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

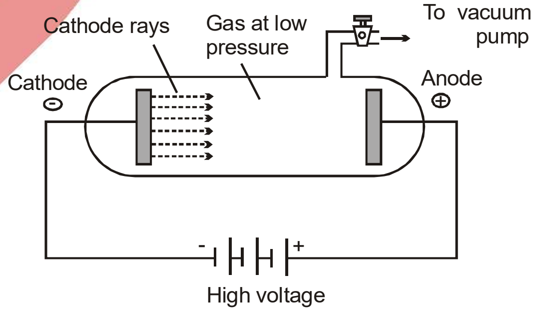
(a) Electron was discovered by Sir J.J. Thomson
(b) The charge on the electron is 1.6×10^{-19} coulomb/gm (Millikan)
(c) The molar mass of electron is 5.48×10^{-4} gm/mole
(d) The mass of electron in motion is expressed as

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

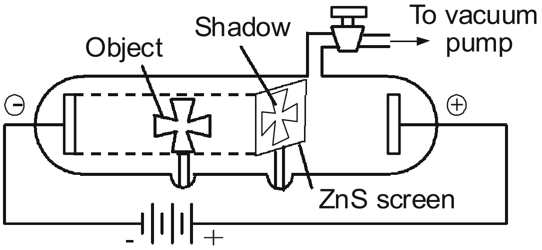
where m′ = mass of the electron in motion
m = rest mass, v = velocity of the electron
c = velocity of light

(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.

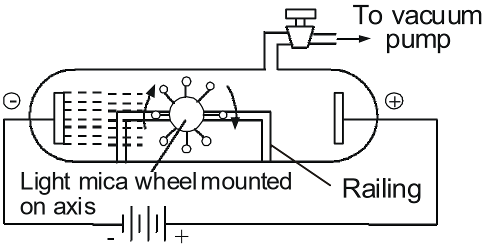
3.1.1 Cathode rays



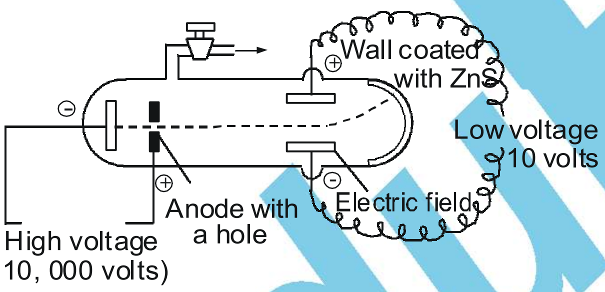
(a) The electron was discovered as a result of the studies of the passage of electricity through gases at extremely low pressures known as discharge tube experiments.
(b) When a high voltage of the order of 10,000 volts or more was impressed across the electrodes, some sort of invisible rays moved from 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.



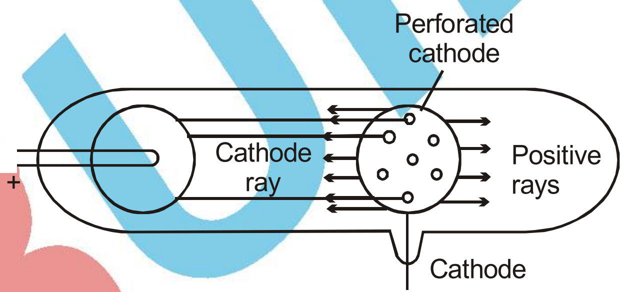
(iv) Cathode rays produce X-rays when they strike against hard metals like tungsten, copper etc.
(v) When the cathode rays are allowed to strike a thin metal foil, it gets heated up. Thus the cathode rays possess heating effect.
(vi) They produce a green glow when strike the glass wall beyond the anode. 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 $+1.602 \times 10^{-19}$ coulomb, i.e., one unit positive charge.
(c) Mass of proton is 1.672×10^{-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.

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

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

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

4.2Neutrino 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

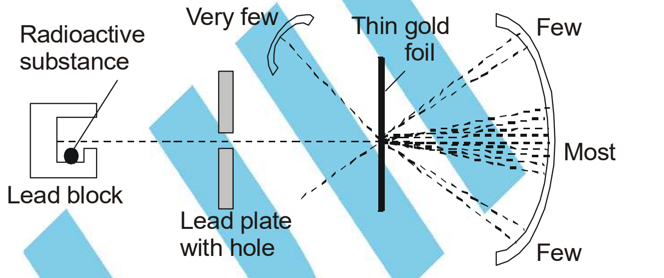
(a) After discovery of electron and proton attempts were made to find out their arrangement in an atom. The first simple model was proposed by J.J. Thomson known as Thomson's atomic model.

(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

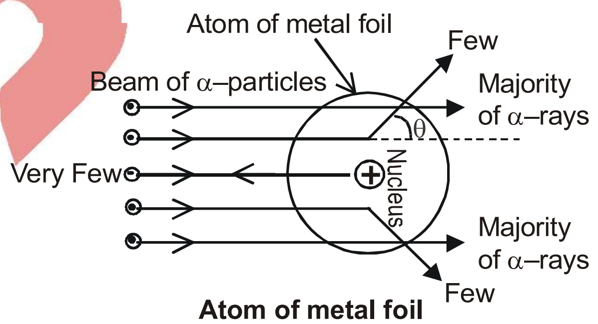
Rutherford carried out experiment on the bombardment of atoms by high speed positively charged α - particles emitted from radium and gave the following observations, which was based on his experiment.



(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.

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

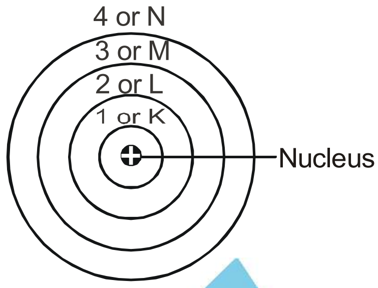
On the another side of that statement since mass of a proton or a neutron is not a whole number (on atomic weight scale), atomic weight is not necessarily a whole number.

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 $h/2\pi$ (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 $\quad \Delta E = E_2 - E_1$

$\uparrow \qquad\qquad \downarrow$

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

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

$\downarrow \qquad\qquad \downarrow$

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

$$= \frac{1}{4\pi\varepsilon_0} = 9 \times 10^9 \ Nm^2/C^2$$

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

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

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

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

From the postulate of Bohr,

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

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

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

From equation (1) and (2) :

on solving, we will get

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

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

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

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

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

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

0In C.G.S. unit K = 1

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

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

Then,

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

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

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

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. :

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

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

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

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)

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

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

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

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

(d) T. E. = – 313.6 × $\frac{Z^2}{n^2}$ 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

$\Delta E = E_2 - E_1$

$h\nu = E_2 - E_1$

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

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

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

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

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 × 10^{-28} gram

by the theoretical value of R_H = 109737 / cm

e = 4.8 × 10^{-10} e.s.u.

by the practical value of R_H = 109677 / cm

c = 3 × 10^{10} cm/sec

by the calculative value of R_H = 109700 / cm

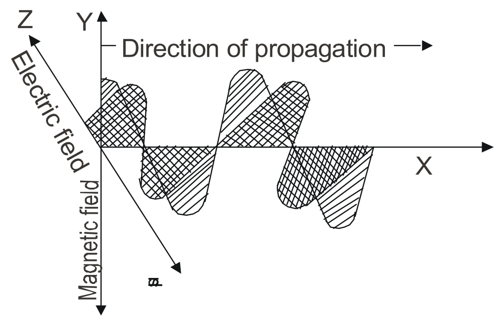
h = 6.625 × 10^{-27} erg-sec

Rydberg constang for other atom R = R_H x 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

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

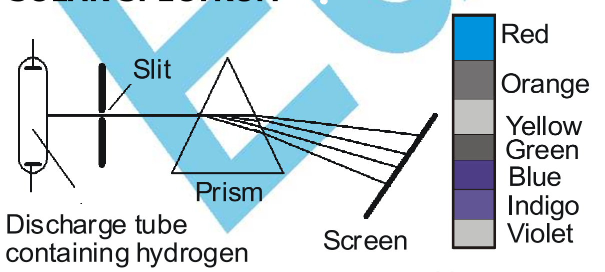
(c) Wave No. ($\bar{\nu}$) : It is defined as the number of wave per cm and it is equal to the inverse of wavelength. Its unit is cm^{-1}.

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

(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} toby 3×10^{14}	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 to7600	
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 lines consist of bright coloured lines separated by dark spaces.	It is consist of dark 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 $= \frac{n_1^2}{R_H} = \frac{5^2}{R_H} = \frac{25}{R_H}$

(f) $\frac{1}{\lambda} = R_H\left[\frac{1}{5^2} - \frac{1}{n_2^2}\right]$ 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 $= \frac{n_1^2}{R_H} = \frac{6^2}{R_H} = \frac{36}{R_H}$

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

$$E \propto \nu \text{ or } E = h\nu$$

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 ......

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

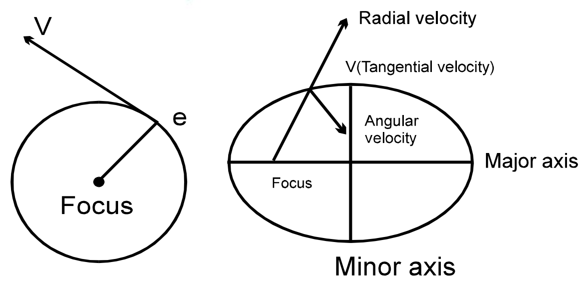
(b) He proposed that the moving electron might describe elliptical orbits in addition to circular, orbits and the nucleus is situated at one of the foci.

(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 $= n_r \dfrac{h}{2\pi}$, Azimuthal momentum $= n_\phi \dfrac{h}{2\pi}$

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 $\qquad$ ϕ = change

r = change $\qquad$ r = constant

$$P = P_r + P_\phi = \frac{n_r h}{2\pi} + \frac{n_\phi h}{2\pi} \qquad n = n_r + n_\phi$$

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

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

$$E = h\nu = \frac{h\,.\,c}{\lambda}$$

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

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

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

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

where h is planck's constant

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

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

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.

