NEET Companion

For NEET and AIIMS

CHEMISTRY

Atomic Structure

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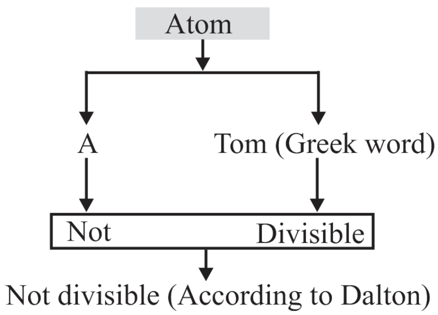
1

ATOMIC STRUCTURE

1. INTRODUCTION

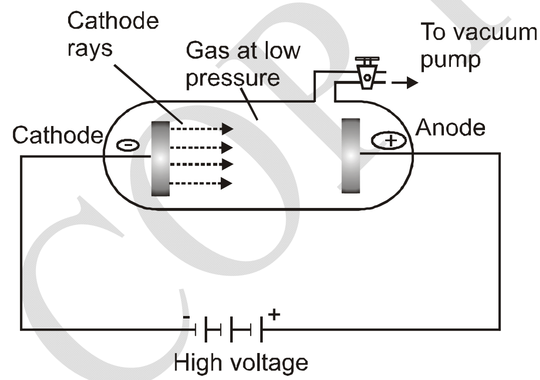


- **Atom is a Greek word**
  and its meaning is **Indivisible** i.e. an ultimate particle which cannot be further subdivided.
  John Dalton (1803 - 1808) considered that " all matter was composed of small particles called atoms.

- **Daltons Atomic Theory :-**
  This theory is based on law of mass conservation and law of definite proportions.
  The salient feature's of this theory are :-
  (1) Each element is composed of extremely small particles called atoms.
  (2) Atoms of a particular element are like but differ from atom's of other element.
  (3) Atom of each element is an ultimate particle and it has a characteristic mass but is structureless
  (4) Atom's are indestructible i.e. they can neither be created nor be destroyed.
  (5) Atom of element's take part in chemical reaction to form molecule.

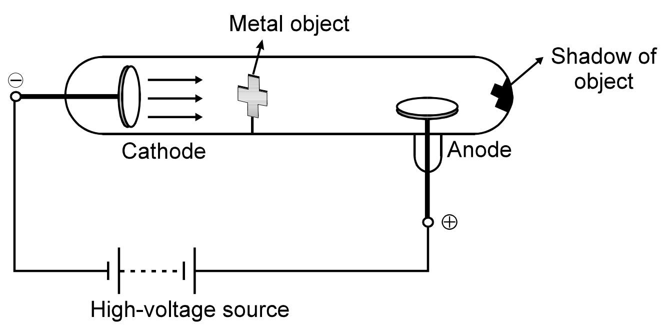
Fundamental particle's :
Atom has many types of particles i.e. electron, proton, neutron, positron, antiproton, neutrino, meson, antineutrino etc. The main fundamental particles are electron, proton, neutron.

2. CATHODE RAYS - DISCOVERY OF ELECTRON



- In 1859 Julius Plucker started the study of conduction of electricity through gases at low pressure (10–4atm) in a discharge tube 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 electrode these rays are called as cathode rays.

3. PROPERTIES OF CATHODE RAYS:



- **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 small paddle 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. The direction of deflection showed that cathode rays consist of negatively charged particles called electrons.

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

(v) Cathode rays penetrate through thin sheets of aluminium and metals.

(vi) They affect the photographic plates

(vii) The ratio of charge(e) to mass(m) i.e. charge/mass is same for all cathode rays irrespective of the gas used in the tube. $e/m = 1.76 \times 10^{11}\ Ckg^{-1}$

Thus, it can be concluded that electrons are basic constituent of all the atoms.

Some important point of Electron:

(i) The specific charge (e/m ratio) of electrons (cathode rays) was determined by Thomson as 1.76×10^{8} coulombs/gram. The specific charge of electron decreases with increase in its velocity because increase in velocity increases the mass of electron. However the specific charge of electron was found to be independent of the nature of gas and electrode used which points out that electrons are present in all atoms.

$$m = \frac{m_0}{\sqrt{1-\frac{v^2}{c^2}}},$$

where m_0 = rest mass of electron,
v = velocity of electron,
c = velocity of light

(ii) Density of electron is found to be $2.17\times10^{17}\ g/c.c$

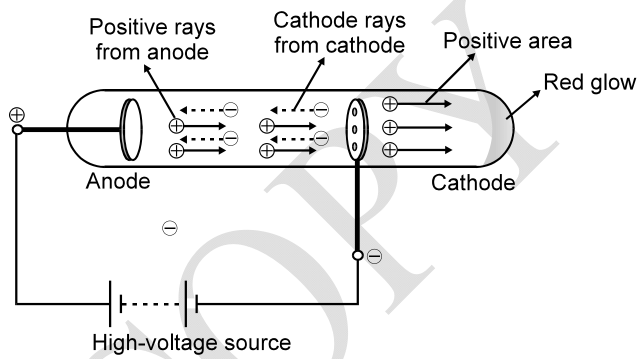
(iii) Mass of one mole of electrns $(6.023\times10^{23}$ electron) is nearly 0.55 mg.

(iv) Charge on one mole of electrons is ≈ 96500 coulombs or 1 Faraday.

(v) Mass of electron is $\frac{1}{1837}$ times that of proton.

4. PRODUCTION OF ANODE RAYS (DISCOVERY OF PROTON)

- Goldstein (1886) repeated the experiment with a discharge tube filled with a perforated cathode and found that new type of rays came out through the hole in the cathode.



When this experiment is conducted, a faint red glow is observed on the wall behind the cathode. Since these rays originate from the anode, they are called anode rays.

5. PROPERTIES OF ANODE RAYS

- ***Anode rays travel along straight paths*** and hence they cast shadows of object placed in their path.

- ***They rotate a light paddle wheel placed in their path.*** This shows that anode rays are made up of material particles.

- They are deflected towards the negative plate of an electric field. This shows that these rays are positively charged.

- For different gases used in the discharge tube, the charge to mass ratio (e/m) of the positive particles constituting the positive rays is different. When hydrogen gas is taken in the discharge tube, the e/m value obtained for the positive rays is found to be maximum. Since the value of charge (e) on the positive particle obtained from different gases is the same, the value of m must be minimum for the positive particles obtained from hydrogen gas. Thus, the positive particle obtained from hydrogen gas is the lightest among all the positive particles obtained from different gases. This particle is called the proton.

Some important point of Proton :

(i) The specific charge of a proton is 9.58×10^4 coulombs/g. However, the specific charge of the anode rays is not constant but changes with the gas in the tube. It is maximum when gas present in the discharge tube is hydrogen.

(ii) Mass of 1 mole of protons is nearly 1.007 g.

6. DISCOVERY OF NEUTRON

- Later, a need was felt for the presence of electrically neutral particles as one of the constituent of atom. These particles were discovered by Chadwick in 1932 by bombarding a thin sheet of Beryllium with α-particles, when electrically neutral particles having a mass slightly greater than that of the protons were emitted. He named these particles as neutrons.

$$^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ^{1}_{0}n$$

Some important point of Neutron:

(i) Neutron is slightly heavier (0.18%) then proton. Mass of neutron is 1.675 × 10–24g or 1.675 × 10-27 kg or 1.00866 amu.

(ii) Density of a neutron is 1.5 × 1014 g/c.c.

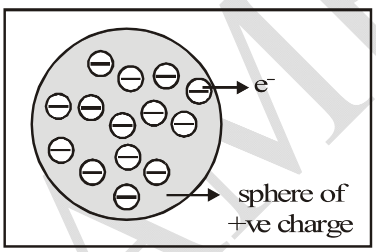
(iii) Mass of 1mole of neutrons is nearly 1.008g.

(iv) Of all the elementary particles present in an atom, neutron is the heaviest and least stable particle. Actually, isolated neutron is unstable and distintegrates into electron, proton and neutrino.

Particles	Symbol	Mass	Charge	Discoverer
Electron	$_{-1}e^0$ or β	9.1096×10^{-31} kg 0.000548 amu	-1.602×10^{-19} Coulombs -4.803×10^{-10} esu	J.J. Thompson Stoney Lorentz 1887
Proton	$_1H^1$	1.6726×10^{-27} kg 1.00757 amu	$+1.602 \times 10^{-19}$ Coulombs $+4.803 \times 10^{-10}$ esu	Goldstein Rutherford 1907
Neutron	$_0n^1$	1.6749×10^{-27} kg 1.00893 amu 1 amu $= 1.66 \times 10^{-27}$ kg	Neutral 0	James Chadwick 1932

7. ATOMIC MODELS

(A) Thomson's Model of Atom [1904]

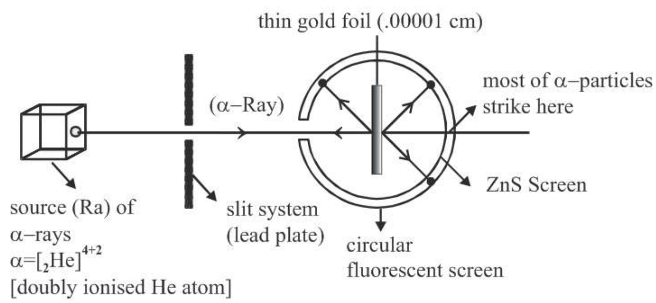


- Thomson was the first to propose a detailed model of the atom.
- Thomson proposed that an atom consists of a uniform sphere of positive.
- This model of atom is known as "Plum -Pudding Model" or "Raisin Pudding Model" or "Water Melon Model".

Drawbacks:

- An important drawback of this model is that the mass of the atoms is considered to be evenly spread over that atom.
- It is a static explain the stability of an atom.
- It couldn't explain the stability of an atom.

(B) Rutherford's Scatting Experiment

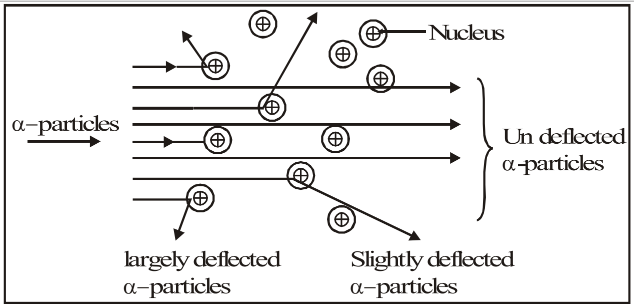


(i) Most of the α-particles passed through the gold foil undeflected.

(ii) A small fraction of the α-particles were deflected by small angles.

(iii) A very few α-particles (~1 in 20,000) bounced back, that is, were deflected by nearly 180°.

Following conclusions were drawn from the above observations-



(i) Since most of the α-particles went straight through the metal foil undeflected, it means that there must be very large empty space within the atom.

(ii) Since few of the α-particles were deflected from their original paths through moderate angles; it was concluded that whole of the occupied by this positive charge is very small in the atom.

- When α-particles come closer to this point, they suffer a force of repulsion and deviate from their paths.
- The positively charged heavy mass which occupies only a small volume in an atom is called **nucleus**. It is rigid and α-particles recoil due to direct collision with the heavy positively charged mass.

(C) Rutherford's Atomic Model

On the basis of scatting experiments, Rutherford proposed model of the atom, which is known as nuclear atomic model. According to this model-

(i) An atom consists of a heavy positively charged nucleus where all the protons and neutrons are present. Protons & neutrons are collectively referred to as nucleons. Almost whole of the mass of the atom is contributed by these nucleons. The magnitude of the +ve charge on the nucleus is different for different atoms.

(ii) The volume of the nucleus is a very small and is only a minute fraction of the total volume of the atom. Nucleus has a diameter of the order of 10^{-12} to 10^{-13} cm and the atom has a diameter of the order of 10^{-8} cm.

$$\frac{D_A}{D_N} = \frac{\text{Diameter of the atom}}{\text{Diameter of the nucleus}}$$

$$= \frac{10^{-8}}{10^{-13}} = 10^5, \; D_A = 10^5 \, D_N$$

Thus diameter (size) of the atom is 10^5 times the diameter of the nucleus.

- The radius of a nucleus is proportional to the cube root of the number of nucleons within it.

$$R \propto A^{1/3} \qquad \Rightarrow \qquad R = R_0 A^{1/3}$$

Where $R_0 = 1.33 \times 10^{-13}$ cm (a constant) and A = mass number (p + n) and R = radius of the nucleus.

$$\boxed{R = 1.33 \times 10^{-13} \times A^{1/3} \text{cm}}$$

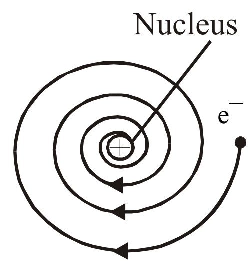
(iii) There is an empty space around the nucleus called extra nuclear part. In this part electrons are present. The number of electrons in an atoms is always equal to number of protons present in the nucleus. As the nuclear part of atom is responsible for the mass of the atom, the extra nuclear part is responsible for its volume.

The volume of the atom is about 10^{15} times the volume of the nucleus.

$$\frac{\text{Volume of the atom}}{\text{Volume of the nucleus}} = \frac{(10^{-8})^3}{(10^{-13})^3} = 10^{15}$$

(iv) Electrons revolve around the nucleus in closed orbits with high speeds. The centrifugal force acting on the revolving electrons is being counter balanced by the force of attraction between the electrons as planets.

Drawbacks of Rutherford model-



(i) This theory could not explain the stability of atom. According to Maxwell, electron loose its energy continuously in the form of electromagnetic radiations. As a result of this, the e^- should loose energy at every turn and move closer and will fall into the nucleus, thereby making the atom unstable.

(ii) If the electrons loose energy continuously, the observed spectrum should be continuous but the actual observed spectrum consists of well defined lines of definite frequencies. Hence, the loss of energy by electron is not continuous in the atom.

8. ATOMIC NUMBER AND MASS NUMBER

(a) Atomic Number

It is represent by Z. The number of protons present in the nucleus is called atomic number of an element.

For neutral atom : Number of electrons
= Number of protons

For an ion :
Number of electrons = Z – (Charge on ion)
Z = number of protons only

(b) Mass Number

It is represent by capital A. The sum of neutrons and protons is called mass number of the element. It is also known as number of nucleons because neutrons & protons are present in nucleus.

Formula :

A = number of protons + number of neutrons
Number of neutrons = A-Z

Note : A is always a whole number.

Atom

Inside the nucleus [n, p] | Outside the nucleus [e]

An atom of the element is represented by $^{A}_{Z}X$

Where, X = Symbol of element
Z = Atomic number = number of protons
= number of electrons (If atom is neutral)
A = Mass number = Number of Neutrons + Atomic number

Ex. $_{11}Na^{+}$ $(p \rightarrow 11)$ $(e \rightarrow 10)$ $_{9}F^{-}$ $(p \rightarrow 9)$ $(e \rightarrow 9 + 1 = 10)$

Ex. $^{12}_{6}C$ $(p \rightarrow 6)$ $(n \rightarrow 12 - 6 = 6)$ $(e \rightarrow 6)$ $^{16}_{8}O$ $(p \rightarrow 8)$ $(n \rightarrow 16 - 8 = 8)$ $(e \rightarrow 8)$

Method for Analysis of atomic weight:

Ex. $^{12}_{6}C$

$p \rightarrow 6$ Weight of proton $= 6 \times 1.00727$
$n \rightarrow 6$ Weight of neutron $= 6 \times 1.00866$
$e \rightarrow 6$

$$\frac{\text{Weight of neutron} = 6 \times 0.000549}{\text{Weight of } ^{12}_{6}\text{C atom} = 12.099\text{amu}}$$

Mass number of $^{12}_{6}C$ atom = 12[p and n]

Note: Mass no. of atom is always a whole number but atomic weight may be in decimal.

