavelength of a moving electron having 4.55×10^{-25} J of kinetic energy -

Sol. Kinetic energy = ($\frac{1}{2}mu^2$) = 4.55×10^{-23} J

$$\therefore \quad u^2 = \frac{2 \times 4.255 \times 10^{-25}}{9.108 \times 10^{-31}}$$

$\therefore \quad u = 10^{3}$ m sec^{-1}

$$\therefore \quad \lambda = \frac{h}{mu} = \frac{6.625 \times 10^{-34}}{9.108 \times 10^{-31} \times 10^{3}}$$

$= \mathbf{7.27 \times 10^{-7}}$ **meter**

Ex.16 The minimum energy required to overcome the attractive forces electron and surface of Ag metal is 7.52×10^{-19} J. What will be the maximum K.E. of electron ejected out from Ag which is being exposed to U.V. light of λ = 360 Å –

(A) 36.38×10^{-19} Joule

(B) 6.92×10^{-19} Joule

(C) 57.68×10^{-19} Joule

(D) 67.68×10^{-19} Joule **(Ans B)**

Sol. Energy absorbed $= \frac{hc}{\lambda}$

$$= \frac{6.625 \times 10^{-27} \times 3.0 \times 10^{10}}{360 \times 10^{-8}}$$

$= 5.52 \times 10^{-11}$ erg

$= 5.52 \times 10^{-18}$ Joule

$= (7.52 \times 10^{-19}) - (.552 \times 10^{-19})$

$= 6.92 \times 10^{-19}$ Joule

Ex.17 In hydrogen atom , an electron in its normal state absorbs two times of the energy as if requires to escape (13.6 eV) from the atom. The wave length of the emitted electron will be -

(A) 1.34×10^{-10} m (B) 2.34×10^{-10} m

(C) 3.34×10^{-10} m (D) 4.44×10^{-10} m

(Ans C)

Sol. Energy absorbed by an atom

= 2 × 13.6 = 27.2 eV

Energy consumed in escape

= 13.6 eV

Energy converted into K.E.

$= 13.6 \times 1.6 \times 10^{-19}$ J

$$v = \sqrt{\frac{2KE}{m}} = \sqrt{\frac{2(13.6 \times 16 \times 10^{-19})}{9.1 \times 10^{-31}}}$$

ms^{-1} $= 2.18 \times 10^{6}$ ms^{-1}

$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}}{9.1 \times 10^{-31} \times 2.1 \times 10^{6}}$$

$= \mathbf{3.34 \times 10^{-10}}$ **m**

Ex.18 Show that the wavelength of a 150 g rubber ball moving with a velocity 50 m sec^{-1} is short enough to be observed -

Sol. $\because \quad \lambda = \frac{h}{mu}$

Given u = 50 m sec^{-1}

= 50 × 10^2 cm sec^{-} ; m = 150 g

$\therefore \quad \lambda = \frac{6.625 \times 10^{-27}}{150 \times 50 \times 10^2}$ = **8.83 × 10^{-33} cm**

The wavelength is much longer than the l of visible region and thus it will not be visible.

Ex.19 If an electron is present in n = 6 level. How many spectral lines would be observed in case of H atom–

(A) 10 (B) 15
(C) 20 (D) 25 **(Ans B)**

Sol. The no. of spectral lines is given by $\frac{n(n-1)}{2}$

when n = 6 then, the no. of spectral lines

$= \frac{6\ x\ (6-1)}{2} = \frac{6\ x\ 5}{2} = 15$

Ex.20 An electron beam can undergo diffraction by crystals. Through what potential should a beam of electrons be accelerated so that its wavelength becomes equal to 1.54 Å -

Sol. We know that

$$\frac{1}{2}mu^2 = eV$$

and $\lambda = \frac{h}{mu}$ or $u = \frac{h}{m\lambda}$ or $u^2 = \frac{h^2}{m^2\lambda^2}$

$$\therefore \quad \frac{1}{2}m \times \frac{h^2}{m^2\lambda^2} = eV$$

or $V = \frac{1}{2}m \times \frac{h^2}{m^2\lambda^2 \times e} = \frac{1}{2} \times \frac{h^2}{m\lambda^2 \times e}$

Substituting the values, we get

$$V = \frac{1}{2} \times \frac{(6.62 \times 10^{-34})^2}{9.108 \times 10^{-31} \times (1.54 \times 10^{-10})^2 \times 1.602 \times 10^{-19}}$$

= **63.3 volt**

Ex.21 What designation will you assign to an orbital having following quantum number –

(a) n = 3, ℓ = 1, m = –1
(b) n = 4, ℓ = 2, m = +2
(c) n = 5, ℓ = 0, m = 0
(d) n = 2, ℓ = 1, m = 0

Sol. (a) Since ℓ = 1 corresponds to p-orbital and m = –1 shows orientation either in x or y axis, thus this
orbital refers to $3p_x$ or $3p_y$

(d) $4d_{xy}$ or $4d_{x^2-y^2}$

(c) 5s

(d) $2p_z$

Ex.22 How many electrons in a given atom can have the following quantum numbers –

(a) n = 4, ℓ = 1
(b) n = 2, ℓ = 1, m = – 1, s = + ½
(c) n = 3
(d) n = 4, ℓ = 2, m = 0

Sol. (a) ℓ = 1 refers to p - subshell which has three orbitals (p_x, p_y and p_z) each having two electrons. Therefore, total number of electrons are **6**.