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

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 = $(n + \ell)$. The electron enters in that sub orbit whose $(n + \ell)$ 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 $-\ell$ to $+\ell$ 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

18.1Shape of orbitals on the basis of quantum number

18.1.1Shape 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.

18.1.2Shape 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.

18.1.3Shape 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.

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 -

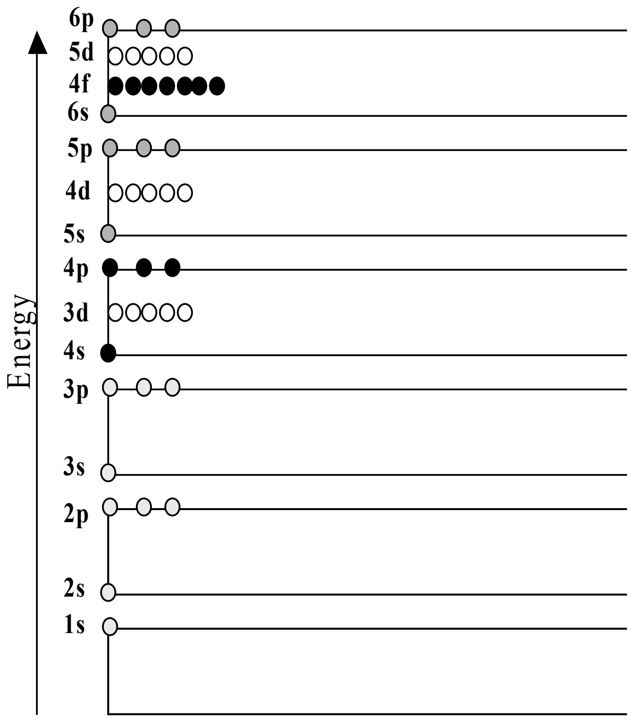


Diagram of multielectron atoms reveals the following points

(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 4th shell.

Energy level order

4f	> 4d	> 4p	> 4s
$\ell = 3$	$\ell = 2$	$\ell = 1$	$\ell = 0$

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

(i) The subshell with lower values of $(n + \ell)$ possess lower energy level.

For 3d $n = 3$ $\ell = 2$

$\therefore$ $n + \ell = 5$

For 4s $n = 4$ $\ell = 0$

$n + \ell = 4$

(ii) If the value of $(n + \ell)$ 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

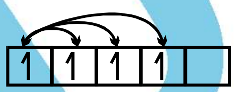
(a) The e^- in various subshells can exchange their positions, since e^- in the same subshell have equal energies.

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

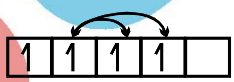
(c) In case of half filled and completely filled orbitals, the exchange energy is maximum and is greater then the loss of orbital energy due to the transfer of electron from a higher to a lower sublevel e.g. from 4s to 3d orbitals in case of Cu and Cr

(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

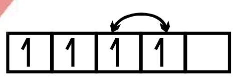
(e) Let us count the number of exchange that are possible in d^4 and d^5 configuration among electrons with parallel spins :

d^4 (1) 

3 exchanges by 1st e^-

(2) 

2 Exhanges by 2nd e^-

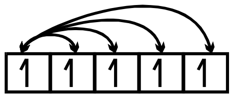
(3) 

only 1 exchange by 3rd e^-

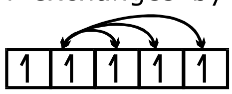
Total number of possible exchanges

= 3 + 2 + 1

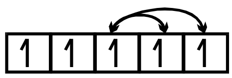
= 6

d^5 (1) 

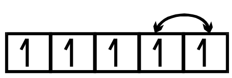
4 exchanges by 1st e^-

(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

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.

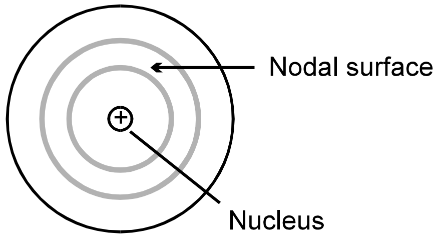
(c) When this ratio exceeds 1.5, then nucleus becomes unstable and radioactive.

(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 sulphate	Hepta hydrate Ferrous 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

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

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

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

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 2nd orbit to 3rd 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 PROBLEMS

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

$r_n = r_0 \times n^2$

$\therefore$ $r_3 = 0.529 \times 10^{-8}$ cm $\times (3)^2$

($\because$ $r_0 = 0.529 \times 10^{-8}$ cm)

Also we know that

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

$$\therefore\ u_3 = \frac{2.19 \times 10^8}{3} \quad (\because\ u_0 = 2.19 \times 10^8 \text{ cm sec}^{-1})$$

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

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

Substituting the values of the different constants

No. of waves in one round

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

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 $\times Z^2_{Li}$ = E_1 for H $\times$ 9 E_1 for He^+

$= E_1$ for H $\times Z^2_{He}$ = E_1 for H $\times$ 4

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

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

$= \mathbf{44.10 \times 10^{-18}}$ **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.

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

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

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 $\times Z^2$ [for Li, Z = 3]

$= 13.6 \times 9$

$=$ **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

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

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

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

Now for H atom

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

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

$R = 1.0974 \times 10^7$ m^{-1} and $h = 6.626 \times 10^{-34}$ 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

$\Delta E = (1.0974 \times 10^7 \text{ m}^{-1}) \times (9)$

$\times (6.626 \times 10^{-34}$ J.S.$) \times (3 \times 10^8 \text{ ms}^{-1})$

$= 1.964 \times 10^{-17}$ J

For one mole of ions, we have

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

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

$= 1.118 \times 10^7$ 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

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

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

∴ 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
∴ E_2 for H = (– 13.6/2^2) = – 13.6/4 = – 3.4 eV
∴ $E_2 - E_1$ = – 3.4 – (– 13.6) = **+ 10.2 eV**
Also for transition of H like atom ; λ = 3.0 x 10^{-8} m

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

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

∴ 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
∴ λ = 911.7×10^{-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 = 6.62×10^{-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, ΔE = hν

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

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

λ = 6.6×10^{-5} cm
Since, 1Å = 10^{-8} cm
λ = 6.6×10^3 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 = nhν
= n× 6.62×10^{-34} J sec × 4.75 × 10^{13} sec^{-1}
= n × 31.445 × 10^{-21} J
Energy required to melt 100 g ice = 350 J × 100 = 35000 J
n × 31.445 × 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 (λ of photon = 5893 Å) -

Sol. Energy emitted by sodium lamp in one sec.
= Watt. × sec = 60 × 1 J
Energy of photon emitted $= \frac{hc}{\lambda}$

$$= \frac{6.625 \times 10^{-34} \times 3 \times 10^{8}}{5893 \times 10^{-10}} = 3.37 \times 10^{-19}\ \text{J}$$

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

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

$= 17.8 \times 10^{19} \times 10 \times 60 \times 60 =$ **6.41×10^{24}**

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

Sol. Kinetic energy = (½mu^2) = 4.55×10^{-23} J

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

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

$$\therefore \lambda = \frac{h}{mu} = \frac{6.625 \times 10^{-34}}{9.108 \times 10^{-31} \times 10^{3}} = \mathbf{7.27 \times 10^{-7}\ meter}$$

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 $= \dfrac{hc}{\lambda}$

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

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

$= 2 \times 13.6 = 27.2$ eV

Energy consumed in escape

$= 13.6$ eV

Energy converted into K.E.

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

$$v = \sqrt{\frac{2KE}{m}} = \sqrt{\frac{2(13.6 \times 16 \times 10^{-19})}{9.1 \times 10^{-31}}}\ \text{ms}^{-1}$$

$= 2.18 \times 10^6$ 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. $\because \lambda = \dfrac{h}{mu}$

Given u = 50 m sec^{-1}

$= 50 \times 10^2$ cm sec^{-} ; m = 150 g

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

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

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

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

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

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

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

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

Substituting the values, we get

$$V = \frac{1}{2} \times \frac{(6.62 \times 10^{-34})^2}{9.108 \times 10^{-31} \times (1.54 \times 10^{-10})^2 \times 1.602 \times 10^{-19}} = \mathbf{63.3\ volt}$$

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. $2 \times 3^2 = 18$ 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 = $(4/3)pr^3$

$= (4/3)p \times (10^{-13})^3\ cm^3$

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

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

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**

EXERCISE # I — UNSOLVED PROBLEMS

Q.1 The study of cathode rays (i.e. electronic discharge through gases) shows that -
(A) Alpha particles are heavier than protons
(B) All forms of matter contain electrons
(C) All nuclei contain protons
(D) e/m is constant

Q.2 Proton is -
(A) Nucleus of deuterium
(B) Ionised hydrogen molecule
(C) Ionised hydrogen atom
(D) An α-particle

Q.3 Which is not deflected by magnetic field -
(A) Neutron (B) Positron
(C) Proton (D) Electron

Q.4 According to Dalton's atomic theory, an atom can –
(A) Be created
(B) Be destroyed
(C) Neither be created nor destroyed
(D) None

Q.5 Rutherford's experiment on scattering of alpha particles showed for the first time that atom has -
(A) Electrons (B) Protons
(C) Nucleus (D) Neutrons

Q.6 α - particles are represented by –
(A) Lithium atoms (B) Helium nuclei
(C) Hydrogen nucleus (D) None of these

Q.7 The energy of electron in first Bohr's orbit of H-atom is –13.6 eV. What will be its potential energy in n = 4th orbit -
(A) – 13.6 eV (B) –3.4 eV
(C) –0.85 eV (D) –1.70 eV

Q.8 The frequency of line spectrum of sodium is 5.09×10^{14} sec^{-1}. Its wave length (in nm) will be - [c = 3×10^{8} m/sec]-
(A) 510 nm (B) 420 nm
(C) 589 nm (D) 622 nm