9. SOME IMPORTANT DEFINITIONS

(a) Isotopes: They are atoms of a given elelment which have the same atomic number but differ in their mass number.

Eg.
- $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$
- $^{16}_{8}O$, $^{17}_{8}O$, $^{18}_{8}O$
- $^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}O$

Explanation 1:

	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
$p \rightarrow$	6	6	6
$e \rightarrow$	6	6	6
$n \rightarrow$	6	7	8

[**Note**: Isotopes have the same number of protons but differ in the number of neutrons in the nucleus]

Explanation 2:

	$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$ (Radioactive element)
	Protium(H)	Deuteruim(D)	Tritium(T)
$p \rightarrow$	1	1	1
$e \rightarrow$	1	1	1
$n \rightarrow$	0	1	2

- Neutron is not available in Protium
- No. of Nucleus
 = No. of Neutrons + No. of protons
 = n + p

Atomic Weight: The atomic weight of an element is the average of mass of all the isotopes of that element.

- If an element have three isotopes y_1, y_2 and y_2 and their isotopes weights are w_1, w_2, w_3, and their percentage/possibility/ratio of occurrence in nature are x_1, x_2, x_3 respectively, then the average atomic weight of element is

Average atomic weight

$$= \frac{w_1x_1 + w_2x_2 + w_3x_3}{x_1 + x_2 + x_3}$$

Eg.	^{35}**Cl**		^{37}**Cl**
Probability	75%		25%
Ratio	3	:	1

$$\text{Average atomic weight} = \frac{35 \times 3 \times 37 \times 1}{3 + 1}$$

$$= \frac{142}{4} = 35.5$$

(b) Isobars

Isobars are the atoms of different element which have the same mass number but different atomic number i.e. they have different number of electrons, protons & neutrons but sum of number of neutrons & protons remains same.

Ex.1 $^{3}_{1}H$ $^{3}_{2}H$

p = 1	p = 2
e = 1	e = 2
n = 2	n = 1
p + n = 3	n + p = 3

Ex.2 $^{40}_{19}K$ $^{40}_{20}Ca$

p = 19	p = 20
e = 19	e = 20
n = 21	n = 20
n + p = 1	n – p = 40

(c) Isodiapheres

They are the atoms of different element which have the same difference of the number of neutrons & protons.

Ex.1 $^{11}_{5}B$ $^{13}_{6}C$

p = 5	p = 6
e = 5	e = 6
n = 6	n = 7
n – p = 1	n – p = 1

Ex.2 $^{15}_{7}N$ $^{19}_{9}F$

p = 7	p = 9
e = 7	e = 9
n = 8	n = 10
n + p = 1	n – p = 1

(d) Isotones/Isoneutronic Species/Isotonic

They are the atoms of different element which have the same number of neutrons.

Ex.1 $^{3}_{1}H$ $^{4}_{2}He$

p = 1	p = 2
e = 1	e = 2
n = 2	n = 2

Ex.2 $^{39}_{19}K$ $^{40}_{20}Ca$

p = 19	p = 20
e = 19	e = 20
n = 20	n = 20

(e) Isosters

They are the molecules which have the same number of atoms & electrons.

Ex.1 CO_2 N_2O

Atoms = 1 + 2 = 3	Atoms = 2 + 1 = 3
Electron = 6 + 8 × 2 = 22 e	Electrons= 7 × 2 + 8 = 22 e

Ex.2 CaO KF

Atoms = 2	Atoms = 2
Electron= 20 + 8 = 28 e	Electrons= 19 + 9 = 28 e

(f) Isoelectronic Species

They are the atoms, molecules or ions which have the same number of electrons.

Ex.1 **Cl^-** 18 e **Ar** 18 e

Ex.2 **H_2O** (2 + 8) = 10 e **NH_3** (7 + 3) = 10 e

Ex.3 **BF_3** (5 × 9 + 3) = 32e **SO_2** (16 + 8 × 2) = 32 e

Example-01

If no. of protons in X^{-2} is 16, then number of electrons in X^{+2} will be-

(1) 14 (2) 16 (3) 18 (4) None

Solution :

∵ No. of protons in X^{-2} is = 16

∴ No. of electrons in X^{+2} is = 14

Example-02

An element have three isotopes and their isotopic weight are 11, 12, 13 unit and their percentage of occurrence in nature is 84, 10, 5 respectively, then calculate the average atomic weight of element.

Solution :

Average Atomic weight

$$= \frac{11\times85+12\times10+13\times5}{85+10+5} = \frac{935+120+65}{100}$$

$$\text{Average weight} = \frac{1120}{100} = 11.2$$

Example-03

Average atomic weight of an element M is 51.7. If two isotopes of M are ^{50}M and ^{52}M, then calculate the percentage of occurrence of ^{50}M in nature.

Solution :

^{50}M ^{52}M

$x_1 + x_2 = 100$

$x_2 = (100 - x_1)$

Average atomic weight $= \dfrac{w_1x_1 + w_2x_2}{x_1 + x_2} = 51.7$

$= \dfrac{50 \times x_1 + 52 \times x_2}{x_1 + x_2} = 51.7$

$= \dfrac{50 \times x_1 + 52(100 - x_1)}{x_1 + x_2}$

$5170 = 50\,x_1 + 5200 - 52\,x_1$

$5170 = -2x_1 + 5200$

$2x_2 = 30$

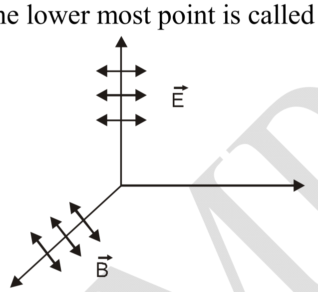
$x_1 = 15\%$ $\qquad {}^{52}M = 85\%$

10. ELECTROMAGNETIC WAVES (EM WAVES) OR RADIANT ENERGY

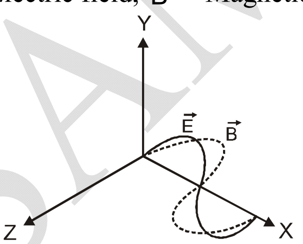
According to this theory, the energy is transmitted from one body to another in the form of waves are known as Electromagnetic waves or radiant energy.

Ex. Radio waves, micro waves, Infra-red rays, ultraviolet rays, X-rays, gamma rays.

- The radiant energy do not need any medium for propagation.
- The radiant energy have electric and magnetic fields and travel at right angle to these fields.
- The upper most point of the wave is called crest and the lower most point is called trough.



$\vec{E}$ = Electric field, $\vec{B}$ = Magnetic field



Direction of wave propagation.

Some of the terms employed in dealing with the waves are described below.

(1) Wavelength (λ) (Lambda): It is defined as the distance between two nearest crest or trough.
It is measured in terms of Å (Angstrom), pm (picometre), nm (nanometer), cm (centimeter), m (meter)

$1Å = 10^{-10}$ m, $\qquad 1pm = 10^{-12}$ m,

$1nm = 10^{-9}$ m, $\qquad 1cm = 10^{-2}$ m

(2) Wave number ($\bar{v}$)(nu bar): It is the reciprocal of the wavelength, that is number of waves present in unit length $\boxed{\bar{v} = \dfrac{1}{\lambda}}$

It is measured in terms of cm^{-1}, m^{-1} etc.

(3) Frequency (v) (nu): Frequency of a wave is defined as the number of waves which pass through a point in 1 s. It is measured in terms of Hertz(Hz), s^{-1} or cycle /s(cps) (1Hertz = $1s^{-1}$)

(4) Time period (T) : Time taken by a wave to pass through one point.

$$\boxed{T = \frac{1}{v}\,\text{second}}$$

(5) Velocity (c): Valocity of a wave is defined as distance covered by a wave in 1 second

c $= \lambda/T = \lambda v$ $\qquad$ or $\qquad$ $v = c/\lambda$

or $c = v(s^{-1}) \times \lambda(m)$ $\qquad$ or $c = v\lambda\ (ms^{-1})$

Since c is constants i.e. frequency is inversely proportional to λ

(6) Amplitude (a): The amplitude of a wave is defined as the height of crust or depth of trough.

Important note:

$$\boxed{v = \frac{c}{\lambda} = c\bar{v}} \left(\bar{v} = \frac{1}{\lambda}\right)$$

- **Order of wavelength in electromagnetic spectrum**
  Cosmic rays < γ – rays < X-rays < Ultraviolet rays < Visible < Infrared < Micro waves < Radio waves.
- **Particle Nature of Electromagnetic Radiation :**
  Some of the experimental phenomenon such as diffraction and interference can be explained by the wave nature of the electromagnetic radiation.
  However, following are some of the observations which could not be explained
  (i) the nature of emission of radiation from hot bodies (black - body radiation)
  (ii) ejection of electrons from metal surface when radiation strikes it (photoelectric effect)

Example-04

The Vividh Bharti station of All India Radio broadcasts on a frequency of 1368 Kilo hertz. Calculate the wavelength of the electromagnetic waves emitted by the transmitter.

Solution :

As we know velocity of light (c) = 3×10^8 m/s

Given v (frequency) = 1368 kHz = 1368×10^3Hz

= 10^3 Hz = 1368×10^3 s^{-1}

$$\because\ \lambda = \frac{c}{v}$$

$$\therefore\ \lambda = \frac{3\times 10^8\,ms^{-1}}{1368\times 10^3 s^{-1}} = 219.3m$$

Example-05

Calculate $\bar{v}$ in cm^{-1} and v of yellow radiation having a wavelength of 5800 Å

Solution :

A we know

$$\bar{v} = \frac{1}{\lambda} = \frac{1}{5800Å} = \frac{1}{5800\times 10^{-8} cm} = \frac{10^8}{5800} cm^{-1} = 17241.37 cm^{-1}$$

$v = c\ \bar{v} = 1\ \times 10^{10}$ cm s^{-1} $\times 10^4$ cm^{-1}

$= 3\times 1.7\times\ \times 10^{14} = 5.1\times 10^{14}$ s^{-1}

Example-06

A particular radiostation broadcast at a frequency of 1120 Kilo hertz. Another radio station broadcast at a frequency of 98.7 mega hertz. What are the wavelength of radiation from each station.

Solution :

$$\text{Station } 1^{st}\ \lambda = \frac{c}{v} = \frac{3\times 10^8 ms^{-1}}{1120\times 10^3 s^{-1}} = 267.86m$$

$$\text{Station } 2^{nd}\ \lambda = \frac{c}{v} = \frac{3\times 10^8 ms^{-1}}{98.7\times 10^7 s^{-1}} = 3.0395m$$

11. PLANCK'S QUANTUM THEORY

According to Planck's quantum theory:

(1) The radiant energy emitted or absorbed by a body not continuously but discontinuously in the form of small discrete packets of energy and these packets are called quantum.

(2) In case of light, the smallest packet of energy is called as quantum.

(3) The energy of each quantum is directly proportional to frequency of the radiation i.e.

$$E \propto v \quad \Rightarrow \quad E = hv$$

$$\text{Or } E = \frac{hc}{v} \quad \left\{\because v = \frac{c}{\lambda}\right\}$$

h is proportionality constant or Planck's constant

$h = 6.626\times 10^{-37}$ KJs or 6.626×10^{-34} Js

Or 6.626×10^{-27} ergs

(4) Total amount of energy transmitted from one body to another will be some integral multiple of energy of a quantum.

$$\boxed{E = nhv = \frac{nhc}{\lambda} = nhc\bar{v}}$$

Where n = Positive integer

= Number of quanta

Example-07

Calculate the energy of a photon of sodium light of wave length 5.862×10^{-16} m in joule.

Sol. $\lambda = 5.886\times 10^{-16}$ m, $c = 3\times 10^8 ms^{-1}$

$$E = nhv \quad \text{or} \quad \frac{nhc}{\lambda} \qquad \{\because n = 1\}$$

$$\therefore E = \frac{hc}{\lambda} = \frac{1\times 6.6\times 10^{-34} ms^{-1}}{5.862\times 10^{-16} m}$$

$$= \frac{6.6\times 3}{5.862}\times 10^{-10} J = 3.38\times 10^{-10} J$$

Example-08

Calculate the frequency & energy of a photon of wavelength 4000 Å

Solution :

(a) Calculation of frequency:

$$\lambda = 4000Å = 4000\times 10^{-10} m$$

$$\because v = \frac{c}{\lambda} = \frac{3\times 10^8 m/s}{4\times 10^{-7} m}$$

$$= 0.75\times 10^{-15} s^{-1} = 7.5\times 10^{14} s^{-1}$$

(b) Calculation of energy :

$E = hv = 6.626\times 10^{-34}$ joule second $\times 7.5\times 10^{14} s^{-1}$

$= 4.962\times 10^{-19}$ joule

Example-09

Calculate the λ and frequency of a photon having an energy of 2 electron volt

Solution :

$\because 1eV = 1.602\times 10^{-19}$ J

$\therefore 2eV = 3.204\times 10^{-19}$ J = E

Example-10

Which has a higher energy?

(a) A photon of violet light with wave length 4000 Å

(b) A photon of red light with wave length 7000 Å

Solution :

(a) Violet light : $E_{violet} = \frac{hc}{\lambda}$

$= \frac{6.626 \times 10^{-34}\,Js \times 10^{8}\,ms^{-1}}{4000 \times 10^{-10}\,m}$

$= 4.97 \times 10^{-19}$ joule

$= 6.204 \times 10^{-7}$ m

(b) Red light :

$E_{red} = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34}\,Js \times 3 \times 10^{8}\,ms^{-1}}{4000 \times 10^{-10}\,m}$

$= 2.8 \times 10^{-19}$ joule

So, $E_{violet} > E_{red}$

Example-11

Calculate number of photon coming out per sec. from the bulb of 100 watt. If it is 50% efficient and wavelength coming out is 600 nm.

Solution :

Energy = 100J

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

$= \frac{6.625 \times 10^{-34} \times 3 \times 10^{8}}{600 \times 10^{-9}}$

$= \frac{6.625}{2} \times 10–19$

Number of photon = $\frac{100}{6.625} \times 1019$

= 15.09 × 1019

Example-12

Certain sun glasses having small of AgCl incorporated in the lenses, on exposer to light of appropriate wavelength turns to gray colour to reduce the glare following the reactions:

$AgCl \xrightarrow{h\nu} Ag\,(Gray) + Cl$

If the heat of reaction for the decomposition of AgCl is 248 kJ mol^{-1}, what maximum wavelength is needed to induce the desired process?

Solution :

Energy needed to change = 248×10^{3} J/mol

If photon is used for this purpose, then according to Einstein law one molecule absorbs one photon. Therefore,

$\therefore N_A \cdot \frac{hc}{\lambda} = 248 \times 10^{3}$

$\lambda = \frac{6.625 \times 10^{-34} \times 3.0 \times 10^{8} \times 6.023 \times 10^{23}}{248 \times 10^{3}}$

$= 4.83 \times 10^{-7}$ m

12. BOHR'S ATOMIC MODEL

Some important formulae:

- This model was based on quantum theory of radiation and classical laws of physics.
- Bohr model is applicable only for single electron species like H, He^{+}, Li^{2+} etc.
- Bohr model is based on particle nature of electron.