(b) ℓ = 1 refers to p - subshell, m = – 1 refers to p_x or p_y orbital whereas, s = +½ indicate for only 1 electron.

(c) For n = 3, ℓ = 0, 1,

2 ℓ = 0	m = 0	2 electrons
ℓ = 1	m = –1	6 electrons
ℓ = 2	m = –2 , –1, 0, +1, +2	10 electrons
Total electrons		18 electrons

Alternatively, number of electrons for any energy level is given by

$2n^2$ i.e. 2 × 3^2 = 18 electrons

(d) ℓ = 2 means d-subshell and m = 0 refer to dz^2 orbital

$\therefore$ Number of electrons are **2.**

Ex.23 Which of the following set of quantum numbers are not permitted -

(a) n = 3, ℓ = 2, m = – 1, s = 0
(b) n = 2, ℓ = 2, m = +1, s = – ½
(c) n = 2, ℓ = 2, m = + 1, s = – ½
(d) n = 3, ℓ = 2. m = – 2, s = + ½

Sol. (a) This set of quantum number is not permitted as value of 's' cannot be zero.

(b) This set of quantum number is not permitted as the value of '*l*' cannot be equal to 'n'.

(c) This set of quantum number is not permitted as the value of '*l*' cannot be equal to 'n'.

(d) This set of quantum number is permitted.

Ex.24 Naturally occuring boron consists of two isotops whose atomic weights are 10.01 and 11.01. The atomic weight of natural boron is 10.81. Calculate the percentage of each isotope in natural boron-

Sol. Let the percentage of isotope with atomic wt. 10.01 = x

$\therefore$ Percentage of isotope with atomic wt. 11.01 = 100 – x

Average atomic wt. = $\dfrac{m_1x_1 + m_2x_2}{x_1 + x_2}$

or Average atomic wt. = $\dfrac{x \times 10.01 + (100 - x) \times 11.01}{100}$

$10.81 = \dfrac{x \times 10.01 + (100 - x) \times 11.01}{100}$ $x = 20$

$\therefore$ % of isotope with atomic wt. 10.01 = **20**

% of isotope with atomic wt. 11.01 = 100 - x = **80**

Ex.25 From the following list of atoms, choose the isotopes, isobars and isotones -

${}^{16}_{8}O$, ${}^{39}_{19}K$, ${}^{235}_{92}U$, ${}^{40}_{19}K$, ${}^{14}_{7}N$, ${}^{18}_{8}O$, ${}^{14}_{6}C$, ${}^{40}_{20}Ca$, ${}^{238}_{92}U$

Sol. Isotopes :

(${}^{16}_{8}O$, ${}^{18}_{8}O$), (${}^{39}_{19}K$, ${}^{40}_{19}K$), (${}^{235}_{92}U$, ${}^{238}_{92}U$)

Isobars : (${}^{40}_{19}K$, ${}^{40}_{20}Ca$) , (${}^{14}_{7}N$, ${}^{14}_{6}C$)

Isotones : (${}^{39}_{19}K$, ${}^{40}_{20}Ca$), (${}^{14}_{6}C$, ${}^{16}_{8}O$)

Ex.26 Atomic radius is the order of 10^{-8} cm. and nuclear radius is the order of 10^{-13} cm. Calculate what fraction of atom is occupied by nucleus -

Sol. Volume of nucleus = (4/3)pr^3

= (4/3)p × $(10^{-13})^3$ cm^3

volume of atom = 4/3 pr^3 = (4/3) p × $(10^{-8})^3$ cm^3

$\therefore \dfrac{V_{nucleus}}{V_{atom}} = \dfrac{10^{-39}}{10^{-24}} = 10^{-15}$

or $\mathbf{V_{nucleus} = 10^{-15} \times V_{atom}}$

Ex.27 Nitrogen atom has Atomic number 7 & oxygen has Atomic number 8. Calculate the total number of electrons in nitrate ion -

Sol. No. of electrons in NO_3^-

= (Electrons in N) + (3 × electrons in O)

+ [1(due to negative charge)]

= 7 + 3 × 8 + 1 = **32**