Q.9 The spectrum of He-atom may be considered similar to the spectrum of -
(A) H (B) Li^{+} (C) Na (D) He^{+}

Q.10 Supposing the energy of fourth shell for hydrogen atom is - 50 a.u. (arbitrary unit). What would be its ionization potential -
(A) 50 (B) 800 (C) 15.4 (D) 20.8

Q.11 Supposing the ionization energy of hydrogen atom is 640 eV. Point out the main shell having energy equal to – 40 eV -
(A) n = 2 (B) n = 3 (C) n = 4 (D) n = 5

Q.12 A 1-kW radio transmitter operates at a frequency of 880 Hz. How many photons per second does it emit -
(A) 1.71×10^{21} (B) 1.71×10^{33}
(C) 6.02×10^{23} (D) 2.85×10^{26}

Q.13 On Bohr's stationary orbits -
(A) Electrons do not move
(B) Electrons move emitting radiations
(C) Energy of the electron remains constant
(D) Angular momentum of the electron is $h/2\pi$

Q.14 The value of Bohr radius of hydrogen atom is -
(A) 0.529×10^{-7}cm (B) 0.529×10^{-8}cm
(C) 0.529×10^{-9}cm (D) 0.529×10^{-10} cm

Q.15 On the basis of Bohr's model, the radius of the 3rd orbit is -
(A) Equal to the radius of first orbit
(B) Three times the radius of first orbit
(C) Five times the radius of first orbit
(D) Nine time the radius of first orbit

Q.16 The correct expression derived for the energy of an electron in the nth energy level is -
(A) $E_n = \frac{2\pi^2me^4}{n^2h^2}$ (B) $E_n = -\frac{2\pi^2me^4}{nh^2}$
(C) $E_n = -\frac{2\pi^2me^2}{n^2h^2}$ (D) $E_n = -\frac{2\pi^2me^4}{n^2h^2}$

Q.17 Ionization energy for hydrogen atom in ergs, Joules and eV respectively is -
(A) 21.8×10^{-12} , 218×10^{-20} , 13.6
(B) $13.6 \times 218 \times 10^{-20}$, 21.8×10^{-13}
(C) 21.8×10^{-20}, 13.6 , 21.8×10^{-13}
(D) 21.8×10^{-13}, 13.6, 21.8×10^{-20}

Q.18 The velocity of an electron in the third orbit of hydrogen atom -
(A) 7.28×10^{7} cm sec^{-1}
(B) 7.08×10^{7} cm sec^{-1}
(C) 7.38×10^{7}cm sec^{-1}
(D) 7.48×10^{7}cm sec^{-1}

Q.19 The ionization energy of a hydrogen atom is 13.6eV. The energy of the third-lowest electronic level in doubly ionized lithium (Z = 3) is -
(A) –28.7 eV (B) –54.4 eV
(C) –122.4 eV (D) –13.6 eV

Q.20 The momentum of a photon with energy 20 eV is-
(A) 10.66×10^{-27} Kg m sec^{-1}
(B) 10.55×10^{-27} Kg m sec^{-1}
(C) 10.60×10^{-27} Kg m sec^{-1}
(D) 10.80×10^{-27} Kg m sec^{-1}

Q.21 For ionising an excited hydrogen atom, the energy required in eV will be -
(A) 3.4 or less (B) More than 13.6
(C) Little less than 13.6 (D) 13.6

Q.22 A gas absorbs a photon of 300 nm and then re-emitts two photons. One photon has a wavelength 600 nm. The wavelength of second photon is -
(A) 300 nm (B) 400 nm
(C) 500 nm (D) 600 nm

Q.23 The specific charge of a proton is 9.6 x 10^7C kg^{-1}, then for an α-particles it will be -
(A) 2.4 x 10^7C kg^{-1} (B) 4.8 x 10^7C kg^{-1}
(C) 19.2 x 10^7C kg^{-1} (D) 38.4 x 10^7C kg^{-1}

Q.24 For H– atom, the energy required for the removal of electron from various sub-shells is given as under– The order of the energies would be –
(A) $E_1 > E_2 > E_3$ (B) $E_3 > E_2 > E_1$
(C) $E_1 = E_2 = E_3$ (D) None of these

Q.25 The wave number of the first line of Balmer series of hydrogen is 15200 cm^{-1}. The wave number of the first Balmer line of Li^{2+} ion is-
(A) 15200cm^{-1} (B) 60800 cm^{-1}
(C) 76000 cm^{-1} (D) 136800 cm^{-1}

Q.26 The wavelength of the third line of the Balmer series for a hydrogen atom is -
(A) $\frac{21}{100R_H}$ (B) $\frac{100}{21\ R_H}$ (C) $\frac{21R_H}{100}$ (D) $\frac{100R_H}{21}$

Q.27 Wave number of a spectral line for a given transition is x cm^{-1} for He^+, then its value for Be^{3+} for the same transition is -
(A) 4x cm^{-1} (B) x cm^{-1}
(C) x/4 cm^{-1} (D) 2x cm^{-1}

Q.28 A photon was absorbed by a hydrogen atom in its ground state and the electron was promoted to the fifth orbit. When the excited atom returned to its ground state, visible and other quanta were emitted. Other quanta are -
(A) $2 \to 1$ (B) $5 \to 2$
(C) $3 \to 1$ (D) $4 \to 1$

Q.29 Wave-length of the first line of Paschen Series hydrogen spectrum is - (R = 109700 cm^{-1}) -
(A) 18750 (Å) (B) 2854 (Å)
(C) 3452 (Å) (D) 6243 (Å)

Q.30 What is the change in the orbit radius when the electron in the hydrogen atom (Bohr model) undergoes the first Paschen transition -
(A) 4.23 x 10^{-10} m (B) 0.35 x 10^{-10} m
(C) 3.7 x 10^{-10} m (D) 1.587 x 10^{-10} m

Q.31 If the shortest wavelength of H atom in Lyman series is x, then longest wavelength in Balmer series of He^+ is -
(A) $\frac{9x}{5}$ (B) $\frac{36x}{5}$ (C) $\frac{x}{4}$ (D) $\frac{5x}{9}$

Q.32 Which of the following expressions represents the spectrum of Balmer series(If n is the principal quantum number of higher energy level) -
(A) $\bar{v} = \frac{R(n-1)(n+1)}{n^2} cm^{-1}$
(B) $\bar{v} = \frac{R(n-2)(n+2)}{4n^2} cm^{-1}$
(C) $\bar{v} = \frac{R(n-2)(n+2)}{n^2} cm^{-1}$
(D) $\bar{v} = \frac{R(n-1)(n+1)}{4n^2} cm^{-1}$

Q.33 The maximum number of electron in a principal shell is -
(A) 2n (B) $2n^2$ (C) 2 (D) $\sqrt{2}$ n

Q.34 Which of the following statements concerning the four quantum numbers is false -
(A) n gives idea of the size of an orbital
(B) ℓ gives the shape of an orbital
(C) m gives the energy of the electron in the orbital
(D) s gives the direction of spin of the electron in an orbital

Q.35 How many electrons can fit into the orbitals that comprise the 3rd quantum shell n = 3 -
(A) 2 (B) 8 (C) 18 (D) 32

Q.36 The shape of the orbital is given by -
(A) Spin quantum number
(B) Magnetic quantum number
(C) Azimuthal quantum number
(D) Principal quantum number

Q.37 The set of quantum numbers not applicable for an electron in an atom is –
(A) n = 1, ℓ = 1, m = 1, s = + 1/2
(B) n = 1, ℓ = 0, m = 0, s = + 1/2
(C) n = 1, ℓ = 0, m = 0, s = – 1/2
(D) n = 2, ℓ = 0, m = 0, s = + 1/2

Q.38 Maximum numbers of electrons in a subshell is given by -
(A) $(2\ell + 1)$ (B) $2(2\ell+1)$
(C) $(2\ell+1)^2$ (D) $2(2\ell + 1)^2$

Q.39 The magnetic quantum number for valency electron of sodium atom is -
(A) 3 (B) 2 (C) 1 (D) Zero

Q.40 Which one of the following represents an impossible arrangement –

	n	ℓ	m	s
(A)	3	2	–2	1/2
(B)	4	0	0	1/2
(C)	3	2	–3	1/2
(D)	5	3	0	1/2

Q.41 The set of quantum number for the 19^{th} electrons in chromium is -
(A) n=4, ℓ=0, m=0, s =+1/2 or -1/2
(B) n=3, ℓ=2, m=1, s=+1/2 or -1/2
(C) n=3, ℓ=2 m= -1, s=+1/2 or -1/2
(D) n=4, ℓ=1, m=0, s=+1/2 or -1/2

Q.42 The electronic configuration together with the quantum number of last electron for lithium is -
(A) $1s^2 2s^1$ 2, 0, 0 + $\frac{1}{2}$
(B) $1s^2 2s^1$ 2, 0, 0 + $\frac{1}{2}$ or – $\frac{1}{2}$
(C) $1s^2 2s^0 2p^1$ 2, 1, 0 ± $\frac{1}{2}$
(D) $2s^2 2s^1$ 2,1, 0 ± $\frac{1}{2}$

Q.43 The electronic configuration together with the quantum number of last electron for lithium is -
(A) $1s^2\ 2s^1$ 2, 0, 0, + $\frac{1}{2}$
(B) $1s^2\ 2s^1$ 2, 0, 0, + $\frac{1}{2}$ or – $\frac{1}{2}$
(C) $1s^2\ 2s^0\ 2p^1$ 2, 1, 0, ± $\frac{1}{2}$
(D) $2s^2\ 2s^1$ 2, 1, 0, ± $\frac{1}{2}$

Q.44 Four sets of values of quantum numbers (n, ℓ, m and s) are given below. Which set does not provide a permissible solution of the wave equation -
(A) 3, 2, –2, $\frac{1}{2}$ (B) 3, 3, 1, $-\frac{1}{2}$
(C) 3, 2, 1, $\frac{1}{2}$ (D) 3, 1, 1, $\frac{1}{2}$