Coulombic force $= \frac{kq_1q_2}{r^2}$

Centrifugal force $= \frac{mv^2}{r}$

Angular momentum = mvr

Important points postulates:

1st Postulate:

- Atom has a nucleus where all protons and neutrons are present.
- The size of nucleus is very small and it is present at the center of the atom.

2nd Postulate:

- Negativity charged electron revolve around the nucleus in the same way as the planets revolve around the sun.
- The path of electron is circular.
- The attraction force (Coulombic or electrostatic force) between nucleus and electron is equal to the centrifugal force on the electron.
  i.e. Attraction force towards nucleus = centrifugal force away from nucleus.

3rd Postulate:

- Electrons can revolve only in those orbits in which angular momentum (mvr) of the electron is integral multiple of $\frac{h}{2\pi}$

  i.e. $\boxed{mvr = \frac{nh}{2\pi}} = n\hbar$

  $\hbar = \frac{h}{2\pi}$

  where: n = +ve integer number (n = 1, 2, 3, 4, …..) or ($n \in I^{+}$)
  h = planck's constant
  π = Constant

- Angular momentum can have values such as $\frac{h}{2\pi}$, $2\frac{h}{2\pi}$, $3\frac{h}{2\pi}$, $4\frac{h}{2\pi}$, $5\frac{h}{2\pi}$, ………but cannot have fractional values such as $1.5\frac{h}{2\pi}$, $1.2\frac{h}{2\pi}$, $0.5\frac{h}{2\pi}$, …………

4rd Postulate:

- The orbits in which electron can revolve are known as **stationary orbits** because in these orbits energy of electron is always constant.

5th Postulate:

- Each stationary orbit is associated with definite amount of energy therefore these orbits are also called as energy levels and are numbered as 1,2,3,4,5,……. or K,L,M,N,O,…..from the nucleus outwards.

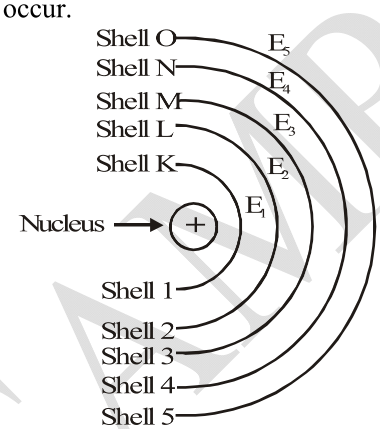
6th Postulate:

- The emission or absorption of energy in the form of photon can only occur when electron jumps from one stationary state to another & it is

$$\boxed{\Delta E = E_{higher} - E_{lower} = E_{n_2} - E_{n_2}}$$

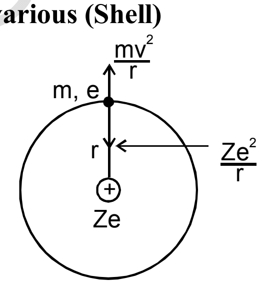
= Energy of a quantum
= $h\nu$
= Bohr's frequency condition

- Energy or absorbed when electron jumps from inner to outer orbit and is emitted when electron moves from outer to inner orbit.
- $n_2 > n_1$ whether emission or absorption of energy will occur.



13. APPLICATION OF BOHR'S MODEL

(A) Radius of various (Shell)



Columbic force = $\frac{Kq_1q_2}{r^2}$

$$= \frac{K.Ze.e}{r^2} = \frac{K.Ze^2}{r^2}$$

Where k = $9\times10^9\,Nm^2/coulomb^2$

As we know:

Coulombic force = Centrifulgal force

$$\frac{K.Ze^2}{r^2} = \frac{mv^2}{r} \quad \text{or } v^2 = \frac{KZe^2}{mr} \quad …..(1)$$

As we know: $mvr = \frac{nh}{2\pi}$ or $v = \frac{nh}{2\pi mr}$ …..(2)

putting the value of v from eq^n.(2) to eq^n.(1)

$$\left(\frac{nh}{2\pi mr}\right)^2 = \frac{KZe^2}{mr}$$

Or $$\frac{n^2h^2}{4\pi^2m^2r^2} = \frac{KZe^2}{mr}$$

$$\boxed{r = \frac{n^2h^2}{4\pi^2KZe^2}} \quad …..(3)$$

Putting the value of h, m, K, & e (Constants) in the above eq^n.(3)

$r = 0.529\times10^{-8}\times\frac{n^2}{Z}$ cm

{$1Å = 10^{-10}\,m = 10^{-8}cm$}

$$\boxed{r_n = 0.529\times\frac{n^2}{Z}Å}$$

This formula is only applicable for hydrogen and hydrogen like species i.e. species containing single electrons.

(B) Velocity of an electrons

Since coulombic force = Centrifugal force

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

or $v^2 = \frac{KZe^2}{mr}$ …….(1)

Putting the value of Angular momentum

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

or $KZe^2 = \frac{nh}{2\pi}(v)$

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

Putting the value of π, k, e & h

$$\boxed{v = 2.188\times10^6\frac{Z}{n}m/s}$$

Example-13

Calculate the radius of 1^{st}, 2^{nd}, 3^{rd}, 4^{th} Bohr's orbit of hydrogen.

Solution :

Radius of Bohr's orbit $r = 0.529 \times \frac{n^2}{Z}$ Å

(a) Radius of 1^{st} orbit :

$$r = 0.529 \times \frac{1^2}{1} \text{Å} = 0.529 \text{Å}$$

(b) Radius of 2^{nd} orbit : $r = 0.529 \times \frac{2^2}{1}$ Å

$$= 0.529 \times 4 \text{Å} = 2.116 \text{Å}$$

(c) Radius of 3^{rd} orbit :

$$r = 0.529 \times \frac{3^2}{1} \text{Å} = 0.529 \times 9 \text{Å} = 4.761 \text{Å}$$

(d) Radius of 4^{th} orbit :

$$r = 0.529 \times \frac{4^2}{1} \text{Å}$$

$$= 0.529 \times 16 \text{Å} = 8.464 \text{Å}$$

Example-14

Calculate the radius ratio of 3^{rd} & 5^{th} orbit of He^+.

Solution :

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

and Atomic Number of He = 2

$$\therefore r_3 = 0.529 \times \frac{3^2}{2} = 0.529 \times \frac{9}{2}$$

and $r_5 = 0.529 \times \frac{5^2}{2} = 0.529 \times \frac{25}{2}$

Therefore $\frac{r_3}{r_5} = \frac{0.529 \times \frac{(3)^2}{2}}{0.529 \times \frac{(5)^2}{2}} = \frac{9}{25}$

Or $r_3 : r_5 = 9 : 25$

Example-15

Calculate the radius ratio of 2^{nd} orbit of hydrogen and 3^{rd} orbit of Li^{+2}.

Solution :

Atomic number of H = 1,

Atomic number of Li = 3,

2^{nd} orbit radius of Hydrogen $(r_2)_H = 0.529 \times \frac{2^2}{1}$

3^{rd} orbit radius of Li^{+2} $(r_3)Li^+ = 0.529 \times \frac{3^2}{3}$

$$\because \frac{(r_2)_H}{(r_3)_{Li^{+2}}} = \frac{0.529 \times \frac{2^2}{1}}{0.529 \times \frac{3^2}{3}} = \frac{4}{3}$$

$$\therefore (r_2)_H : (r_2)_{Li^{+2}} = 4 : 3$$

Example-16

Calculate the radius ratio of 2^{nd} excited state of H & 1^{st} excited state of Li^{+2}.

Solution :

2^{nd} excited state, means e^- is present in 3^{rd} shell of hydrogen $r_3 = 0.529 \times \frac{(3)^2}{1} = 0.529 \times 9$

1^{st} excited state, means e^- exist in 2^{nd} shell of Li^{+2}

$$r_2 = 0.529 \times \frac{(2)^2}{3} = 0.529 \times \frac{4}{3}$$

$$\frac{\text{radius of } 2^{nd} \text{ excited state of hydrogen}}{\text{radius of } 1^{st} \text{ excited state of } Li^{+2}}$$

$$= \frac{(r_3)_H}{(r_2)_{Li^{+2}}} = \frac{0.529 \times \frac{9}{1}}{0.529 \times \frac{4}{3}} = \frac{27}{4}$$

Example-17

Calculate velocity of an electron placed in the third orbit of the hydrogen atom. Also calculate the number of revolutions per second that this electron makes around the nucleus.

Solution :

Velocity of electron in 3^{rd} orbit:

$$\because \ v_n = 2.182 \times 10^6 \times \frac{Z}{n} \text{ms}^{-1}$$

$$\therefore \ v_n = 2.182 \times 10^6 \times \frac{1}{3} \text{ms}^{-1} = 7.27 \times 10^5 \text{ms}^{-1}$$

No. of revolution per second

$$= \frac{V_n}{2\pi r_3} = \frac{V_n}{2\pi r_3} = \frac{v_n}{2\pi\left(\frac{n^2 a_0}{Z}\right)}$$

$$= \frac{7.27 \times 10^5}{2 \times 3.14 \times 9 \times 0529 \times 10^{-10}}$$

$$= 2.43 \times 10^{14} \text{ r.p.s.}$$

Example-18

How much time an e^- will take for one complete revolution in 2^{nd} orbit of He^+?

Solution :

Time taken =distance/velocity $= \dfrac{2\pi r}{v}$

$$= \frac{2\times 3.14\times 0.529\times \frac{4}{2}\times 10^{-10}\,m}{2.18\times 10^{6}\times \frac{2}{2}\,ms^{-1}}$$

$$= 3.05\times 10^{-16}\,s$$

(C) Energy of an electron

Let the total energy of an electron be E. It is the sum of kinetic and potential energy.

i.e. $E = \left(\frac{1}{2}mv^2\right) + \left(\frac{Kq_1q_2}{2}\right)$ $\left[P.E. = -\frac{KZe^2}{r}\right]$

$$E = \frac{1}{2}mv^2 + \frac{K.Ze.(-e)}{r} = \frac{1}{2}mv^2 - \frac{K.Ze.(e)}{r}$$

$$\left[KE = \frac{1}{2}mv^2 = \frac{KZe^2}{2r}\right]$$

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

Putting the value of r from eq. (3)

$$E_n = -\frac{KZe^2 \times 4\pi^2 mKZe^2}{2n^2h^2}$$

or $\boxed{E_n = -\dfrac{2\pi^2 mK^2Z^2e^4}{n^2h^{2a}}}$

Putting the value of , K, e, m, h, we get:

$$\boxed{E_n = -2.18\times\frac{Z^2}{n^2}\,J/atom}$$

or $\boxed{E_n = -1.36\times\dfrac{Z^2}{n^2}\,eV/atom}$

This formula atom & hydrogen like species i.e. single electron species. Since n can have only integral values, it follows that total energy of the e^- is quantised.

The –ve sign indicates that the electron is bonded towards nucleus.

SOME EXTRA POINTS:

(i) $K.E. = \dfrac{KZe^2}{2r}$ i.e. $K.E. \propto \dfrac{1}{r}$

On increasing radius, K.E. decreases.

(ii) $P.E. = -\dfrac{KZe^2}{r}$ i.e. $P.E. \propto -\dfrac{1}{r}$

On increasing redius, P.E. increases.

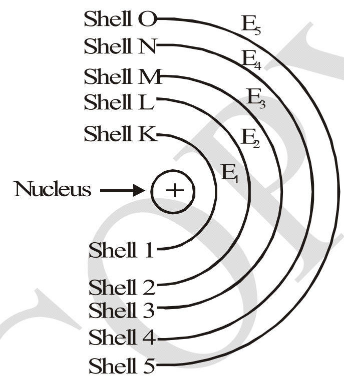
(iii) $E = -\dfrac{KZe^2}{2r}$ i.e. $E \propto -\dfrac{1}{r}$

On increasing radius, total energy increases.

Conclusion: $\boxed{P.E. = (-)2KE}$ $\boxed{K.E. = (-)E}$ $\boxed{P.E. = 2E}$

Energy difference between two energy levels:

$$E_{n_2} - E_{n_2} = -13.6\times Z^2\left[\frac{1}{n_2^2} - \frac{1}{n_1^2}\right]$$



Energy level for H atom can be represented as follows:

n = 6 or P	$E_6 = -0.38eV$
n = 5 or O	$E_5 = -0.54eV$
n = 4 or N	$E_4 = -0.85eV$
	$E_5 - E_4 = 0.31eV$
n = 3 or M	$E_3 = -1.51eV$
	$E_4 - E_3 = 0.66eV$
n = 2 or L	$E_2 = -3.4eV$
	$E_3 - E_2 = 1.89eV$
n = 1 or K	$E_1 = -13.6eV$
	$E_2 - E_1 = 10.2eV$

i.e. $(E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3) > (E_5 - E_4)$……….

Example-19

If the total energy of an electron is -1.51 eV in hydrogen atom then find out K.E., P.E., orbit, radius and velocity of the electron in that orbit.

Solution :

(i) $K.E. = -E = 1.51eV$

(ii) $P.E. = 2\times E = -2\times 1.51 = -3.02eV$

(iii) $\because E = -13.6\times\dfrac{Z^2}{n^2}eV$

$$\text{or } -1.51 = -13.6\times\frac{1^2}{n^2}eV$$

$$\Rightarrow n^2 = \frac{-13.6}{-1.51} = 9$$

$\therefore$ n = 3 i.e. 3^{rd} orbit

(iv) $r = 0.529\times\dfrac{n^2}{Z} = 0.529\times\dfrac{3\times 3}{1}$

$$= 0.529\times 9 = 4.761\text{Å}$$

(v) $v = 2.188\times10^{8}\times\frac{Z}{n}$

$= 2.188\times10^{8}\times\frac{1}{3}\,\text{cm/s}$

$= 0.729\times10^{8}\,\text{cm/s}$

Example-20

Calculate the energy of Li^{+2} ion for 2^{nd} excited state

Solution :

$E = -13.6\times\frac{Z^2}{n^2}$

$\because$ Z = 3 and electron exist in 2^{nd} excited state, means electron present in 3^{rd} shell

i.e. n = 3

$\therefore\ E = -13.6\times\frac{(3)^2}{(3)^2} = -13.6\,\text{eV/atom}$

Example-21

Calculate the ratio of energies of He^{+} for 1^{st} & 2^{nd} excited state.

Sol. $\frac{\text{Energy of }(He^{+})\ 1^{st}\text{ Excitedstate}}{\text{Energy of }(He^{+})\ 2^{nd}\text{ Excitedstate}}$

$= \frac{\text{Energy of }(He^{+})\ 2^{nd}\text{ shell}}{\text{Energy of }(He^{+})\ 3^{rd}\text{ shell}}$

$= \frac{-13.6\times\frac{(2)^2}{(2)^2}}{-13.6\times\frac{(2)^2}{(3)^2}} = \frac{9}{4}$

Example-22

The ionization energy for the hydrogen atomss is 13.6 eV then the required energy in eV to excite it from the ground state to 1^{st} excited state

Solution :

Ionization energy = 13.6 eV

i.e. Energy in ground state = -13.6eV

Energy of 1^{st} excited state

i.e. 2^{nd} orbit = 3.4eV

so, $E_2 - E_1 = 3.4 + 13.6 = 10.2\,\text{eV}$

Example-23

What is the orbit number of H atom if electron having energy is – 3.4 eV? Also report the angular momentum of electron.