Q.45 In presence of magnetic field d-sub orbit is -
(A) 5 - Fold degenate (B) 3-Fold degenerate
(C) 7-Fold degenerate (D) Non- degenerate

Q.46 In which of the following pairs is the probability of finding the electron in xy-plane zero for both orbitals ?
(A) $3d_{yz}, 4d_{x^2-y^2}$ (B) $2p_z, dz^2$
(C) $4d_{zx}$, $3p_z$ (D) All of these

Q.47 For $4p_y$ orbital : There are -
nodal plane = ... and azimuthal quantum number ℓ =
(A) 1, 0 (B) 0, 1 (C) 1, 1 (D) 2, 1

Q.48 The maximum probability of finding electron in the d_{xy} orbital is -
(A) Along the x axis
(B) Along the y axis
(C) At an angle of 45º from the x and y axis
(D) At an angle of 90º from the x and y axis

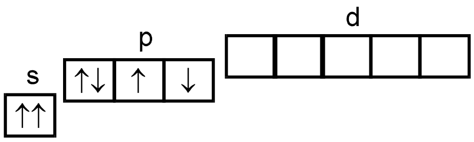
Q.49 An electron has a spin quantum number + 1/2 and a magnetic quantum number –1. It cannot be present in -
(A) d-Orbital (B) f-Orbital
(C) s-Orbital (D) p-Orbital

Q.50 If the electronic structure of oxygen atom is written as $1s^2$, $2s^2$ 2p [↑↓|↑↓|] it would violate -
(A) Hund's rule
(B) Paulis exclusion principle
(C) Both Hund's and Pauli's principles
(D) None of these

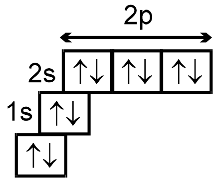
Q.51 The energy of an electron of $2p_y$ orbital is -
(A) Greater than $2p_x$ orbital
(B) Less than $2p_z$ orbital
(C) Equal to 2s orbital
(D) Same as that of $2p_x$ and $2p_z$ orbitals

Q.52 The number of unpaired electrons in carbon atom is -
(A) 2 (B) 4 (C) 1 (D) 3

Q.53 When 4 d orbital is complete, the newly entering electrons goes in to -
(A) 5f (B) 5s
(C) 5p (D) 6d Orbital

Q.54 

The above configuration is not correct as it violates -
(A) Only Hund's rule
(B) Only Pauli's exclusion principle
(C) (n + ℓ) rule
(D) (Hund + Pauli) rule

Q.55 Which of the following elements is represented by the electronic configuration -


(A) Nitrogen (B) Fluorine
(C) Oxygen (D) Neon

Q.56 The electronic configurations of Cr^{24} and Cu ^{29} are abnormal -
(A) Due to extra stability of exactly half filled and exactly fully filled sub shells
(B) Because they belong to d-block
(C) Both the above
(D) None of the above

Q.57 The electronic configuration of chromium (Z = 24) is -
(A) $[Ne]3s^2 3p^6 3d^4 4s^2$
(B) $[Ne]\ 3s^2 3p^6 3d^5 4s^1$
(C) $[Ne]3s^2 3p^6 3d^1 4s^2$
(D) $[Ne]\ 3s^2 3p^6 4s^2 4p^4$

Q.58 The number of d-electrons in Fe^{2+} (At. no. 26) is not equal to that of the --
(A) p-Electrons in Ne (At. No. 10)
(B) s-Electrons in Mg (At No. 12)
(C) d-Electrons in Fe atom
(D) p-Electrons in Cl^- ion (At. No. 17)

Q.59 In an electron microscope, electron are accelerated to great velocities. Calculate the wavelength of an electron travelling with a velocity of 7.0 megameters per second . The mass of an electron is 9.1×10^{-28} g -
(A) 1.0×10^{-13}m
(B) 1.0×10^{-7}m
(C) 1.0 m
(D) 1.0×10^{-10}m

Q.60 A 200g cricket ball is thrown with a speed of 3.0×10^3 cm sec^{-1}. What will be its de Broglie's wavelength - [$h = 6.6 \times 10^{-27}$g cm^2 sec^{-1}]
(A) 1.1×10^{-32} cm
(B) 2.2×10^{-32} cm
(C) 0.55×10^{-32} cm
(D) 11.0×10^{-32} cm

Q.61 Which is the de-Broglie equation -
(A) $h = p\lambda$
(B) $h = p\lambda^{-1}$
(C) $h = \lambda p^{-1}$
(D) $h = p + \lambda$

Q.62 Which of the following has the largest de Broglie wavelength given that all have equal velocity -
(A) CO_2 molecule
(B) NH_3 molecule
(C) Electron
(D) Proton

Q.63 A ball has a mass of 0.1 kg its velocity is 40 m/s, find out de Broglie wave length -
(A) 1.66×10^{-34}m
(B) 2×10^{-34} m
(C) 3×10^{-34} m
(D) 4×10^{-34} m

Q.64 If the uncertainty of position for an electron is zero, what is the uncertainty of the momentum-
(A) Zero (B) $\hbar$ (C) h (D) Infinite

Q.65 Which of the following is the most correct expression for Heisenberg's uncertainty principle
(A) $\Delta x.\Delta p = \frac{h}{4\pi}$
(B) $\Delta x.\Delta p \geq \frac{h}{4\pi}$
(C) $\Delta x.\Delta p \leq \frac{h}{4\pi}$
(D) $\Delta x \cdot \Delta v = \frac{h}{4\pi}$

Q.66 The Heisenberg uncertainty principle can be applied to -
(A) A cricket ball
(B) A foot ball
(C) A jet aeroplane
(D) An electron

Q.67 Velocity of helium atom at 300 K 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

Q.1 The wave character of electron was experimentally verified by -
(A) de - Broglie (B) A. Einstein
(C) Germer (D) Schrodinger

Q.2 Cathode rays are -
(A) Electromagnetic waves
(B) Radiations
(C) Stream of α-particles
(D) Stream of electrons

Q.3 The e/m ratio for cathode rays -
(A) Is constant
(B) Varies as the atomic number of the element forming cathode in the discharge tube changes
(C) Varies as atomic number of the gas in the discharge tube varies
(D) Has the smallest value when the discharge tube is filled with hydrogen

Q.4 Arrange the orbitals of H-atom in the increasing order of their energy -
$3p_x$, 2s, $4d_{xy}$, 3s, $4p_z$, $3p_y$, 4s
(A) $2s < 3s = 3p_x = 3p_y < 4s = 4p_z = 4d_{xy}$
(B) $2s < 3s < 3p_x = 3p_y < 4s = 4p_z = 4d_{xy}$
(C) $2s < 3s < 3p_x = 3p_y < 4s = 4p_z = 4d_{xy}$
(D) $2s < 3s < 3p_x = 3p_y < 4s < 4p_z < 4d_{xy}$

Q.5 Electron, Proton and Neutron were respectively discovered by -
(A) James Chadwick, John Dalton, J.J. Thomson
(B) J.J. Thomson, Goldsteine, John Dalton
(C) J.J. Thomson, William Crookes, Goldsteine
(D) J.J. thomson, Goldstein , James Chadwick

Q.6 If the I.P. of Li^{+2} is 122.4 eV. Find out 6th I.P. of carbon -
(A) 122.4 × 4eV (B) 122.4 × 2eV
(C) 122.4 × 3eV (D) 122.4 × 5eV

Q.7 If W is the atomic mass and N is the atomic number of an element, the number of -
(A) Electrons = W – N (B) Neutrons = W – N
(C) Protons = W – N (D) Electrons = W

Q.8 It is known that atoms contain protons, neutrons and electrons. If the mass of neutron is assumed to be half of its original value whereas that of electron is assumed to be twice of this original value. The atomic mass of $_6C^{12}$ will be -
(A) Twice (B) 75% less
(C) 25% less
(D) One-half of its original value

Q.9 The energy difference between two electronic states is 46 .12 kcal /mole. What will be the frequency of the light emitted when an electron drops from the higher to the lower energy state (Planck' constant = 9.52 x 10^{-14} kcal sec mole $^{-1}$)
(A) 4.84 x 10^{15} cycles sec $^{-1}$
(B) 4.84 x 10^{-5} cycles sec^{-1}
(C) 4.84x 10^{-12} cycles sec^{-1}
(D) 4.84 x 10^{14} cycles sec^{-1}

Q.10 Electromagnetic radiation with maximum wave length is -
(A) Ultraviolet (B) Radio wave
(C) X - rays (D) Infrared

Q.11 Multiple of fine structure of spectral lines is due to -
(A) Presence of main energy levels
(B) Presence of sub- levels
(C) Presence of electronic configuration
(D) Is not a characteristics of the atom.

Q.12 The quantum number not obtained from the Schrodinger's wave equation is -
(A) n (B) l (C) m (D) s

Q.13 Wave mechanical model of the atom depends upon -
(A) de-Broglie concept of dual nature of electron
(B) Heisenberg uncertainty principle
(C) Schrodinger uncertainty principle
(D) All

Q.14 The correct Schrodinger's wave equation for an electron with total energy E and potential energy V is given by -
(A) $\frac{\delta^2\psi}{\delta\chi^2}+\frac{\delta^2\psi}{dy^2}+\frac{\delta^2\psi}{\delta z^2}+\frac{8\pi^2}{mh^2}(E-V)\psi=0$
(B) $\frac{\delta^2\psi}{\delta x^2}+\frac{\delta^2\psi}{dy^2}+\frac{\delta^2\psi}{dz^2}+\frac{8\pi m}{h^2}(E-V)\psi=0$
(C) $\frac{\delta^2\psi}{\delta x^2}+\frac{\delta^2\psi}{\delta y^2}+\frac{\delta^2\psi}{\delta z^2}+\frac{8\pi^2 m}{h^2}(E - V)\ \psi=0$
(D) None of the above.

Q.15 Which one of the statement of quantum numbers is false -
(A) Quantum number were proposed out of necessity in Bohr model of the atom.
(B) Knowing *n* and ℓ it possible to designated a subshell .
(C) The principal quantum number alone can give the complete energy of an electron in any atom.
(D) Azimuthal quantum number refers to the subshell to which an electron belongs and describes the motion of the electron.

Q.16 Which orbital is dumb-bell shaped -
(A) s-Orbital (B) p-Orbital
(C) d-Orbital (D) f-Orbital

Q.17 Which of the following subshell can accommodate

as many as 10 electrons -
(A) 2d (B) 3d (C) 3dxy (D) $3dz^2$

Q.18 "No two electrons in an atom can have the same set of four quantum numbers. "This principle was enunciated by -
(A) Heisenberg (B) Pauli
(C) Maxwell (D) de Broglie.

Q.19 How many spherical nodes are present in a 4s orbital in hydrogen atom -
(A) 0 (B) 1 (C) 2 (D) 3

Q.20 Minimum core charge is shown by the atom -
(A) O (B) Na (C) N (D) Mg

Q.21 I.P. of hydrogen atom is equal to 13.6 eV. What is the energy required for the process :
$He^+ + \text{energy} \longrightarrow He^{+2} + e^-$
(A) 2 × 13.6 eV (B) 1 × 13.6 eV
(C) 4 × 13.6 eV (D) None of these

Q.22 If elements with principal quantum number n > 4 is not allowed in nature, the number of possible elements would be -
(A) 60 (B) 32 (C) 64 (D) 50

Q.23 If the values of ($n + \ell$) is not > 3, then the maximum number of electron in all the orbital would be -
(A) 12 (B) 10 (C) 2 (D) 6

Q.24 It is not possible to explain the pauli's exclusion principle with the help of this atom -
(A) B (B) Be (C) C (D) H

Q.25 Uncertainty in position and momentum are equal. Uncertainty in velocity is -
(A) $\sqrt{\frac{h}{\pi}}$ (B) $\sqrt{\frac{h}{2\pi}}$ (C) Both (D) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$

Q.26 The nucleus of an atom is located at x = y = z = 0. If the probability of finding an s-orbital electron in a tiny volume around x = a, y = z = 0 is 1×10^{-5}, what is the probability of finding the electron in the same sized volume around x = z = 0, y = a -
(A) 1×10^{-5} (B) $1 \times 10^{-5} \times a$
(C) $1 \times 10^{-5} \times a^2$ (D) $1 \times 10^{-5} \times a^{-1}$

Q.27 How fast is an electron moving if it has a wavelength equal to the distance it travels in one second -
(A) $\sqrt{\frac{h}{m}}$ (B) $\sqrt{\frac{m}{h}}$
(C) $\sqrt{\frac{h}{p}}$ (D) $\sqrt{\frac{h}{2(KE)}}$

Q.28 Electron corpuscular nature is not connected with -
(A) Diffraction phenomenon
(B) Photo electric effect
(C) Compton effect
(D) Mechanical effect by cathode rays

Q.29 The correct statement(s) about Bohr's orbits of hydrogen atom is/are -
(A) $r = \left[\frac{n^2h^2}{4\pi^2me^2}\right]$
(B) K.E. of the electron = −1/2 (P. E. of the electron)
(C) Angular momentum (L) = $n\left(\frac{h}{2\pi}\right)$
(D) All the above

Q.30 In centre-symmetrical system, the orbital angular momentum, a measure of the momentum of a particle travelling around the nucleus, is quantised. Its magnitude is -
(A) $\sqrt{\ell(\ell+1)}\ \frac{h}{2\pi}$ (B) $\sqrt{\ell(\ell-1)}\ \frac{h}{2\pi}$
(C) $\sqrt{s(s+1)}\ \frac{h}{2\pi}$ (D) $\sqrt{s(s-1)}\ \frac{h}{2\pi}$

Q.31 Each orbital has a nodal plane. Which of the following statements about nodal planes are not true -
(A) A plane on which there is zero probability that the electron will be found
(B) A plane on which there is maximum probability that the electron will be found
(C) Both
(D) None

Q.32 An electron, a proton and an alpha particle have kinetic energies of 16E, 4E and E respectively. What is the qualitative order of their de Broglie wavelengths -
(A) $\lambda_e > \lambda_p = \lambda_\alpha$ (B) $\lambda_p = \lambda_a > \lambda_e$
(C) $\lambda_p < \lambda_e > \lambda_a$ (D) $\lambda_a < \lambda_e \gg \lambda_p$

Q.33 If Hund's rule is followed, magnetic moment of Fe^{2+}, Mn^+ and Cr all having 24 electrons will be in order -
(A) $Fe^{2+} < Mn^+ < Cr$ (B) $Fe^{2+} < Cr = Mn^+$
(C) $Fe^{2+} = Mn^+ < Cr$ (D) $Mn^{2+} = Cr < Fe^{2+}$

Q.34 One energy difference between the states n = 2 and n = 3 is E eV, in hydrogen atom. The ionisation potential of H atom is -
(A) 3.2 E (B) 5. 6E (C) 7.2 E (D) 13.2 E

Q.35 Magnetic moments of V(Z = 23), Cr(Z = 24), Mn(Z = 25) are x, y, z. Hence -
(A) x = y = z (B) x < y < z
(C) x < z < y (D) z < y < x

Q.36 The speed of a proton is one hundredth of the speed of light in vacuum. What is its de-Broglie wavelength? Assume that one mole of protons has a mass equal to one gram [h = 6.626×10^{-27} erg sec] -
(A) 13.31×10^{-3} Å (B) 1.33×10^{-3} Å
(C) 13.13×10^{-2} Å (D) 1.31×10^{-2} Å

Q.37 The ratio of $E_2 - E_1$ to $E_4 - E_3$ for the hydrogen atom is approximately equal to -
(A) 10 (B) 15 (C) 17 (D) 12

Q.38 Consider the following ions -
(1) Ni^{2+} (2) Co^{2+} (3) Cr^{2+} (4) Fe^{3+}
(Atomic numbers : Cr = 24, Fe = 26, Co = 27, Ni = 28)
The correct sequence of the increasing order of the number of unpaired electrons in these ions is -
(A) 1, 2, 3, 4 (B) 4, 2, 3, 1
(C) 1, 3, 2, 4 (D) 3, 4, 2, 1

Q.39 What are the values of the orbital angular momentum of an electron in the orbitals 1s, 3s, 3d and 2p-
(A) 0, 0, $\sqrt{6}$ h, $\sqrt{2}$ h (B) 1, 1, $\sqrt{4\hbar}$, $\sqrt{2\hbar}$
(C) 0, 1, $\sqrt{6\hbar}$, $\sqrt{3\hbar}$ (D) 0, 0, $\sqrt{20\hbar}$, $\sqrt{6\hbar}$

Q.40 In an atom two electron move around the nucleus in circular orbits of radii R and 4R. The ratio of the time taken by them to complete one revolution -
(A) 1 : 4 (B) 4 : 1 (C) 1 : 8 (D) 8 : 7

Q.41 A beam of electrons is accelerated by a potential difference of 10000 volts. The wavelength of the wave associated with it will be -
(A) 0.0123 Å (B) 1.23 Å
(C) 0.123 Å (D) None of these

Q.42 If the number of electrons in p-orbital are two, then which one of the following is in accordance with Hund's rule -
(A) $p_x^2\, p_y^0\, p_z^0$ (B) $p_x^0\, p_y^2\, p_z^0$
(C) $p_x^0\, p_y^0\, p_z^2$ (D) $p_x^1\, p_y^1\, p_z^0$

Q.43 If there are six energy levels in H–atom then the number of lines its emission spectrum in ultra voilet region will be -
(A) 6 (B) 5 (C) 4 (D) 3

Q.44 Magnetic moment of X^{3+} ion of 3d series is $\sqrt{35}$ BM. What is atomic number of X^{3+} ?
(A) 25 (B) 26 (C) 27 (D) 28

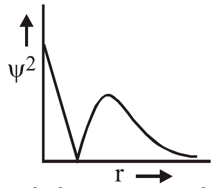
Q.45 An electron is moving with the velocity equal to 10% of the velocity of light. Its de-Broglie wave length will be -
(A) 2.4×10^{-12}cm (B) 2.4×10^{-18}cm
(C) 2.4×10^{-9}Cm (D) None of these

Q.46 Correct set of four quantum numbers for the valence (outer most) electron of rubidium (Z = 37) is -
(A) 5,0,0, $+\frac{1}{2}$ (B) 5,1,0, $+\frac{1}{2}$
(C) 5,1,1, $+\frac{1}{2}$ (D) 5,0,0, $-\frac{3}{2}$

Q.47 Ratio of time period of electron in first and second orbit of H-atom would be -
(A) 1 : 18 (B) 1 : 8
(C) 1 : 2 (D) 2 : 1

Q.48 If x is the velocity of an electron in first Bohr's orbit. What would be the velocity of the electron in third Bhor's orbit -
(A) $\frac{x}{9}$ (B) $\frac{x}{3}$ (C) 3x (D) 9 x

Q.49 The following graph between Ψ^2 probability density and distance from the nucleus represents -



(A) 2s (B) 3s (C) 1s (D) 2p

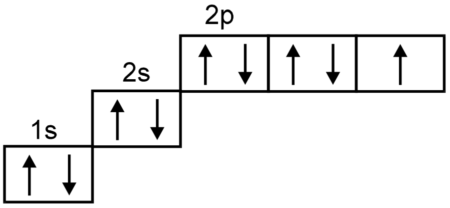
Q.50 The wave length of the last line of Paschen series would be -
(A) 9540 Å (B) 8204 Å
(C) 3650 Å (D) 912 Å

Q.51 The wavelength of X-ray is 10^{-6} cm. Its frequency in Giga Hertz (GHz) will be (Hint : 1 Giga = 10^9)
(A) 3×10^7 (B) 3×10^9
(C) 3×10^{10} (D) 3×10^8

Q.52 Which electronic level would allow the hydrogen atom to absorb a photon but not to emit a photon
(A) 3s (B) 2p (C) 2s (D) 1s

Q.53 Pauli's exclusion principle states that -
(A) Nucleus of an atom contains no nagative charge
(B) Electrons move in circular orbits around the nucleus
(C) Electrons occupy orbitals of lowest energy
(D) All the four quantum numbers of two electrons in an atom cannot be equal.