Solution :

E_1 for H = – 13.6 eV

Now $E_n = \frac{E_1}{n^2}$

$\therefore\ -3.4 = \frac{-13.6}{n^2}$

$\therefore\ n = 2$

Now, Angular momentum (mvr) = $n.\frac{h}{2\pi}$

$= \frac{2\times6.626\times10^{-34}}{2\times3.14} = 2.1\times10^{-34}\ \text{J - sec}^{-1}$

Example-24

A single electron system has ionization energy 11180 kJ mol^{-1} . Find the number of protons in the nucleus of the system.

Solution :

I.E. = $\frac{Z^2}{n^2}\times 21.69\times10^{-19}$ J

$\frac{11180\times10^{3}}{6.023\times10^{23}} = \frac{Z^2}{1^2}\times21.69\times10^{-19}$

Example-25

Which state of the triply ionized Beryllium (Be3+) has the same orbit radius as that of the ground state of hydrogen atom ?

Solution :

Radius of ground state of hydrogen atom

= 0.529 Å

$0.529 = 0.529\times\frac{n^2}{Z}$

$0.529 = 0.529\times\frac{n^2}{4}$

$\therefore$ n = 2

Ans. n = 2

Example-26

The excitation energy of first excited state of a hydrogen like atom is 40.8 eV. Find the energy needed to remove the electron to form the ion.

Solution :

$40.8 = (\Delta E)_{2\to1}\times Z^2$

$\Rightarrow\ 40.8 = 10.2\times Z^2$

$\Rightarrow\ Z^2 = 4$ or $Z = 2$

IE = $13.6\,Z^2 = 13.6\times4 = 54.4$ eV

- Bohr's atomic model is applicable only for monoelectonic species like H, He^{+}, Li^{+2}, Na^{10+}, u^{91+} etc.

  $E_{z,n} = E_H\times\frac{Z^2}{n^2}$

  If z is same; $E_n = E_H\times\frac{1}{n^2}$

  If n is same; $E_z = E_H\times Z^2$

14. DEFINITION VALID FOR SINGLE ELECTRON SYSTEM

1. **Ground state (G.S.):** In any single electron species n = 1 is called ground state.
   Ground state energy of H–atom = – 13.6 ev
   Ground state energy of He^+ Ion = – 54.4 ev

2. **Excited state (E.S.):** In single electron species n>1 is called excited state.
   n = 2, called first excited state
   n = 3, called second excited state
   n = 4, called third excited state
   For the nth shell = $(n – 1)^{th}$ excited state.

3. **Excitation energy:** Energy required to excite an electron from its ground state to any excited state is called excitation energy.
   $E_2 – E_1$ = first excitation energy
   $E_3 – E_1$ = second excitation energy
   $E_4 – E_1$ = third excitation energy

4. **Binding Energy or Seperation energy :–** Energy required to move an electron from any state to $n=\infty$ is called binding energy of that state.
   Binding energy of ground state=I.E. of atom or Ion.
   $E_\infty – E_2 = 0 – E_2 = –E_2$ = first seperation energy
   $E_\infty – E_3 = 0 – E3 = –E_3$ = second seperation energy
   $E_\infty – E_4 = 0 – E_4 = –E_4$ = third seperation energy

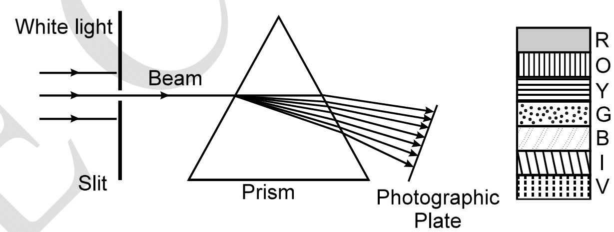
5. **Ionization energy (I.E):** Energy required to remove an electron from its ground state.
   For single electron species I.E. = $0 – E_1 = – E_1$
   For multi electron species
   Ionisation energy of H–atom = 13.6 ev
   $I.E._1 < I.E._2 < I.E._3$ ....
   Ionisation energy of He^+ ion = 54.4 ev
   Ionisation energy of Li^{+2} ion = 122.4 ev

15. HYDROGEN SPECTRUM

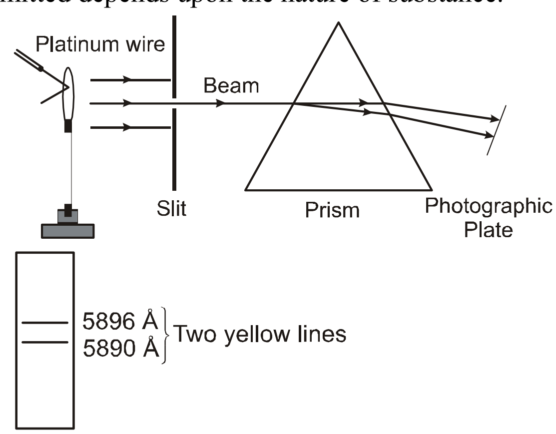
- **Study of Emission and Absorption Spectra :**
  An instrument used to separate the radiation of different wavelengths (or frequencies) is called spectroscope or a spectrograph. Photograph (or the pattern) of the emergent radiation recorded on the film is called a spectrogram or simply a spectrum of the given radiation. The branch of science dealing with the study of spectra is called spectroscopy.

- **Emission spectra :**
  When the radiation emitted from some source e.g. from the sun or by passing electric discharge through a gas at low pressure or by heating some substance to high temperature etc, is passed directly through the prism and then received on the photographic plate, the spectrum obtained is called 'Emission spectrum'.
  Depending upon the source of radiation, the emission spectra are mainly of two type:

- **Continuous spectra:**
  When white light from any source such as sun, a bulb or any hot glowing body is analysed by passing through a prism it is observed that it splits up into seven different wide band of colours from violet to red. These colours are so continuous that each of them merges into the next. Hence the spectrum is called continuous spectrum.

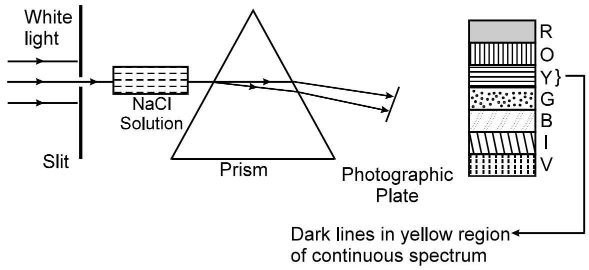


- **Line spectra:**
  When some volatile salt (e.g., sodium chloride) is placed in the Bunsen flame or an electric discharge is passed through a gas at low pressure, light emitted depends upon the nature of substance.

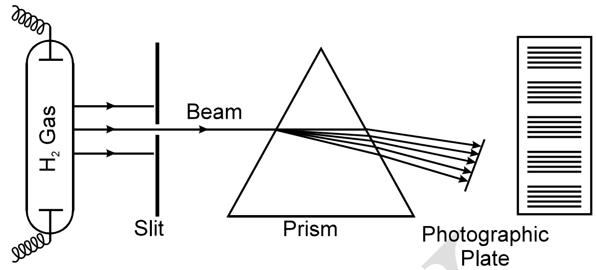


It is found that no continuous spectrum is obtained but some isolated coloured lines are obtained on the photographic plate separated from each other by dark spaces. This spectrum is called 'Line emission spectrum' or simply Line spectrum.

- **Absorption spectra:**
  When white light from any source is first passed through the solution or vapours of a chemical substance and then analysed by the spectroscope, it is observed that some dark lines are obtained in the continuous spectrum. These dark lines are supposed to result from the fact that when white light (containing radiations of many wavelengths) is passed through the chemical substance. Radiations of certain wavelengths are absorbed, depending upon the nature of the substance.

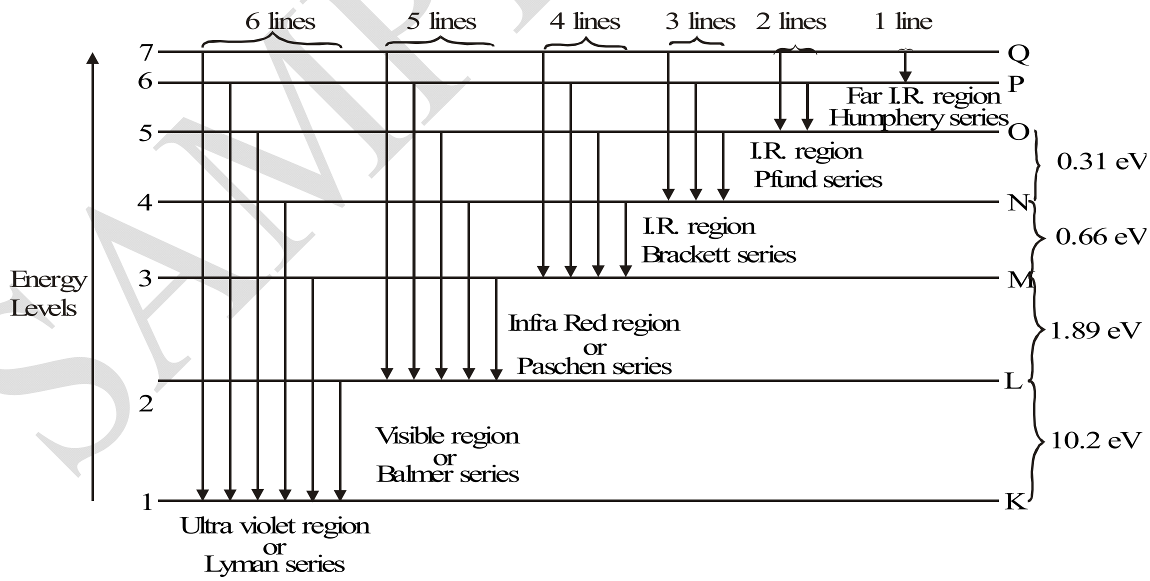


- **Emission spectrum of Hydrogen:**



When hydrogen gas at low pressure is taken in the discharge tube and the light emitted on passing electric discharge is examined with a spectroscope, the spectrum obtained is called the emission spectrum of hydrogen.

Series	**Discovered by**	**regions**	**$n_2 \rightarrow n_1$**	**Number of lines**
Lyman	Lyman	U.V. region	n_2 = 2,3,4 ... / n_1=1	$n_2 - 1$
Balmer	Balmer	Visible region	n_2 = 3,4,5 ... / n_1=2	$n_2 - 2$
Paschen	Paschen	Infra red (I.R.)	n_2 = 4,5,6 ... / n_1=3	$n_2 - 3$
Brackett	Brackett	I.R. region	n_2 = 5,6,7 ... / n_1=4	$n_2 - 4$
Pfund	Pfund	I.R. region	n_2 = 6,7,8 ... /n_1=5	$n_2 - 5$
Humphery	Humphery	Far I.R. region	n_2 = 7,8,9 ... / n_1=6	$n_2 - 6$



Similar words

- First line/Starting line/Initial line ($\lambda_{max.}$ and $v_{min.}$)

- Last line/Limiting line/Marginal line ($\lambda_{min.}$ and $v_{max.}$)
- First line of any series = α line
  Second line of any series = β line
  Third line of any series = γ line

Calculating of number of spectral lines

(a) Total number of spectral lines

$$= 1 + 2 + \ldots. (n_2 - n_1)$$

$$= \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

Where: n_2 = higher energy level;
n_1 = lower energy level
If $n_1 = 1$ (ground state)

Total number of spectral lines $= \frac{(n_2 - 1)n_2}{2}$

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

(b) Number of spectral lines which falls in a particular series = $(n_2 - n_1)$
where n_2 = higher energy level, n_1 = Fixed lower energy level of each series.

16. RYDBERG FORMULA

In 1890, Rydberg gave a very simple theoretical equation for the calculation of the wavelength of various lines of hydrogen like spectrum

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

where R = Rydberg constant = 109678 cm^{-1}

$= 109700\ cm^{-1} = 10970000\ m^{-1} = 1.1\times10^7\ m^{-1}$

$$\frac{1}{R} = 9.12\times10^{-6}\,cm = 912\ \text{Å}$$

Derivation of Rydberg formual:

$$\Delta E = E_{n_2} - E_{n_1}$$

$$\Delta E = \frac{-2\pi^2 mK^2Z^2e^4}{n_2^2h^2} - \left[\frac{-2\pi^2 mK^2Z^2e^4}{n_1^2h^2}\right]$$

$$= \frac{2\pi^2 mK^2Z^2e^4}{n_1^2} - \frac{2\pi^2 mK^2Z^2e^4}{n_2^2h^2}$$

$$\left(\because \Delta E = hv = \frac{hc}{\lambda}\right)$$

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

$$\text{or } \frac{1}{\lambda} = \frac{2\pi^2 mK^2e^4Z^2}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

where $\frac{2\pi^2 mK^2e^4}{ch^3}$ is a costant which is equal to Rydberg constant (R).

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

17. LIMITATION OF THE BOHR'S MODEL

(1) Bohr's theory does not explain the spectrum of multi electron atom.
(2) Why the angular momentum of the revolving electron is equal to $\frac{nh}{2\pi}$, has not been explained by Bohr's theory.
(3) Bohr's inter related quantum theory of relation and classical laws of physics without any theoretical explanation.
(4) Bohr's theory does not explain the fine structure of the spectral lines. Fine structure of the spectral line is obtained when spectrum is viewed by spectroscope of more resolution power.
(5) Bohr theory does not explain the splitting of spectral lines in the presence of magnetic field (Zemman's effect) of electrical field (Stark's effect).

18. DEVELOPMENT LEADING TO QUANTUM OR WAVE MECHANICAL MODEL OF ATOM

- In view of the short comings of Bohr model of atom, efforts were made to develop a new model of atom which could overcome the limitations of Bohr model. The development of new model was mainly based on following two concepts that had been put forward:
  (1) De-Broglie concept (dual nature of matter)
  (2) Heisenberg uncertainty principle.
  The new branch of science which was developed taking into consideration the above two concepts is known as quantum mechanics or wave mechanics. That is why the new model of atom is known as quantum or wave mechanical model of atom.
  Thus, just as Bohr model of atom was developed on the basis of planck's quantum theory, the quantum mechanical model of atom has been developed on the basis of quantum mechanics. It is therefore, important to first understand the above two concepts as explained below.

(A) THE DUAL NATURE OF MATTER (THE WAVE NATURE OF ELECTRON)

- In 1924, a French physical, **Louis de-Broglie** suggested that if the nature of light is both that of a particle and of a wave, then this dual behavior should be true also for the matter.

(1) The wave nature of light rays and X-rays is proved on the basis of their interference and diffraction and many facts related to radiations can only be explained when the beam of light rays is regarded as compared of energy corpuscles or photons whose velocity is $3\times10^{10}\,cm/s$.