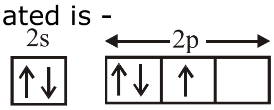
Q.54 Which element is represented by the following electronic configuration -

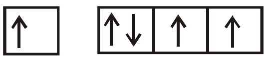


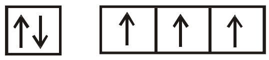
(A) Nitrogen (B) Oxygen
(C) Fluorine (D) Neon

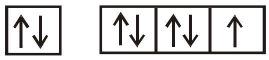
Q.55 When an electron jumps from L to K shell -
(A) Energy is absorbed
(B) Energy is released
(C) Energy is neither absorbed nor released
(D) Energy is sometimes absorbed and some times released

Q.56 The orbital diagram in which 'Aufbau principle' is violated is -

(A) 

(B) 

(C) 

(D) 

Q.57 d^6 configuration will result in total spin of -
(A) 3/2 (B) 1/2 (C) 2 (D) 1

Q.58 Bohr's model can explain -
(A) The spectrum of only hydrogen atom
(B) The spectrum of the atoms of all the elements
(C) The spectrum of only sodium atom
(D) The spectrum of atomic or ionic species having one electron.

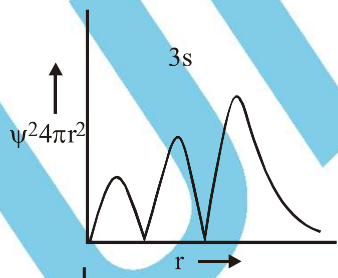
Q.59 The mass number of dispositive Zn ion is 70. The total number of neutrons is -
(A) 34 (B) 40 (C) 36 (D) 38

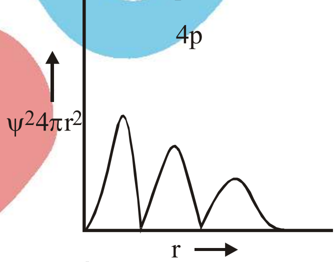
Q.60 The energy required to separate an electron from the level n = 3 of an atom is 9.69 × 10^{-19} Joules. What will be its energy in first excited state -
(A) $-\ 9 \times 9.69 \times 10^{-19}$J
(B) $-\ \dfrac{9 \times 9.69 \times 10^{-19}}{4}$ J
(C) $\dfrac{4}{9} \times 9.69 \times 10^{-19}$ J
(D) None of these

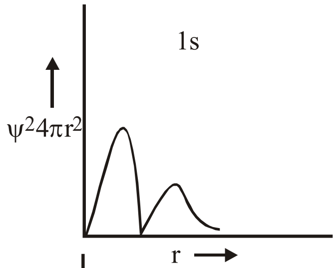
Q.61 When beryllium is bombarded with alpha particles (Chadwick's experiment) extremely penetrating radiations which can not be deflected by electrical or magnetic field are given out. These are -
(A) A beam of protons
(B) Alpha rays
(C) A beam of neutrons
(D) A beam of neutrons and protons

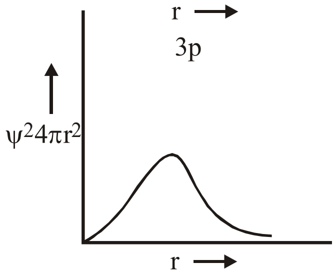
Q.62 The potential energies of first, second and third Bohr's orbits of He^+ cation are E_1, E_2 and E_3. The correct sequence of these energies is -
(A) $E_1 > E_2 > E_3$ (B) $E_1 = E_2 > E_3$
(C) $E_1 = E_2 = E_3$ (D) $E_3 > E_2 > E_1$

Q.63 Which of the following is correct radial probability distribution curve for various orbitals ?

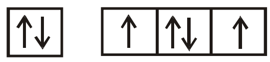
(A) 

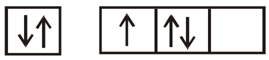
(B) 

(C) 

(D) 

Q.64 In which of the following orbital diagrams are both Pauli's exclusion principle and Hund's rule violated ?

(A) 

(B) 

(C) ↑↓ | ↑↓ | ↑↑ | ↓

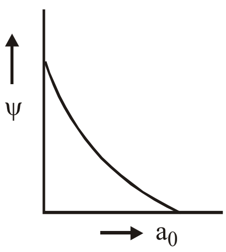
(D) ↓↓ | ↑↓ | ↑ |

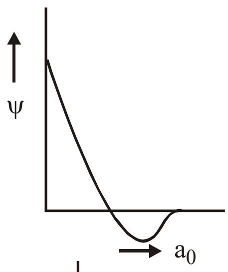
Q.65 What are the values of the orbital angular momentum of an electron in the orbitals 1s, 3s, 3d, and 2p ?

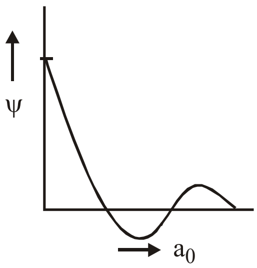
(A) 0, 0, $\sqrt{6}\hbar$, $\sqrt{2}\hbar$ (B) 1, 1, $\sqrt{4}\hbar$, $\sqrt{2}\hbar$

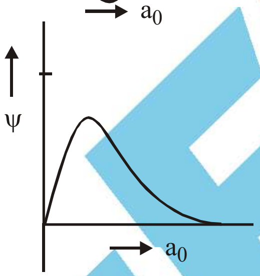
(C) 0, 1 $\sqrt{6}\hbar$, $\sqrt{3}\hbar$ (D) 0, 0, $\sqrt{20}\hbar$, $\sqrt{6}\hbar$

Q.66 Which of the following graphs correspond to one node ?

(A)


(B)


(C)


(D)


Q.67 The number of elliptical orbits excluding circular orbits in the N-shell of an atom is -

(A) 3 (B) 4 (C) 2 (D) 1

Q.68 A compound of vanadium has a magnetic moment 1.73 B. M. The electronic configuration of vanadium ion in the compound is -

(A) [Ar] $3d^2$ (B) [Ar] $3d^1$

(C) [Ar] $3d^3$ (D) [Ar] $3d^0 4s^1$

Q.69 How many lines in the spectrum will be observed when electrons return from 7th shell to 2nd shell ?

(A) 13 (B) 14 (C) 15 (D) 16

Q.70 In Ca atom how many e^- contains m = 0

(A) 12 (B) 8 (C) 10 (D) 18

Q.71 In Ne how many e^- contains m = –1

(A) 4 (B) 2 (C) 0 (D) 1

Q.72 When the wavelength of incident light on metallic plate is halved, the K.E. of emitted photoelectron will be -

(A) halved (B) doubled

(C) unchanged

(D) increased more than double

Q.73 The number of electrons in Na, having $n+\ell = 3$

(A) 4 (B) 6 (C) 7 (D) 8

Q.74 Which orbital has 1 nodal plane -

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

Q.75 How many s electron are there in Cu^+ -

(A) 2 (B) 4 (C) 6 (D) 10

Q.76 If the total energy of an electron in a hydrogen atom in excited state is –3.4 eV, then the de Broglie wavelength of the electron is -

(A) 3.3×10^{-10} cm (B) 6.6×10^{-10} cm

(C) 3.3×10^{10} cm (D) 9.3×10^{-12} cm

Q.77 The correct set of quantum numbers to the unpaired electron of fluorine atom -

(A) n = 3, ℓ = 0, m = 0

(B) n = 3, ℓ = 1, m = 1

(C) n = 2, ℓ = 0, m = 0

(D) n = 2, ℓ = 1, m = 1

Q.78 Which of the following statement is correct -

(A) Number of angular nodes = n – ℓ – 1

(B) Number of radial nodes = ℓ

(C) Total number of nodes = n – 1

(D) All

Q.79 The total energy associated per quanta with light of wavelength 600 nm -

(A) 3.3×10^{-12} erg (B) 3.3×10^{-6} erg

(C) 6.6×10^{-12} erg (D) 6.6×10^{-6} erg

Q.80 Iodine molecule dissociates into atom after absorbing light of 4500 Å. The K.E. of iodine atoms if B.E. of I_2 is 240 kJ mol^{-1} -

(A) 0.43×10^{-19} J (B) 0.216×10^{-19} J

(C) 4.3×10^{-16} J (D) 2.16×10^{-16} J

Q.81 The number of revolution/sec. made by electron in 3rd orbit of hydrogen atom -

(A) 4.88×10^{14} (B) 2.44×10^{14}

(C) 9.9×10^{14} (D) 2.44×10^{12}

Q.82 Angular and spherical nodes in 3s -

(A) 1, 1 (B) 1, 0 (C) 2, 0 (D) 0, 2

Q.83 The magnetic moment of V^{4+} ion -

(A) 1.73 (B) 1.41 (C) 3.46 (D) 2

Q.84 Which orbital represents the following set of quantum numbers n = 3, ℓ = 0, m = 0, s = +1/2 -

(A) 3p (B) 2s (C) 3s (D) 2p

Q.85 The number of unpaired electrons in Zn^{+2} -

(A) 0 (B) 1 (C) 2 (D) 3

Q.86 The uncertainly in velocity of electron having uncertainty in its position of 1Å -

(A) 5.8×10^5 m/s (B) 5.8×10^6 m/s

(C) 5.8×10^7 m/s (D) 5.8×10^8 m/s

Q.87 If ionisation energy of hydrogen atom is 13.6 eV. I.E. of Li^{+2} will be -
(A) 13.6 eV (B) 10.4 eV
(C) 40.8 eV (D) 122.4 eV

Q.88 The wavelength of third lyman series of hydrogen atom is approximately -
(A) 1×10^{-7} m (B) 1×10^{-8} m
(C) 1×10^{-6} m (D) 1×10^{-5} m

Q.89 The number of waves made by a Bohr electron in one complete revolution in its 3rd orbit -
(A) 1 (B) 2 (C) 3 (D) 4

Q.90 If potential energy of an electron in hydrogen atom is –x eV, then its kinetic energy will be -
(A) x eV (B) –x eV
(C) 2x eV (D) x/2 eV

Q.91 The number of orbitals in n = 3 are -
(A) 1 (B) 3 (C) 5 (D) 9

Passage :

Orbital is the region in an atom where the probability of finding the electron is maximum. It represents three-dimensonal motion of an electron around the nucleus. Orbitals do not specify a definite path according to the uncertainty principle. An orbital is described with the help of wave function ψ. Whenever an electron is described by a wave function, we say that an electron occupies that orbital. Since many wave functions are possible for an electron, there are many atomic orbitals in an atom. Orbitals have different shapes; except s-orbitals, all other orbitals have directional character. Number of spherical nodes in an orbital is equal to $(n-\ell-1)$. Orbital angular momentum of an electron is $\sqrt{\ell(\ell+1)}\,\hbar$.