(2) According to de-Broglie, the wavelength λ of an electron is inversely proportional to its momentum p.

$$\lambda \propto \frac{1}{p} \qquad \text{or} \qquad \lambda = \frac{h}{p}$$

(here h = Planck's constant, p = momentum of electron)

$\because$ Momentum (p) = Mass (m) $\times$ Velocity (v)

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

(3) The above relation can be proved as follows by using Einstein's equation, Planck's quantum theory and wave theory of light.
Einstein's equation, $E = mc^2$ where E is energy, m is mass of a body and c is its velocity.

$\because E = hv = h\times\frac{c}{\lambda}$ (According to planck's quantum theory) …..(i)

and $c = v\lambda$ (According to wave theory of light)

But according to Einstein's equation $E = mc^2$ …. (ii)

From equation (i) & (ii): $mc^2 = h\times\frac{c}{\lambda}$

or $mc = \frac{h}{\lambda}$ or $p = \frac{h}{\lambda}$ or $\boxed{\lambda = \frac{h}{p}}$

(4) It is clear from the above equation that the value of λ decreases on increasing either m or v or both. The wavelength of many fast-moving objects like an aeroplane or a cricket ball, is very low because of their high mass.

Bohr's theory and de-Broglie concept :

(1) According to de-Broglie, the nature of an electron moving around the nucleus is like a wave that flows in circular orbits around the nucleus.

(2) If an electron is regarded as a wave, the quantum condition as given by Bohr in his theory is readily fulfilled.

(3) If the radius of a circular orbit is r, then its circumference will be $2\pi r$.

(4) We know that according to Bohr theory,

$$mvr = \frac{nh}{2\pi} \text{ or } 2\pi r = \frac{nh}{mv}$$

($\because$ mv = p momentum)

or $2\pi r = \frac{nh}{p}$ ($\because \frac{h}{p} = \lambda$ de-Broglie equation)

$\therefore$ $2\pi r = n\lambda$ (where n = total number of waves 1, 2, 3, 4, 5, … ∞ and λ = wavelength)

(5) $\because 2\pi r = \frac{nh}{mv}$ or $mvr = \frac{nh}{2\pi}$

$\because$ mvr = Angular momentum

Thus mvr = Angular momentum, which is a integral multiple of $\frac{h}{2\pi}$.

(6) It is clear from the above description that according to de-Broglie there is similarity between wave theory and Bohr theory.

$\therefore$

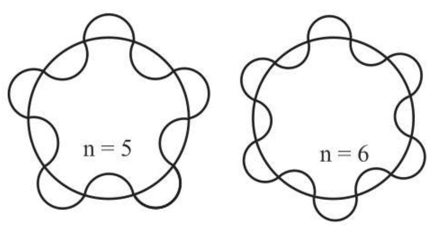


Figure : Similarity between de-Broglie waves and Bohr's orbit

(B) HEISENBERG UNCERTAINTY PRINCIPAL

- Bohr's theory considers an electron as a material particle. Its position and momentum can be determined with accuracy. But, when an electron is considered in the form of wave as suggested by de-Broglie, it is not possible to ascertain simultaneously the exact position and velocity of the electron more precisely at a given instant since the wave extends throughout a region of space.

- In 1927, Werner Heisenberg presented a principal known as Heisenberg uncertainly principle which states that : "It is impossible to measure simultaneously the exact momentum of a body as small as an electron."
- The uncertainity in measurement of position, (Δx), and the uncertainity in momentum (Δp) are related by Heisenberg's relationship as

  $$F \times \Delta t \times \Delta x \geq \frac{h}{4\pi} \text{ or } \Delta E \times \Delta t \geq \frac{h}{4\pi}$$

  where h is Planck's constant.

  (i) When $\Delta x = 0, \ \Delta v = \infty$

  (ii) When $\Delta v = 0, \ \Delta x = \infty$ So, if the position is known quite accurately, i.e., Δx is very small, Δv becomes large and vice-versa.
- **De-Broglie wavelength in terms of kinetic energy.**

  Kinetic Energy (K. E.) $= \frac{1}{2}mv^2$

  or $m \times K.E. = \frac{1}{2}m^2v^2$

  or $m^2v^2 = 2m\,K.E.$ or $mv = \sqrt{2mK.E.}$

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

  $\therefore \lambda = \frac{h}{\sqrt{2mK.E.}}$ $\left(\because mv = \sqrt{2mK.E.}\right)$

- **When a charged particle carrying Q coulomb charge is accelerated by applying potential differences of V volts, then:**

  K.E. $= Q \times V$ Joule

  But $\lambda = \frac{h}{\sqrt{2mK.E.}}$ $\therefore \lambda = \frac{h}{\sqrt{2mQV}}$

  $\boxed{\text{For electron} \left(\lambda = \sqrt{\frac{150}{V}}\,\text{Å}\right)} = \frac{12.25}{\sqrt{V}}\,\text{Å}$
- The wave nature of electron was verified experimentally by Davisson and Germer.
- de-Broglie hypothesis is applicable to macroscopic as well microscopic objects but it has no physical significance for macroscopic objects.
- Remember $\boxed{\frac{h}{4\pi} = 0.527 \times 10^{-34}\,J\,sec}$

Example-27

The mass of a particle is 1 mg and its velocity is 4.5×10^5 cm per second. What should be the wavelength of this particle if h = 6.652×10^{-27} erg second.

(1) 1.4722×10^{-24} cm (2) 1.4722×10^{-29} cm

(3) 1.4722×10^{-32} cm (4) 1.4722×10^{-34} cm

Solution :

Given that

$m = 1\text{ mg} = 1 \times 10^{-3}\,g$,

$v = 4.5 \times 10^5\,cm\,s^{-1}$,

$h = 6.652 \times 10^{-27}$ erg s.

$$\therefore \lambda = \frac{h}{mv} = \frac{6.625 \times 10^{-27}\,ergs}{1 \times 10^{-3}\,g \times 4.5 \times 10^5\,cms^{-1}}$$
$$= 1.4722 \times 10^{-29}\text{ cm}$$

Example-28

A ball weighting 25g moves with a velocity of 6.6×10^4 cm s^{-1} then find out the de-Broglie λ associated with it.

Solution :

$$\lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34} \times 10^7\,ergs}{25 \times 6.6 \times 10^4\,cms^{-1}}$$
$$= 0.04 \times 10^{-31}\,cm = 4 \times 10^{-33}\,cm$$

Example-29

If the uncertainity in position of a moving particle is 0 then find out Δp –

Solution :

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

$$\text{or } \Delta p \geq \frac{h}{4\pi \times 0} \text{ or } \Delta p \geq \infty$$

Example-30

Calculate the uncertainity in the position of a particle when the uncertainity in momentum is

(a) 1×10^{-3} g cm s^{-1}, (b) Zero

Solution :

(a) Given $\Delta p = 1 \times 10^{-3}$ g cm s^{-1},

$h = 6.62 \times 10^{-27}$ erg s, $\pi = 3.142$

According to uncertainity principle

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

$$\text{or } \Delta x \geq \frac{h}{4\pi}.\frac{1}{\Delta p} \geq \frac{6.627 \times 10^{-27}}{4 \times 3.142} \times \frac{1}{10^{-3}}$$
$$\geq 0.527 \times 10^{-24}\,cm$$

(b) When the value of $\Delta p = 0$, the value of Δx will be infinity.

Example-31

Calculate the uncertainity in velocity of a cricket ball of mass 150 g if the uncertainity in its position is of the order of 1Å ($h = 6.6\times10^{-34}$ kg m^2 s^{-1}).

Solution :

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

$$= \frac{6.6\times10^{-34}}{4\times3.143\times10^{-10}\times0.150} = 3..499\times10^{-24}\,ms^{-1}$$

19. QUANTUM NUMBERS:

- The set of four numbers required to define an electron completely in an atom are called quantum numbers. The first three have been derived from Schrodinger wave equation.
  (a) Principal quantum number (n) $\rightarrow$ shell (Orbit) (proposed by Bohr)
  (b) Azimuthal quantum number (ℓ) $\rightarrow$ Sub shell (proposed by Sommerfield)
  (c) Magnetic quantum number (m) $\rightarrow$ Orbitals (proposed by Linde)
  (d) Spin quantum number (s) $\rightarrow$ Spin of electron (proposed by Goldschmidt & uhlenbeck)

(a) Principal Quantum Number (n)

Given By $\rightarrow$ Bohr

- It represents the name and energy of the shell to which electron belongs and size of orbital.
- The value of n lies between 1 to ∞ i.e. n = 1,2,3,4,……. ∞ corresponding name of shells are K,L,M,N,O,……
- Greater the value of n, greater is the distance from the nucleus.

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

$$r_1 < r_2 < r_3 < r_4 < r_5 \ldots\ldots\ldots$$

- Greater the value of n, greater is the energy of shell.

$$E = -13.6\times\frac{Z^2}{n^2}\text{eV / atom}$$

$$E_1 < E_2 < E_3 < E_4 \ldots\ldots\ldots$$

- Velocity of electron $v = 2.18\times10^6\,\frac{Z}{n}\,m/s$

$$v_1 > v_2 > v_3 \ldots\ldots..$$

- The angular momentum of a revolving electron is $mvr = \frac{nh}{2\pi}$

  Where n = Principal quantum shell is equal to $2n^2$
- The number of electrons in a particular shell is equal to $2n^2$.

(b) Azimuthal quantum number / Angular quantum / Secondary quantum number / Subsidiary quantum number (ℓ)

Given by – Somerfield

- It represents the name of the subshell, shape of orbitals and orbital angular momentum.
- Possible values of 'ℓ' are:

i.e. ℓ = 0, 1 , 2……. (n-1)

ℓ = 0 (0 Subshell)

ℓ = 0 (p Subshell)

ℓ = 1 (d Subshell)

ℓ = 2 (f Subshell)

- Value of ℓ lies between 0 to (n – 1) in a particular nth shell:

Ex. If n = 1 then $\ell = 0 \Rightarrow$ 1s, i.e. in n = 1 shell, only one subshell 's' & 'p' are present.

- If n = 2 then $\ell = 0, 1 \Rightarrow$ 2s, 2p, i.e. in n = 2 shell, two subshell 's' & 'p' are present.
- If n = 3 then $\ell = 0, 1, 2 \Rightarrow$ 3s, 3p, 3d, i.e. in n = 3 shell, three subshell 's' , 'p' & 'd' are present.
- If n = 4 then $\ell = 0,1,2,3 \Rightarrow$ 4s, 4p, 4d, 4f i.e. in n = 4 shell, four subshell 's', 'p', 'd' & 'f' are present.
- If the value of n is same then the order of energy of the various subshell will be s < p < d < f [valid only for multi-electron species]

Ex. 1s < 2s < 3s < 4s < 5s < 6s
3d < 4d < 5d < 6d
4p < 5p < 6p

- The orbital angular momentum$=\sqrt{\ell(\ell+1)}\ \frac{h}{2\pi}$

  or $\sqrt{\ell(\ell+1)}\hbar$ $\left\{\because \hbar = \frac{h}{2\pi}\right\}$ {$\hbar$ is called as 'hash'}

Orbital angular momentum: For a subshell = 0

For p subshell = $\sqrt{2}\dfrac{h}{2\pi}$ or $\sqrt{2}h$

- The number of electrons in a particular subshell is equal to $2(2\ell+1)$

  for s subshell number of electrons = 2 e^-
  for p subshell number of electrons = 6 e^-
  for d subshell number of electrons = 10 e^-
  for f subshell number of electrons = 14 e^-
- Shape of the orbital:
  S $\rightarrow$ spherical
  p $\rightarrow$ dumb bell shape
  d $\rightarrow$ double dumb bell shape
  f $\rightarrow$ complex shape

(c) Magnetic Quantum Number/ Orientation Quantum Number (m):
Given by linde

- It represents the orientation of electron cloud (orbital).
- Under the influence of magnetic field each subshell is further subdivided into orbitals (The electron cloud is known as orbital)
- Value of m = all integral value from $-\ell$ to $+\ell$ including zero.

  i.e. Value of m = $-\ell$ to $+\ell$

Orbital: 3D space around the nucleus where the probability of finding electrons is maximum is called an orbital. An orbital can be represented by 3 set of quantum numbers = $\psi_{n,l,m}$

Ex.1: $2p_x$; n = 2, ℓ = 1, m = –1 or m = +1

Ex.2: $3d_{z^2}$; n = 3, ℓ = 2, m = 0

Ex.3: $\psi_{(3,2,0)}$; n = 3, ℓ = 2, m = 0; $3d_{z^2}$;

Note: It is point / line / plane / surface in which probability of finding electron is zero.

Total number of nodes = $\boxed{n-1}$

They are of 2 types:

(i) Radial nodes / spherical nodes / Nodal surface

$\boxed{\text{number of radial node} = n-\ell-1}$

(ii) Angular nodes / Nodal planes

$\boxed{\text{number of angular nodes/nodal planes} = \ell}$

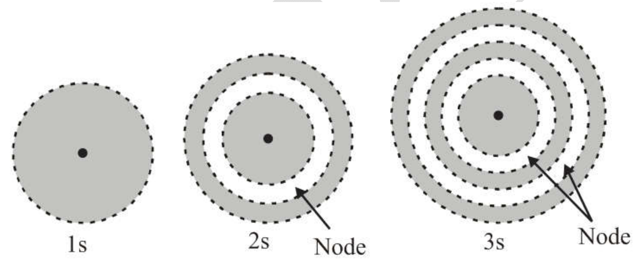
- Nucleus and ∞ (infinite) are not considered as node.

Types of orbital:

Case-I: If ℓ = 0 then m = 0, it implies that s subshell has only one orbital called as s orbital.

Shape of s-orbitals:

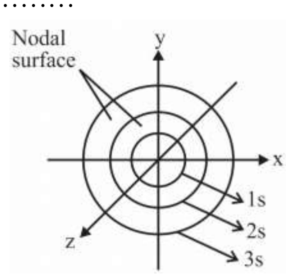
The s-orbitals are spherical about the nucleus, i.e., the probability of finding electron is same in all directions from the nucleus. The size of the orbitals depends on the value of principle quantum number. The 1s orbital is smaller than 2s-orbitals and 2s-orbital is smaller than 3s, but all are spherical in shape as shown in figure.



Although the s-orbitals belonging to different shells are spherically symmetrical about the neucleus, yet they differ in certain respects as explained below:

(i) The probability of finding 1s electron is found to be maximum near the nucleus and decreases as the nucleus and then deceases to zero as the distance from the nucleus increases, The intermediate region (a spherical shell) where the probability is zero is called a nodal surface or simply node. Thus, 2s-orbital differs from 1s-orbital in having one node within it. Similarity, 3s has two nodes. In general, any ns orbital has (n – 1) nodes.

(ii) The size and energy of the s-orbital increases as the principal quantum number increases, i.e., the size and energy of s-orbital increases in the order 1s < 2s < 3s ………



The s orbital of higher energy levels are also symmetrically spherical and can be represented as above.

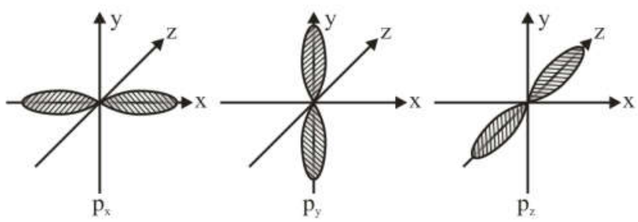
Case-II: If $\ell = 1$ (p-Subshell) then

m=	–1	0	+1
	p_x	p_z	p_y

It implies that, p subshell have three orbital called as p_x, p_y, p_z.