Q.1 The nodes present in 5p orbital are -
(A) one planar, five spherical
(B) one planar, four spherical
(C) one planar, three spherical
(D) four spherical

Q.2 When an atom is placed in a magnetic field, the possible number of orientations for an orbital of azimuthal quantum number 3 is -
(A) three (B) one (C) five (D) seven

Q.3 Orbital angular momentum of *f*-electrons is-
(A) $\sqrt{2}\,\hbar$ (B) $\sqrt{3}\,\hbar$ (C) $\sqrt{12}\,\hbar$ (D) $2\hbar$

Q.4 Which of the following orbitals has/have two nodal planes ?
(A) d_{xy} (B) d_{yz}
(C) d_{xz} (D) All of these

True or False:

Q.5 The species Na^+, Mg^{2+}, Al^{3+}, O^{2-} and F^- are iso-electronic.

Q.6 The nuclear reaction ${}^{9}_{4}Be + {}^{4}_{2}He \rightarrow {}^{12}_{6}C + {}^{1}_{0}n$ was used by curie to discover neutron.

Q.7 Lyman series of hydrogen spectrum lies in the visible region.

Q.8 All the four quantum number have been derived from Schrodinger wave equation.

Q.9 The outer electronic configuration of chromium atom is $3d^4 4s^2$.

Q.10 The electron density in xy plane of $3d_{x^2-y^2}$ orbital is zero.

Q.11 All the atomic orbitals are directional in nature.

Q.12 The designation of an orbital, n = 4 and ℓ = 0 is 4s.

Q.13 Chromium atom has six unpaired electrons.

Q.14 The energies of various subshells in the same shell are in the order of s > p > d > *f*.

Each of the questions given below consist of Statement – I and Statement – II. Use the following Key to choose the appropriate answer.

(A) If both Statement- I and Statement- II are true, and Statement - II is the correct explanation of Statement– I.
(B) If both Statement - I 8 and Statement - II are true but Statement - II is not the correct explanation of Statement – I.
(C) If Statement - I is true but Statement - II is false.
(D) If Statement - I is false but Statement - II is true.

Q.15 **Statement I :** 2p orbitals do not have any spherical node.
Statement II : The number of nodes in p-orbitals is given by (n – 2) where n is the principal quantum number.

Q.16 **Statement I :** All p-orbitals have only one planar node.
Statement II : The number of radial nodes depends on the principal quantum number only.

Q.17 **Statement I :** A spectral line will be seen for a $2p_x - 2p_y$ transition.
Statement II : Energy is released in the form of waves of light when the electron drops from $2p_x$ to $2p_y$ orbital.

Q.18 **Statement I :** Hydrogen has one electron in its orbit but it produces several spectral lines.
Statement II : There are many excited energy levels available.

Q.19 **Statement I :** The 19th electron in potassium atom enters into 4 s-orbital and not the 3d-orbital.
Statement II : (n + ℓ) rule is followed for determining the orbital of the lowest energy state.

Q.20 **Statement I :** The free gaseous Cr atom has six unpaired electrons.
Statement II : Half-filled s-orbital has greater stability.

Q.21 **Statement I :** The electronic configuration of the nitrogen atom is represented as

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

and not as

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$

Statement II : The electronic configuration of the ground state of an atom is the one which has the greatest multiplicity.

Q.22 Statement I : For n = 3, ℓ = 0, 1 and 2 and m may be 0; 0, ± 1 and 0, ± 1 and ± 2.
Statement II : For each value of n, there are 0 to (n – 1) possible values of ℓ and for each value of ℓ, there are 0 to ± ℓ values of m.

Q.23 Statement I : The radial probability distribution curves of 2s, 3p, 4d and 5f orbitals are identical in shape.
Statement II : The number of planar nodes present in these orbitals is different.

Q.24 Statement I : $2p_x$, $2p_y$ and $2p_z$ each have one nodal plane.
Statement II : These orbital are degenrate orbitals.

Column Matching:

Q.25

Column-I	**Column-II**
(A) 2s	(P) Angular node = 1
(B) 1s	(Q) Radial node = 0
(C) 2p	(R) Radial node = 1
(D) 3p	(S) Angular node = 0

Q.26

Column-I	**Column-II**
(A) n = 6 $\rightarrow$ n = 3	(P) 10 lines in the spectrum
(B) n = 7 $\rightarrow$ n = 3	(Q) Spectral lines in visible region
(C) n = 5 $\rightarrow$ n = 2	(R) 6 lines in the spectrum
(D) n = 6 $\rightarrow$ n = 2	(S) Spectral lines in infrared region

SECTION : A

Q.1 An atom has a mass of 0.02 kg & uncertainity in its velocity is 9.218×10^{-6} m/s then uncertainity in position is (h = 6.626×10^{-34} J - s)
(A) 2.86×10^{-28} m (B) 2.86×10^{-32} cm
(C) 1.5×10^{-27} m (D) 3.9×10^{-10} m

Q.2 Energy of H–atom in the ground state is –13.6 eV , Hence energy in the second excited state is –
(A) –6.8 eV (B) –3.4 eV
(C) –1.51 eV (D) –4.3 eV

Q.3 Unertainty in position of a particle of 25 g in space is 10^{-5} m. Hence uncertainty in velocity (ms^{-1}) is (Planck's constant h = 6.6×10^{-34} Js)
(A) 2.1×10^{-28} (B) 2.1×10^{-34}
(C) 0.5×10^{-34} (D) 5.0×20^{-24}

Q.4 The orbital angular momentum for an electron revolving in an orbit is given by $\sqrt{l(l+1)}.\frac{h}{2\pi}$. This momentus for an s-electron will be given by -
(A) $\frac{h}{2\pi}$ (B) $\sqrt{2}.\frac{h}{2\pi}$ (C) $+\frac{1}{2}.\frac{h}{2\pi}$ (D) zero

Q.5 In Bohr series of lines of hydrogen spectrum, third line from the red end corresponds to where one of the following inter-orbit jumps of electron for Bohr orbits in an atom of hydrogen.
(A) $4 \to 1$ (B) $2 \to 5$
(C) $3 \to 2$ (D) $5 \to 2$

Q.6 The de Broglie wavelength of a tennis ball mass 60 g moving with a velocity of 10 mt. per second is approximately -
(A) 10^{-16} metres (B) 10^{-25} metres
(C) 10^{-33} metres (D) 10^{-31}metres

Q.7 Which of the following sets of quantum numbers is correct for an electron in 4f orbital ?
(A) n = 4, ℓ = 3 , m = + 4, s = + $\frac{1}{2}$
(B) n = 4, ℓ = 4 , m = – 4, s = – $\frac{1}{2}$
(C) n = 4, ℓ = 3 , m = + 1, s = + $\frac{1}{2}$
(D) n = 4, ℓ = 3 , m = – 2, s = + $\frac{1}{2}$

Q.8 Consider the ground state of Cr atom (Z = 24). The numbers of electrons with the azimuthal quantum numbers, ℓ =1 and 2 are, respectively
(A) 12 and 4 (B) 12 and 5
(C) 16 and 4 (D) 16 and 5

Q.9 The wavelength of the radiation emitted, when in a hydrogen atom electron falls from infinity to stationary state 1, would be
(Rydberg constant = 1.097×10^{7} m^{-1})
(A) 91 nm (B) 192 nm
(C) 406 nm (D) 9.1×10^{-8} nm

Q.10 Which one of the following sets of ions represents the collection of isoelectronic species?
(A) K^+, Ca^{2+}, Sc^{3+}, Cl^-
(B) Na^+, Ca^{2+}, Sc^{3+}, F^-
(C) K^+, Cl^-, Mg^{2+}, Sc^{3+}
(D) Na^+, Mg^{2+}, Al^{3+}, Cl^-
(Atomic nos.: F = 9, Cl = 17, Na = 11, Mg = 12, Al = 13, K = 19, Ca = 20 , Sc = 21)

Q.11 In a multi-electron atom, which of the following orbitals described by the three quantum members will have the same energy in the absence of magnetic and electric fields ?
(a) n = 1, ℓ = 0, m = 0 (b) n = 2, ℓ = 0, m = 0
(c) n = 2, ℓ = 1, m = 1 (d) n = 3, ℓ = 2, m = 1
(e) n = 3, ℓ = 2, m = 0
(A) (b) and (c) (B) (a) and (b)
(C) (d) and (e) (D) (c) and (d)

Q.12 Of the following sets which one does NOT contain isoelectronic species ?
(A) CN^-, N_2, C_2^{2-}
(B) PO_4^{3-}, SO_4^{2-}, ClO_4^-
(C) BO_3^{3-}, CO_3^{2-}, NO_3^-
(D) SO_3^{2-}, CO_3^{2-}, NO_3^-