Shape of p-orbitals:

- There are three p-orbitals, commonly referred to as p_x, p_y, and p_z. These three p-orbitals, possess equivalent energy and therefore, have same relation with the nucleus. They, differ in their direction & distribution of the charge.



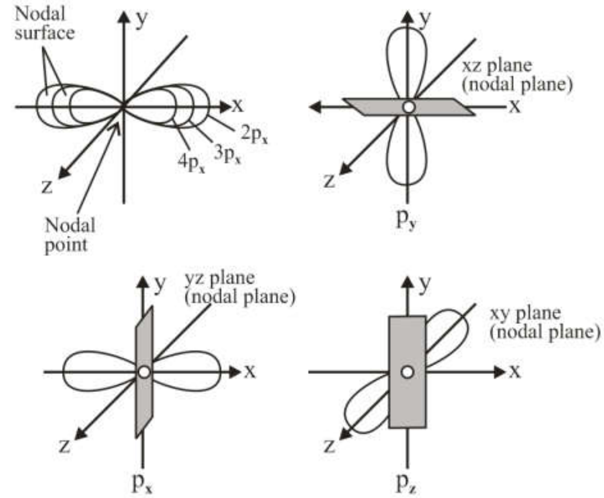
- These three p-orbitals are situated at right angle to one another and are directed along x, y and z axis (figure)
- Each p orbital has dumb bell shape (2 lobes which are which are separated from each other by a point by a point of zero probability called nodal point or node or nucleus).
- The two lobes of each orbital are separated by a plane of zero electron density called nodal plane.
- Each p orbital of higher energy level are also dumb bell shape but they have nodal surface.

Nodal surface:

Orbital	Nodal surface
$3p_x$	1
$4p_x$	2
np_x	(n–2)

Nodal plane:

Orbital	Nodal plane
p_x	yz plane
p_y	xz plane
p_z	xy plane

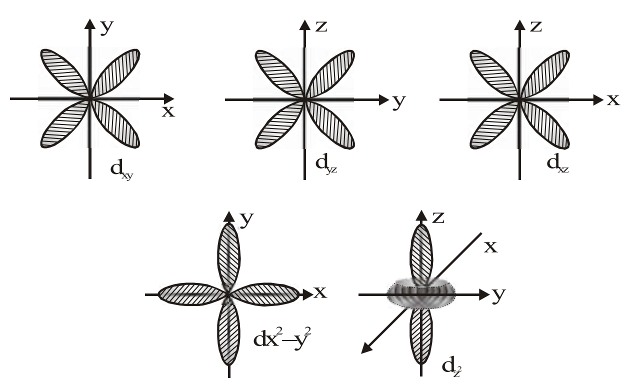


Case III When $\ell = 2$, 'm' has values -2, -1, 0, +1, +2. It implies that d subshell of any energy shell has five orbitals. All the five orbitals are not identical in shape. Four of the d orbitals d_{xy}, d_{yz}, d_{xz}, $d_{x^2-y^2}$ contain four lobes while fifth orbital d_{z^2} consists of only two lobes. The lobes of d_{xy} orbital lie between x and y axes. Similar is the case for d_{yz} and d_{xz}. Four lobes of d_{z^2} orbital are lying along z axis and contain a ring of negative charge surrounding the nucleus in xy plane. Geometry of d orbital is double dumbbell.

m =	–2	–1	0	+1	+2
	d_{xy}	d_{yz}	d_{z^2}	d_{xz}	d_{x2-y2}

Shape of d-orbits

It implies that d subshell has 5 orbitals i.e., five electron cloud and can be represented as follows



Each d-orbital of higher energy level also have double dumbbell shape but they have nodal surface.

In d-orbitals:

(i) Nodal Point → 1

(ii) Nodal surface → 3 d_{xy} → 0 Nodal surface
4 d_{xy} → 1 Nodal surface
5 d_{xy} → 2 Nodal surface
n d_{xy} → (n – 3)

Number of nodal surface = n – ℓ – 1

(iii) Nodal plane:
d_{xy} → xz & yz nodal plane:
d_{xz} → xy & zy nodal plane:
d_{zy} → zx & yx nodal plane:
$d_{x^2-y^2}$ → 2, nodal plane:
d_{z^2} → 0, nodal plane:

Note: Orbitals of d subshell are equivalent in energy.

(d) Spin Quantum number (s):
Given by **Goudsmit** and **Unlenbeck**

- It represents the direction of electron spin around its own axis.

  For clockwise spin/spin up(↑) electron $\pm\frac{1}{2}$
- For anti-clockwise spin/spin down(↓) electron $\pm\frac{1}{2}$

  Spin angular momentum of an electron

  $$= \sqrt{s(s+1)}.\frac{h}{2\pi} \text{ or } \sqrt{s(s+1)}.h$$
- Each orbital can accommodate 2 electrons with opposite spin or spin paired.

  Correct $\boxed{\uparrow\downarrow}$ Spin paired e^-

  Wrong $\boxed{\uparrow\uparrow}$ Spin parallel e^-
- Maximum spin of atom = number of unpaired e^-

Example-32

Calculate the value of n, ℓ and m for $7p_y$ orbital?

Solution :
n = 7, ℓ = 1, m = +1 or –1

Example-33

Calculate the value of n, ℓ and m for 3s orbital?

Solution :
n = 3, ℓ = 0, m = 0

Example-34

Calculate the value of n, ℓ and m for $5d_{z^2}$ orbital?

Solution :
n = 5, ℓ = 2, m = 0

Example-35

Which of the following set of quantum numbers is not possible?

(a) n = 2, ℓ = 0, m = –1, s = $-\frac{1}{2}$

(b) n = 3, ℓ = 2, m = 0, s = $\pm\frac{1}{2}$

(c) n = 2, ℓ = 3, m = –2, s = $\pm\frac{1}{2}$

Solution :
(a) not possible
(b) possible
(c) not possible

20. RULES FOR FILLING OF ELECTRONS

(a) Pauli's exclusion principle
(b) (n+ ℓ) rule
(c) Aufbau Principle
(d) Hund's maximum multiplicity principle

(a) Pauli's Exclusion Principal
In 1925 Pauli started that no two electron in an atom can have same values of all four quantum numbers.
i.e., an orbital can accommodate maximum 2 electrons with opposite spin.

Ex.1 $^{12}_{6}C$ →

		$1s^2$	$2s^2$	$2p^2$
		$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}\boxed{\uparrow}\boxed{}$
n	=	1	2	2
ℓ	=	0	0	1
m	=	0	0	+1, 0, –1
s	=	$+\frac{1}{2},-\frac{1}{2}$	$+\frac{1}{2},-\frac{1}{2}$	$+\frac{1}{2},+\frac{1}{2}$

Ex.2 $_{17}Cl$ →

		$1s^2$	$2s^2$	$2p^6$
		$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$
n	=	1	2	2
ℓ	=	0	0	1
m	=	0	0	+1, 0, –1
s	=	$+\frac{1}{2},-\frac{1}{2}$	$+\frac{1}{2},-\frac{1}{2}$	$+\frac{1}{2},-\frac{1}{2}+\frac{1}{2},-\frac{1}{2}+\frac{1}{2},-\frac{1}{2}$

		$3s^2$	$3p^5$
		$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$
n	=	3	3
ℓ	=	0	0
m	=	0	+1, 0, –1
s	=	$+\frac{1}{2},-\frac{1}{2}$	$+\frac{1}{2},-\frac{1}{2}+\frac{1}{2},-\frac{1}{2}+\frac{1}{2}$

(b) (n + ℓ) Rule (For multi electron species)

According to it the sequences in which various subshell are filled up can also be determined with the help of (n + ℓ) value for a given subshell.

- **Principal of (n + ℓ) rule:**

  The subshell with lowest (n + ℓ) value is filled up first. When two or more subshell have same (n + ℓ) value then the subshell with lowest value of n is filled up first.

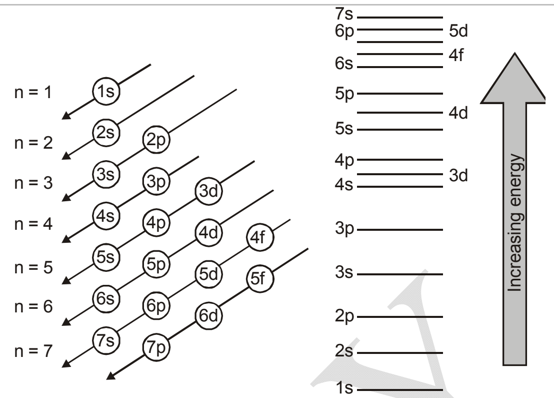
- **In case of H-atom:**
  Energy only depends on principal quantum number
  1s < 2s = 2p < 3s = 3p = 3d < 4s = 4d = 4f <….

Sub Shell	n	ℓ	n+ℓ	
1s	1	0	1	
2s	2	0	2	
2p	2	1	3	(1)
3s	3	0	3	(2)
3p	3	1	4	(1)
4s	4	0	4	(2)
3d	3	2	5	(1)
4p	4	1	5	(2)
5s	5	0	5	(3)
4d	4	2	6	(1)
5p	5	1	6	(2)
6s	6	0	6	(3)

Order: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^{26}, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^2, 5f^{214}, 6d^{10}$,……

(c) Aufbau Principal
Aufbau is a German word and its meaning is 'Building up'

- Aufbau principal gives a sequence in which various subshell are filled up depending on the relative order of the energies of various subshell.
- Principle: The subshell with minimum energy is filled up first when this subshell obtained maximum quota of electrons then the next subshell of higher energy starts filling.
- The sequence in which various subshell are filled are as follows.



Order: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^{26}, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^2, 5f^{214}, 6d^{10}$,……

For Example
Electronic Configuration

$_1H \rightarrow 1s^1$
$_2He \rightarrow 1s^2$
$_3Li \rightarrow 1s^2, 2s^1$
$_4Be \rightarrow 1s^2, 2s^2$
$_5B \rightarrow 1s^2, 2s^2, 2p^1$
$_6C \rightarrow 1s^2, 2s^2, 2p^2$
$_7N \rightarrow 1s^2, 2s^2, 2p^3$
$_8O \rightarrow 1s^2, 2s^2, 2p^4$
$_9F \rightarrow 1s^2, 2s^2, 2p^5$
$_{10}Ne \rightarrow 1s^2, 2s^2, 2p^6$
$_{11}Na \rightarrow 1s^2, 2s^2, 2p^6, 3s^1$
$_{12}Mg \rightarrow 1s^2, 2s^2, 2p^6, 3s^2$
$_{13}A\ell \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$
$_{14}Si \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$
$_{15}P \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$
$_{16}S \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$
$_{17}Cl \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
$_{18}Ar \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$
$_{19}K \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$
$_{20}Ca \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$
$_{21}Sc \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^1$
$_{22}Ti \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^2$
$_{23}V \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^3$
$_{24}Cr \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^5$ [Exception]
$_{25}Mn \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^5$
$_{26}Fe \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$
$_{27}Co \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^7$
$_{28}Ni \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^8$

$_{29}Cu \rightarrow$ $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$ [Exception]

$_{30}Zn \rightarrow$ $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}$

Electronic configuration can be written by following different methods:

$_{26}Fe \rightarrow$ (1) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$

(2) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^6, 4s^2$

(3) $1s^2$, $2s^2p^6$, $3s^2p^6d^6$, $4s^2$

2 8 14 2

(4) [Ar] $4s^2$ $3d^6$

$$_{26}Fe \rightarrow 1s^2\ \underbrace{2s^2\ 2p^2}_{(n-2)}\ \underbrace{3s^2\ 3p^6\ 3d^6}_{(n-1)}\ \underbrace{4s^2}_{(n)}$$

n $\rightarrow$ Outer mnost Shell or Ultimate Shell or Valence

In this Shell electrons are called as Valence electrons or this is called core charge

(n – 1) $\rightarrow$ Penultimate Shell or core or pre valence Shell

(n – 2) $\rightarrow$ Pre Penultimate Shell

If we remove the last 'n' Shell (ultimate Shell) then the remaining shells are collectively called as Kernal.

Ex. $_{26}Fe \rightarrow \underbrace{1s^2\ 2s^2\ 2p^2\ 3s^2\ 3p^6\ 3d^6}_{\text{Kernel}}\ 4s^2$

- **Exception of Aufbau principle:**

  In some cases it is seen that the electronic configuration is slightly different from the arrangement given by Aufbau principle. A simple reason behind this is that half filled & full filled subshell have got extra stability.

Ex.1 $_{24}Cr \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$4s^2$ [↑↓] $3d^4$ [↑][↑][↑][↑][ ]

(wrong configuration)

$_{24}Cr \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$4s^1$ [↑] $3d^5$ [↑][↑][↑][↑][↑]

(correct configuration)

Ex.2 $_{29}Cu \rightarrow$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$4s^2$ [↑↓] $3d^9$ [↑↓][↑↓][↑↓][↑↓][↑]

(wrong configuration)

$_{29}Cu \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$4s^1$ [↑] $3d^{10}$ [↑↓][↑↓][↑↓][↑↓][↑↓]

(correct configuration)

(d) Hund's Maximum Multiplicity Rule

(Multiplicity: Many of the same kind)

- This rule deals with the filling of electrons into the orbitals belonging to the same subshell (that is, orbitals of equal energy, called degenerate orbitals).
- It states: pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.
- Since there are three p, five d and seven f orbitals, therefore, the pairing of electrons will start in the p, d and f orbitals with the entry of 4^{th}, 6^{th} and 8^{th} electron, respectively.

Ex. $_5B \rightarrow 1s^2$ [↑↓] $2s^2$ [↑↓] $2p^1$ [↑][ ][ ]

$_6C \rightarrow 1s^2$ [↑↓] $2s^2$ [↑↓] $2p^2$ [↑][↑][ ]

$_7N \rightarrow 1s^2$ [↑↓] $2s^2$ [↑↓] $2p^3$ [↑][↑][↑]

$_8O \rightarrow 1s^2$ [↑↓] $2s^2$ [↑↓] $2p^4$ [↑↓][↑][↑]

$_9F \rightarrow 1s^2$ [↑↓] $2s^2$ [↑↓] $2p^5$ [↑↓][↑↓][↑]

$_{10}Ne \rightarrow 1s^2$ [↑↓] $2s^2$ [↑↓] $2p^6$ [↑↓][↑↓][↑↓]

Example-36

Calculate the number of unpaired electrons in Cr

Solution :

$_{24}Cr \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

in $_{24}Cr$, 6 electrons are unpaired.

Example-37

The number of unpaired electrons in Cr^{+3}

Solution :

$Cr^{+3} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^3$

in Cr^{+3}, 3 electrons are unpaired.

Example-38

The number of unpaired electronics in 3d subshell of Cr^{+3}

Solution : 3

Example-39

The number of unpaired electronics in Fe^{+2} & Fe^{+3}

Solution :

$Fe^{+2} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^6$

= 4 unpaired electrons

$Fe^{+3} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^5$

= 5 unpaired electrons

EXERCISE # 1

Based On Calculation related to nucleus

1. The mass charge ratio for A^+ ion is 1.97×10^{-7} kg C^{-1}. Calculate the mass of A atom.
(1) 9.1×10^{-26} (2) 3.1×10^{-25}
(3) 3.16×10^{-26} (4) 9.1×10^{-25}

2. Atomic radius is of the order of 10^{-8} cm and nuclear radius is of the order of 10^{-13} cm. Calculate what fraction of atom is occupied by nucleus ?
(1) 10^{-20} (2) 10^{-15}
(3) 10^{-12} (4) None

3. Atomic weight of an element is not necessarily whole number because
(1) it contains electrons, protons and neutrons
(2) it contains allotropic forms
(3) atoms are no longer considered indivisible
(4) it contains isotopes

4. The charge on the atom having 17 protons, 18 neutrons and 18 electrons is
(1) + 1 (2) – 1
(3) – 2 (4) zero

5. Which of the following are isoelectronic with one another ?
(1) Na^+ and Ne (2) K^+ and O
(3) Ne and O (4) Na^+ and K^+

6. A neutral atom (atomic No. > 1) consists of :
(1) Only protons
(2) Neutrons + protons
(3) Neutrons + electrons
(4) Neutrons + electrons + protons

7. Millikan's oil drop expriments is used to find -
(1) e/m ratio of an electron
(2) Charge on an electron
(3) Mass of an electron
(4) Velocity of an electron

8. An element having atomic number 25 and atomic weight 55 will have –
(1) 25 protons and 30 neutrons
(2) 25 neutrons and 30 protons
(3) 55 protons
(4) 55 neutrons

9. Which of the following is isoelectronic with N_2O
(1) NO (2) N_2O_5 (3) CO_2 (4) CO

10. The fraction of volume occupied by the nucleus with respect to the total volume of an atom is
(1) 10^{-15} (2) 10^{-5} (3) 10^{-30} (4) 10^{-10}

11. Charge on a positron is equal to that of :
(1) Proton (2) Electron
(3) Nucleon (4) Neutron

12. The e/m is not constant for:
(1) Cathode rays (2) Positive rays
(3) α-rays (4) β-rays

13. Which one of the following pairs represents isobars
(1) ${}_2He^3$ and ${}_2He^4$
(2) ${}_{12}Mg^{24}$ and ${}_{12}Mg^{25}$
(3) ${}_{19}K^{40}$ and ${}_{19}K^{39}$
(4) ${}_{19}K^{40}$ and ${}_{18}Ar^{40}$

Based On Quantum theory of light and photoelectric effect

14. AIR service on Vividh Bharati is transmitted on 219 m band. What is its transmission frequency in Hertz?
(1) 1.3×10^6 Hz (2) 1.9×10^6 Hz
(3) 1×10^6 Hz (4) 6.5×10^6 Hz

15. If 10^{-17}J of light energy is needed by the interior of human eye to see an object. The number of photons of green light (λ = 550 nm) needed to see the object are :
(1) 27 (2) 28 (3) 29 (4) 30

16. Which of the following statements is false:
(1) The energy of red photon is more than the energy of violet photon
(2) The momentum of photon is inversely proportional to its wave length
(3) The energy of a photon is inversely proportional to its wave length
(4) The particle nature of electromagnetic radiations is able to explain the photoelectric effect

17. Light of wavelength λ falls on metal having work function hc/λ_0. Photoelectric effect will take place only if :
(1) $\lambda \geq \lambda_0$ (2) $\lambda \geq 2\lambda_0$
(3) $\lambda \geq \lambda_0$ (4) $\lambda \geq \lambda_0/2$

18. Calculate the number of proton emitted in 10 hours by a 60 W sodium lamp (λ or photon = 5893 Å)
(1) 6.4×10^{24} (2) 7.2×10^{24}
(3) 2.1×10^{19} (4) 3.3×10^{19}

19. A photon in X region is more energetic than in the visible region; X is:
(1) IR (2) UV
(3) Microwave (4) Radio wave

20. Electromagnetic radiations of wavelength 242 nm is just sufficient to ionise Sodium atom. Then the ionisation energy of Sodium in KJ mole^{-1} is
(1) 494.65 (2) 400
(3) 247 (4) 600

21. A 1 kW radio transmitter operates at a frequency of 880 Hz. How many photons per second does it emit:
(1) 1.71×10^{21} (2) 1.71×10^{33}
(3) 6.02×10^{23} (4) 2.85×10^{26}

22. A bulb of 40 W is producing a light of wavelength 620 nm with 80% of efficiency then the number of photons emitted by the bulb in 20 seconds are (1eV=1.6×10^{-19} J, hc = 12400 eV Å)
(1) 2×10^{18} (2) 10^{18}
(3) 10^{21} (4) 2×10^{21}

23. Which one of the following is not the characteristic of Planck's quantum theory of radiation-
(1) The energy is not absorbed or emitted in whole number multiple of quantum.
(2) Radiation is associated with energy.
(3) Radiation energy is not emitted or absorbed continously but in the form of small packets called quanta.
(4) This magnitude of energy associated with a quantum is proportional to the frequency.

24. Calculate the energy of a photon of sodium light of wave length 5.826×10^{-16} m in Jules.
(1) $\lambda=3.41\times10^{-12}$ J (2) $\lambda=3.41\times10^{-10}$ J
(3) $\lambda=3.41 \times 10^{-15}$ J (4) $\lambda=3.41\times10^{-20}$ J

25. Calculate the frequency of a photon of wavelength 4000 Å
(1) 7.5×10^{14} s^{-1} (2) 7.5×10^{-16} s^{-1}
(3) 8×10^{-14} s^{-1} (4) 6.5×10^{-15} s^{-1}

26. Calculate the wavelength of a photon having an energy of 2 electron volt
(1) 6.204×10^{-7} m (2) 6.206×10^{-6} m
(3) 6.204×10^{-9} m (4) 6.204×10^{-8} m

27. How many photons of light having a wavelength of 5000 Å are necessary to provide 1 joule of energy.
(1) 2.8×10^{18} photons (2) 2.5×10^{17} photons
(3) 2.5×10^{18} photons (4) 2.6×10^{14} photons

Based On | Bohr Model

28. 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:
(1) 84.2×10^{-18} J/atom
(2) 44.10×10^{-18} J/atom
(3) 63.2×10^{-18} J/atom
(4) 21.2×10^{-18} J/atom

29. Energy required to pull out an electron from 1st orbit of hydrogen atom to infinity is 100 units. The amount of energy needed to pull out the electron from 2nd orbit to infinity is :
(1) 50 units (2) 100 units
(3) 25 units (4) Zero

30. The ionization energy of H-atom is 13.6 eV. The ionization energy of Li^{+2} ion will be :
(1) 54.4 eV (2) 122.4 eV
(3) 13.6 eV (4) 27.2 eV

31. Which of the following electron transition in a hydrogen atom will require the largest amount of energy ?
(1) From n = 1 to n = 2 (2) From n=2 to n = 3
(3) From n = ∞ to n = 1 (4) From n=3 to n = 5

32. The energy of an orbit in a hydrogen atom is given by the relation, $E = \dfrac{\text{Constant}}{n^2}$ (kJ mol^{-1})
Which of the following properties represents the constant in the above relation ?
(1) Electron affinity (2) Ionization energy
(3) Electronegativity (4) Bond energy

33. What is likely to be orbit number for a circular orbit of diameter 20 nm of the hydrogen atom if we assume Bohr orbit to be the same as that represented by the principal quantum number?
(1) 10 (2) 14
(3) 12 (4) 16

34. If velocity of an electron in I orbit of H atom is V, what will be the velocity of electron in 3rd orbit of Li^{+2}

(1) V (2) V/3
(3) 3 V (4) 9 V

35. The species which has its fifth ionisation potential equal to 340 V is

(1) B^+ (2) C^+
(3) B (4) C

36. Match the following

(1) Energy of ground state of He^+ — (i) + 6.04 eV
(2) Potential energy of I orbit of H-atom — (ii) –27.2 eV
(3) Kinetic energy of II excited state of He^+ — (iii) 54.4 V
(4) Ionisation potential of He^+ — (iv) – 54.4 eV

(1) A – (i), B – (ii), C – (iii), D – (iv)
(2) A – (iv), B – (iii), C – (ii), D – (i)
(3) A – (iv), B – (ii), C – (i), D – (iii)
(4) A – (ii), B – (iii), C – (i), D – (iv)

37. **S_1 :** Bohr model is applicable for Be^{2+} ion.
S_2 : Total energy coming out of any light source is integral multiple of energy of one photon.
S_3 : Number of waves present in unit length is wave number.
S_4 : e/m ratio in cathode ray experiment is independent of the nature of the gas.

(1) F F T T (2) T T F F
(3) F T T T (4) T F F F

38. **S_1 :** Potential energy of the two opposite charge system increases with the decrease in distance.
S_2 : When an electron make transition from higher orbit to lower orbit it's kinetic energy increases.
S_3 : When an electron make transtition from lower energy to higher energy state its potential energy increases.
S_4 : 11eV photon can free an electron from the 1st excited state of He^+ -ion.

(1) T T T T (2) F T T F
(3) T F F T (4) F F F F

39. If r_1 is the radius of the first orbit of hydrogen atom, then the radii of second, third and fourth orbits in terms of r_1 are :

(1) r_1^2, r_1^3, r_1^4 (2) $8r_1, 27r_1, 64r_1$
(3) $4r_1, 9r_1\ 16r_1$ (4) $2r_1, 6r_1, 8r_1$

40. Bohr's model can explain :

(1) The spectrum of hydrogen atom only
(2) The spectrum of atom or ion containing one electron only
(3) The spectrum of hydrogen molecule only
(4) The solar spectrum

41. Calculate the de-Broglie wave length of the electrone in the ground state of hydrogen atom , given that its kinetic energy is 13.6 eV ($1eV = 1.602\times10^{-19}J$)

(1) 3.328×10^{-10} m (2) 2.338×10^{-10} m
(3) 3.328×10^{10} m (4) 2.338 x 10m

Based On Spectrum

42. An atom has x energy level then total number of lines in its spectrum are :

(1) 1 + 2 + 3 .... (x + 1)
(2) 1 + 2 + 3 .... $(x)^2$
(3) 1 + 2 + 3 .... (x – 1)
(4) (x + 1) (x + 2) (x + 4)

43. If the series limit of wavelength of the Lyman series for the hydrogen atoms is 912Å, then the series limit of wavelength for the Balmer series of the hydrogen atom is :

(1) 912 Å (2) 912 × 2 Å
(3) 912 × 4 Å (4) 912/2 Å

44. The shortest wave length in H spectrum of Lymen series when $R_H = 109678\ cm^{-1}$ is

(1) 1002.7 Å (2) 1215.67 Å
(3) 1127.30 Å (4) 911.7 Å

45. No. of visible lines when an electron returns from 5th orbit to ground state in H spectrum :

(1) 5 (2) 4
(3) 3 (4) 10

46. In hydrogen spectrum which of the following lies in the wavelength range 350 –700 nm ?

(1) Balmer series (2) Lyman series
(3) Brackett series (4) Paschen series

47. According to Bohr's theory, the angular momentum for an electron in 5th orbit is

(1) 2.5 h/π (2) 5 h/π
(3) 25 h/π (4) 5π /2h

48. Transition of an electron from n = 3 to n = 1 level results in

(1) emission spectrum (2) band spectrum
(3) infrared spectrum (4) X-ray spectrum

49. If r is the radius of first orbit, the radius of n^{th} orbit of H atom is given by -
(1) r n (2) $r n^2$ (3) r/n (4) $r^2 n^2$

50. The radius of hydrogen in ground state is 0.53 Å. In normal state the radius of Li^{2+} (atomic number = 3) in ground state will be :
(1) 1.06 Å (2) 0.265 Å
(3) 0.17 Å (4) 0.53 Å

51. The minimum energy required to excite a hydrogen atom from its ground state is -
(1) 3.4 eV (2) 13.6 eV
(3) – 13.6 eV (4) 10.2 eV

52. The separation energy of the electron present in the shell n = 3 is 1.51 eV. What is the energy in the first excited state –
(1) –1.51 eV (2) –3.4 eV
(3) +1.51 eV (4) +3.4 eV

53. The wavelength of a spectral line for an electronic transition is inversely related to :
(1) number of electrons undergoing transition
(2) the nuclear charge of the atom
(3) the velocity of an electron undergoing transition
(4) the difference in the energy levels involved in the transition

54. In a sample of H-atom electrons make transition from 5^{th} excited state to ground state, producing all possible types of photons, then number of lines in infrared region are
(1) 4 (2) 5
(3) 6 (4) 3

55. Calculate wavelength of 3^{rd} line of Bracket series in hydrogen spectrum
(1) $\frac{784}{33R}$ (2) $\frac{33R}{784}$
(3) $\frac{784R}{33}$ (4) $\frac{33}{784R}$

Based On | De-broglie wavelength & Uncertainity principle

56. What possibly can be the ratio of the de Broglie wavelengths for two electrons each having zero initial energy and accelerated through 50 volts and 200 volts ?
(1) 3 : 10 (2) 10 : 3
(3) 1 : 2 (4) 2 : 1

57. 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]:
(1) 13.31 × 10^{-3} Å (2) 1.33 × 10^{-3} Å
(3) 13.13 × 10^{-2} Å (4) 1.31 × 10^{-2} Å

58. An α-particle is accelerated through a potential difference of V volts from rest. The de-Broglie's wavelength associated with it is
(1) $\sqrt{\frac{150}{V}}$ Å (2) $\frac{0.286}{\sqrt{V}}$ Å
(3) $\frac{0.101}{\sqrt{V}}$ Å (4) $\frac{0.983}{\sqrt{V}}$ Å

59. The uncertainity in position and velocity of a particle are 10^{-10} m and 5.27 × 10^{-24} ms^{-1} respectively. Calculate the mass of the particle (h = 6.625 × 10^{-34} Joule sec.)
(1) 0.099 Kg (2) 0.089 Kg
(3) 0.99 Kg (4) Can not predict

60. It the uncertainity in position of a moving particle is 0 then find out ΔP
(1) 0 (2) 1
(3) ∞ (4) Can not predict

61. The de Broglie equation suggests that an electron has
(1) Particle nature
(2) Wave nature
(3) Particle-wave nature
(4) Radiation behaviour

62. The Uncertainity in the momentum of an electron is 1.0×10^{-5} kg ms^{-1} . The Uncertainity in its position will be: (h = 6.626 × 10^{-34} Js)
(1) 1.05 × 10^{-28} m (2) 1.05 × 10^{-26} m
(3) 5.27 × 10^{-30} m (4) 5.25 × 10^{-28} m

63. Which of the following has least de Broglie λ
(1) e^- (2) p
(3) CO_2 (4) SO_2

64. A helium molecule is moving with a velocity of 2.40 x 10^2 ms^{-1} at 300k. The de-Broglie wave length is about
(1) 0.416 nm (2) 0.83 nm
(3) 803 Å (4) 8000 Å

65. In H-atom, if 'x' is the radius of the first Bohr orbit, de Broglie wavelength of an electron in 3^{rd} orbit is:
(1) 3 π x (2) 6 π x
(3) $\frac{9x}{2}$ (4) $\frac{x}{2}$

66. The wavelength of a charged particle ________the square root of the potential difference through which it is accelerated :
(1) is inversely proportional to
(2) is directly proportional to
(3) is independent of
(4) is unrelated with