Q.13 According to Bohr's theory, the angular momentum of an electron in 5th orbit is -
(A) 1.0 h/π (B) 10 h/π
(C) 2.5 h/π (D) 25 h/π

Q.14 Uncertainty in the position of an electron (mass = 9.1×10^{-31} kg) moving with a velocity 300 m/s, accurate upto 0.001 %, will be (h = 6.63×10^{-34} Js)
(A) 5.76×10^{-2} m (B) 1.92×10^{-2} m
(C) 3.84×10^{-2} m (D) 19.2×10^{-2} m

Q.15 Which of the following sets of quantum numbers represents the highest energy of an atom?
(A) n = 3, ℓ = 1, m = 1, s = +½
(B) n = 3, ℓ = 2, m = 1, s = +½

(C) $n = 4, \ell = 0, m = 0, s = +½$
(D) $n = 3, \ell = 0, m = 0, s = +½$

Q.16 The ionization enthalpy of hydrogen atom is 1.312×10^6 J mol^{-1}. The energy required to excite the electron in the atom from n = 1 to n = 2 is
(A) 6.56×10^5 J mol^{-1}
(B) 7.56×10^5 J mol^{-1}
(C) 9.84×10^5 J mol^{-1}
(D) 8.51×10^5 J mol^{-1}

SECTION : B

Q.1 What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 in the He^+ spectrum ?
(A) n= 4 to n = 1 (B) n = 3 to n = 2
(C) n = 3 to n = 1 (D) n = 2 to n= 1

Q.2 Which of the following is violation of Pauli's exclusion principle ?
(A) [↑↓] [↑↓][ ][ ]
(B) [↑↑] [↑][↑][↑]
(C) [↑↓] [↑][↓][↑]
(D) [↑] [↑][↑][↑]

Q.3 From the given sets of quantum numbers the one that is inconsistent with the theory is
(A) $n = 3 ; \ell = 2 ; m = -3 ; s = +1/2$
(B) $n = 4 ; \ell = 3 ; m = 3 ; s = +1/2$
(C) $n = 2 ; \ell = 1 ; m = 0 ; s = -1/2$
(D) $n = 4 ; \ell = 3 ; m = 2 ; s = +1/2$

Q.4 The orbital angular momentum of an electron in 2s orbital is
(A) $+\frac{1}{2} \cdot \frac{h}{2\pi}$ (B) zero
(C) $\frac{h}{2\pi}$ (D) $\sqrt{2} \cdot \frac{h}{2\pi}$

Q.5 Which of the following has maximum number of unpaired electron ?
(A) Mg^{2+} (B) Ti^{3+} (C) V^{3+} (D) Fe^{2+}

Q.6 The electrons, indentified by quantum number n and ℓ, (i) $n = 4, \ell = 1$ (ii) $n = 4, \ell = 0$ (iii) $n = 3, \ell = 2$ (iv) $n = 3, \ell = 1$ can be placed in order of increasing energy, from the lowest to highest, as
(A) (iv) < (ii) < (iii) < (i)
(B) (ii) < (iv) < (i) < (iii)
(C) (i) < (iii) < (ii) < (iv)
(D) (iii) < (i) < (iv) < (ii)

Q.7 The first use of quantum theory to explain the structure of atom was made by
(A) Heisenberg (B) Bohr
(C) Planck (D) Einstein

Q.8 For a d-electron, the orbital angular momentum is
(A) $\sqrt{6}h/2\pi$ (B) $\sqrt{2}h/2\pi$
(C) $h/2\pi$ (D) $2h/2\pi$

Q.9 The energy of an electron in the first Bohr orbit of H atom is –13.6 eV. The possible energy value (s) of the excited state(s) for electrons in Bohr orbits of hydrogen is (are)
(A) –3.4 eV (B) –4.2 eV
(C) –6.8 eV (D) +6.8 eV

Q.10 The energy of the electron in the first orbit of He^+ is -871.6×10^{-20} J. The energy of the electron in the first orbit of hydrogen would be
(A) -871.6×10^{-20} J (B) -435.8×10^{-20} J
(C) -217.9×10^{-20} J (D) -108.9×10^{-20} J

Q.11 Ground state electronic configuration of nitrogen atom can be represented by
(A) [↑↓] [↑↓] [↑][↑][↑]
(B) [↑↓] [↑↓] [↑][↓][↑]
(C) [↑↓] [↑↓] [↑][↓][↓]
(D) [↑↓] [↑↓] [↓][↓][↓]

Q.12 The electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$. This represents its
(A) excited state (B) ground state
(C) cationic form (D) anionic form

Q.13 The number of nodal planes in a p_x orbital is
(A) one (B) two (C) three (D) zero

Q.14 The wavelength associated with a golf ball weighing 200g and moving at a speed of 5 m/h is of the order
(A) 10^{-10}m (B) 10^{-20}m
(C) 10^{-30} m (D) 10^{-40}m

Q.15 The quantum numbers +1/2 and −1/2 for the electron spin represent
(A) rotation of the electron in clockwise and anticlockwise direction respectively
(B) rotation of the electron in anticlockwise and clockwise direction respectively
(C) magnetic moment of the electron pointing up and down respectively
(D) two quantum mechanical spin states which have no classical analogue

Q.16 Rutherford's experiment, which estabilished the nuclear model of the atom, used a beam of -
(A) β−particles, which impinged on a metal foil and got absorbed
(B) γ−rays, which impinged on a metal foil and ejected electrons
(C) helium atoms, which impinged on a metal foil and got scattered
(D) helium nuclei, which impinged on a metal foil and got scattered

Q.17 If the nitrogen atom had electronic configuration 1s, it would have energy lower than that of the normal ground state configuration $1s^2\ 2s^2\ 2p^3$, because the electrons would be closer to the nucleus. Yet, $1s^7$ is not observed because it violates.
(A) Heisenberg uncertainty principle
(B) Hund's rule
(C) Pauli's exclusion principle
(D) Bohr postulates of stationary orbits.

Q.18 Identify the least stable among the following :]
(A) Li^- (B) Be^- (C) B^- (D) C^-

Q.19 For which of the following the radius will be same as for hydrogen atom n = 1
(A) He^+, n = 2 (B) Li^{2+}, n =2
(C) Be^{3+}, n = 2 (D) Li^{2+} , n = 3

Q.20 The number of radial nodal surface in 3s and 2p
(A) 2, 0 (B) 2, 1
(C) 1 , 0 (D) 0, 2

Q.21 According to Bohr's theory ,
E_n = Total energy ; K_n = Kinetic energy
V_n = Potential energy r_n = Radius of n^{th} orbit
Match the following :
[IIT-2006]

Column I	**Column II**
(A) V_n/K_n = ?	(P) 0
(B) If radius of nth orbital $\propto E_n^x$, x = ?	(Q) −1
(C) Angular momentum in lowest orbital	(R) −2
(D) $\frac{1}{r^n} \propto Z^y, y = ?$	(S) 1

Q.22 Match the entries in Column-I with the correctly related quantum number(s) in Column-II.

Column-I	**Column-II**
(A) Orbital angular momentum of the electron in a hydrogen-like atomic orbital	(P) Principal quantum number
(B) A hydrogen-like one electron wave function obeying Pauli principle	(Q) Azimuthal quantum number
(C) Shape, size and orientation of hydrogen like atomic orbitals	(R) Magnetic quantum number
(D) Probability density of electron of the nucleus in hydrogen-like atom	(S) Electron spin quantum number

ANSWER

ANSWER KEY

LEVEL # 1

Q.No.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	D	C	A	C	C	B	D	C	B	B	C	B	C	B	D
Q.No.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	D	A	A	D	A	A	D	B	C	D	B	D	A	A	C
Q.No.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	A	B	B	C	C	C	A	B	D	C	A	B	B	B	A
Q.No.	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	C	C	C	C	A	D	A	C	D	D	A	B	D	B	A
Q.No.	61	62	63	64	65	66	67								
Ans.	C	B	A	C	D	B	A								

LEVEL # 2

Q.No.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	C	D	A	A	D	A	B	C	D	B	B	D	D	C	C
Q.No.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	B	B	B	D	B	C	A	A	D	D	A	A	A	D	A
Q.No.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	B	A	B	C	C	B	B	A	A	C	C	D	B	B	C
Q.No.	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	A	B	B	A	B	A	D	D	C	B	B	C	D	B	B
Q.No.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75
Ans.	C	D	A	D	A	B	A	B	C	A	C	D	C	B	C
Q.No.	76	77	78	79	80	81	82	83	84	85	86	87	88	89	90
Ans.	B	D	C	A	B	B	D	A	C	A	A	D	A	C	D
Q.No.	91														
Ans.	D														

LEVEL # 3

1. C **2.** D **3.** C **4.** D **5.** True **6.** False **7.** False **8.** False

9. False **10.** False **11.** False **12.** True **13.** True **14.** False **15.** A **16.** C

17. D **18.** A **19.** A **20.** C **21.** A **22.** A **23.** B **24.** B

25. A $\rightarrow$ P,R; B $\rightarrow$ Q,S; C $\rightarrow$ Q,P; D $\rightarrow$ P,R **26.** A $\rightarrow$ R,S; B $\rightarrow$ P,S; C $\rightarrow$ R,Q; D $\rightarrow$ P,Q

LEVEL # 4

SECTION : A

Q.No.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	A	C	A	D	B	C	C,D	B	A	A	C	D	C	B	B
Q.No.	16														
Ans.	C														

SECTION : B

Q.No.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	D	B	A	B	D	A	B	A	A	C	A,D	B	A	C	D
Q.No.	16	17	18	19	20										
Ans.	D	C	B	C	A										

21. A $\rightarrow$ R; B $\rightarrow$ Q; C $\rightarrow$ P; D $\rightarrow$ S

22. A $\rightarrow$ Q; B $\rightarrow$ P,Q,R,S; C $\rightarrow$ P,Q,R; D $\rightarrow$ P,Q