67. Which of the following should be the wavelength of an electron its mass is 9.1×10^{-31} Kg and its velocity 1/10 of that of light and the value of h is 6.6252×10^{-24} joule second?
(1) 2.446×10^{-7} m (2) 2.246×10^{-9}m
(3) 2.246×10^{-11} m (4) 2.246×10^{-13} m

68. A ball weight 25 g moves with a velocity of 6.6×10^4 cm/sec then find out the de Broglie wavelength.
(1) 0.4×10^{-33} cm (2) 0.4×10^{-31} cm
(3) 0.4×10^{-30} cm (4) 0.4×10^{20} cm

69. Calculate the uncertainity in velocity of a cricket ball of mass 150 g if the uncertainity in its position is of the order of 1 Å ($h = 6.6 \times 10^{-34}$ Kg m^2 s^{-1})
(1) 3.499×10^{-24} ms^{-1}
(2) 3.499×10^{-21} ms^{-1}
(3) 3.499×10^{-20} ms^{-1}
(4) 3.499×10^{-21} ms^{-1}

Based On | Quantum numbers & Electronic configuration

70. The orbital angular momentum of an electron in 2s-orbital is -
(1) $\frac{h}{mv}$ (2) zero
(3) $\frac{h}{2\pi}$ (4) $\sqrt{2}\ \frac{h}{2\pi}$

71. Which of the following set of quantum numbers are permitted
(1) n = 3, l = 2, m = – 2, s = +1/2
(2) n = 3, l = 2, m = – 1, s = 0
(3) n = 2, l = 2, m = +1, s = – 1/2
(4) n = 2, l = 2, m = +1, s = – 1/2

72. A given orbital is labeled by the magnetic quantum number m = –1. This could not be
(1) s - orbital (2) p-orbital
(3) d-orbital (4) f-orbital

73. Magnetic quantum number specifies -
(1) Size of orbitals (2) Shape of orbitals
(3) Orientation of orbitals (4) Nuclear stability

74. For the energy levels in an atom which one of the following statements is correct :
(1) The 4s sub-energy level is at a higher energy than the 3d sub-energy level
(2) The second principal energy level can have four orbitals and contain a maximum of 8 electrons
(3) The M-energy level can have maximum of 32 electrons
(4) None of these

75. Which of the following represents the correct set of quantum numbers of a 4d electron ?
(1) 4, 3, 2, $+\frac{1}{2}$ (2) 4, 2, 1, 0
(3) 4, 3, – 2, $+\frac{1}{2}$ (4) 4, 2, 1, $-\frac{1}{2}$

76. A p-orbital can accommodate
(1) 4 electrons
(2) 6 electrons
(3) 2 electrons with parallel spins
(4) 2 electrons with opposite spins

77. An orbital containing electron having quantum number n = 4, l = 3, m = 0 and s = $-\frac{1}{2}$ is called
(1) 3s orbital (2) 3p orbital
(3) 4d orbital (4) 4f orbital

78. The maximum number of electrons in a subshell is given by the expression
(1) $4l - 2$ (2) $4l + 2$
(3) $2l + 2$ (4) $2n^2$

79. The electrons present in K-shell of the atom will differ in
(1) principal quantum number
(2) azimuthal quantum number
(3) magnetic quantum number
(4) spin quantum number

80. Magnetic moment of X^{n+} (Z = 26) is $\sqrt{24}$ B.M. Hence number of unpaired electrons and value of n respectively are:
(1) 4, 2 (2) 2, 4
(3) 3, 1 (4) 0, 2

81. Which of the following statements is wrong :

(1) Kinetic energy of an electron is half of the magnitude of its potential energy

(2) Kinetic energy of an electron is negative of total energy of electron

(3) Energy of an electron decreases with increases in the value of principal quantum number

(4) All of these

82. For an electron, with n = 3 has only one radial node. The orbital angular momentum of the electron will be

(1) 0 (2) $\sqrt{6}\frac{h}{2\pi}$

(3) $\sqrt{2}\frac{h}{2\pi}$ (4) $3\left(\frac{h}{2\pi}\right)$

83. Which of the following pair having same number of orbitals :

(a) N, O (b) O, F

(c) Na, K (d) S, Cl

The correct answer is :

(1) a, b, c (2) b, c, d

(3) c, d, a (4) a, b, d

84. The maximum number of 3d-electrons having spin quantum number s = +1/2 are -

(1) 10 (2) 14

(3) 5 (4) None of these

85. In which (n + ℓ) rules not applicable -

(1) Cu, Cr (2) Cu, Zn

(3) Ag, Zn (4) All of these

EXERCISE # 2

1. For $\ell = 1$, n = 3 the corresponding orbitals are -

(1) s, p_x, p_y (2) s, p_z, p_y
(3) s, p_x, d_{xy} (4) p_x, p_y, p_z

2. If the energy of an electron in hydrogen atom is given by expression, $-1312/n^2$ kJ mol^{-1}, then the energy required to excite the electron from ground state to second orbit is

(1) 328 kJ/mol (2) 656 kJ/mol
(3) 984 kJ/mol (4) 1312 kJ/mol

3. The ionization energy of H atom is 13.6 eV what will be ionization energy of He^+ and Li^{+2} ions-

(1) - 54.4 ev and - 12.2 ev
(2) 122.4 ev and 55.4 ev
(3) 54.4 ev and 122.4 ev
(4) 12.1 ev and 13.6 ev

4. According to classical theory if an electron is moving in a circular orbit around the nucleus-

(1) It will continue to do so for sometime.
(2) Its orbit will continuously shrink.
(3) Its orbit will continuously enlarge.
(4) It will continue to do so for all the time.

5. The increasing order (lowest first) for the value of e/m (charge/mass) is for -

(1) e, p, n, α (2) n, p, e, α
(3) n, p, α, e (4) n, α, p, e

6. Which orbital is non-directional

(1) s (2) p (3) d (4) All

7. Uncertainity in position is twice the Uncertainity in momentum. Uncertainity in velocity is :

(1) $\sqrt{\frac{h}{\pi}}$ (2) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$
(3) $\frac{1}{2m}\sqrt{\hbar}$ (4) $\frac{h}{4\pi}$

8. If n and ℓ are respectively the principal and azimuthal quantum numbers, then the expression for calculating the total number of electrons in any orbit is -

(1) $\sum_{\ell=1}^{\ell=n} 2(2\ell+1)$ (2) $\sum_{\ell=1}^{\ell=n-1} 2(2\ell+1)$
(3) $\sum_{\ell=0}^{\ell=n+1} 2(2\ell+1)$ (4) $\sum_{\ell=0}^{\ell=n-1} 2(2\ell+1)$

9. If wavelength is equal to the distance travelled by the electron in one second, then -

(1) $\lambda = \frac{h}{p}$ (2) $\lambda = \frac{h}{m}$
(3) $\lambda = \sqrt{\frac{h}{p}}$ (4) $\lambda = \sqrt{\frac{h}{m}}$

10. The wave number of electromagnetic radiation emitted during the transition of electron in between two levels of Li^{2+} ion whose principal quantum numbers sum is 4 and difference is 2 is :

(1) 3.5 R_H (2) 4 R_H
(3) 8 R_H (4) R_H

11. De Broglie wavelength of an electron after being accelerated by a potential difference of V volt from rest is

(1) $\lambda = \frac{12.3}{\sqrt{h}}$ Å (2) $\lambda = \frac{12.3}{\sqrt{V}}$ Å
(3) $\lambda = \frac{12.3}{\sqrt{E}}$ Å (4) $\lambda = \frac{12.3}{\sqrt{m}}$ Å

12. After np orbitals are filled, the next orbital filled will be -

(1) (n + 1) s (2) (n + 2) p
(3) (n + 1) d (4) (n + 2) s

13. The correct set of four quantum numbers for the valence electron of Rubidium (Z = 37) is

(1) n = 5, $\ell = 0$, m = 0, s = $+\frac{1}{2}$
(2) n = 5, $\ell = 1$, m = 0, s = $+\frac{1}{2}$
(3) n = 5, $\ell = 1$, m = 1, s = $+\frac{1}{2}$
(4) n = 6, $\ell = 0$, m = 0, s = $+\frac{1}{2}$

14. No. of visible lines when an electron returns from 5th orbit to ground state in H spectrum -

(1) 5 (2) 4
(3) 3 (4) 10

15. If the shortest wave length of Lyman series of H atom is x, then the wave length of the first line of Balmer series of H atom will be -

(1) 9x/5 (2) 36x/5
(3) 5x/9 (4) 5x/36

16. Which of the above statement (s) is/are **false.**

I. Orbital angular momentum of the electron having n = 5 and having value of the azimuthal quantum number as lowest for this principle quantum number is $\frac{h}{\pi}$.

II. If n = 3, ℓ = 0, m = 0, for the last valence shell electron, then the possible atomic number may be 12 or 13.

III. Total spin of electrons for the atom $_{25}Mn$ is $\pm\frac{7}{2}$.

IV. Spin magnetic moment of inert gas is 0

(1) I, II and III (2) II and III only
(3) I and IV only (4) None of these

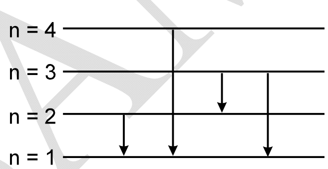
17. In case of $d_{x^2-y^2}$ orbital

(1) Probability of finding the electron along x-axis is zero.
(2) Probability of finding the electron along y-axis is zero.
(3) Probability of finding the electron is maximum along x and y-axis.
(4) Probability of finding the electron is zero in x-y plane

18. Electromagnetic radiations of wavelength 242 nm is just sufficient to ionise sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1}.

(1) 495 kJ/mol (2) 821 kJ/mol
(3) 136 kJ/mol (4) None

19. Suppose that a hypothetical atom gives a red, green, blue and violet line spectrum. Which jump according to figure would give off the red spectral line.



(1) $3 \rightarrow 1$ (2) $2 \rightarrow 1$
(3) $4 \rightarrow 1$ (4) $3 \rightarrow 2$

20. The difference between the wave number of 1st line of Balmer series and last line of paschen series for Li^{2+} ion is :

(1) $\frac{R}{36}$ (2) $\frac{5R}{36}$
(3) 4R (4) $\frac{R}{4}$

21. de-Broglie wavelength of electron in second orbit of Li^{2+} ion will be equal to de-Broglie of wavelength of electron in

(1) n = 3 of H-atom (2) n = 4 of C^{5+} ion
(3) n = 6 of Be^{3+} ion (4) n = 3 of He^{+} ion

22. Match List-I with List-II and select the correct answer using the codes given below the lists (ℓ and m are respectively the azimuthal and magnetic quantum no.)

List-I

(A) Number of value of ℓ for an energy level
(B) Value of ℓ for a particular type of orbital
(C) Number of values of m for ℓ = 2
(D) Value of 'm' for a particular type of orbital

List-II

(1) 0, 1, 2,......... (n - 1)
(2) $+\ell$ to $-\ell$ through zero
(3) 5
(4) n

Code :

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

23. A photon of 300 nm is absorbed by a gas and then emits two photons. One photon has a wavelength 496 nm then the wavelength of second photon in nm :

(1) 759 (2) 859
(3) 959 (4) 659

24. The uncertainty in momentum of an electron is 1 $\times 10^{-5}$ kg.m/s. The uncertainty in its position will be

($h = 6.62 \times 10^{-34}$ kg m^2/s) :

(1) 1.05×10^{-28} m (2) 1.05×10^{-26} m
(3) 5.27×10^{-30} m (4) 5.25×10^{-28} m

25. Which of the following configuration is correct for iron?

(1) $1s^2, 2s^2p^6, 3s^23p^63d^5$
(2) $1s^2, 2s^22p^6, 3s^23p^63p^6, 4s^2, 3d^5$
(3) $1s^2, 2s^22p^6, 3s^23p^6, 4s^2, 3d^7$
(4) $1s^2, 2s^22p^6, 3s^23p^63d^6, 4s^2$

26. Who modified Bohr's theory by introducing elliptical orbits for electron path?
(1) Hund (2) Thomson
(3) Rutherford (4) Sommerfeld

27. The de-Broglie wavelength of a particle with mass 1 g and velocity 100 m/s is:
(1) 6.63×10^{-33} m (2) 6.63×10^{-34} m
(3) 6.63×10^{-35} m (4) 6.65×10^{-36} m

28. The energy of photon is given as : Δe/atom = 3.03×10^{-19} J atom^{-1}, then the wavelength (λ) of the photon is:
(Given, h(Planck's constant)=6.63×10^{-34} J-s,
c (velocity of light) = 3.00×10^{8} m s^{-1})
(1) 6.56 nm (2) 65.6 nm
(3) 656 nm (4) 0.656 nm

29. The following quantum numbers are possible for how many orbital (s) n = 3, l = 2, m = + 2:
(1) 1 (2) 2
(3) 3 (4) 4

30. In hydrogen atom, energy of first excited state is – 3.4 eV. Then, KE of same orbit of hydrogen atom is:
(1) + 3.4 eV (2) + 6.8 eV
(3) – 13.6 eV (4) + 13.6 eV

31. The value of Planck's constant is 6.63×10^{-34} Js. The velocity of light is 3.0×10^{8} ms^{-1}. Which value is closest to the wavelength in nanometers of a quantum of light with frequency of 8×10^{15} s^{-1}?
(1) 2×10^{-25} (2) 5×10^{-18}
(3) 4×10^{1} (4) 3×10^{7}

32. The frequency of the radiation emitted when the electron falls from n = 4 to n = 1 in a hydrogen atom will be (Given ionization energy of H = 2.18×10^{-18} J atom^{-1} and h = 6.625×10^{-34} Js)
(1) 1.54×10^{15} s^{-1} (2) 1.03×10^{15} Js^{-1}
(3) 3.08×10^{15} s^{-1} (4) 2.0×10^{15} s^{-1}

33. Among the following transition metal ions, the one where all the metal ions have $3d^2$ electronic configuration is [At Nos. Ti = 22, V = 23, Cr = 24, Mn = 25]
(1) Ti^{3+}, V^{2+}, Cr^{3+}, Mn^{4+}
(2) Ti^{4+}, V^{4+}, Cr^{+6}, Mn^{7+}
(3) Ti^{4+}, V^{3+}, Cr^{2+}, Mn^{3+}
(4) Ti^{2+}, V^{3+}, Cr^{4+}, Mn^{5+}

34. The energy of the second Bohr orbit of the hydrogen atom is –328 kJ mol^{-1} ; hence the energy of fourth Bohr orbit would be
(1) – 1312 kJ mol^{-1} (2) – 82 kJ mol^{-1}
(3) – 41 kJ mol^{-1} (4) – 164 kJ mol^{-1}

35. **Given :** The mass of electron is 9.11×10^{-31} kg, planck constant is 6.626×10^{-34} J s, the Uncertainity involved in the measurement of velocity within a distance of 0.1 Å is :
(1) 5.79×10^{8} m s^{-1} (2) 5.79×10^{5} m s^{-1}
(3) 5.79×10^{6} m s^{-1} (4) 5.79×10^{7} m s^{-1}

36. The orientation of an atomic orbital is governed by :
(1) azimuthal eqantum number
(2) spin quantum number
(3) magnetic quantum number
(4) principal quantum number

37. Consider the following sets of quantum numbers :

	n	l	m	s
(i)	3	0	0	+1/2
(ii)	2	2	1	+1/2
(iii)	4	3	–2	–1/2
(iv)	1	0	–1	–1/2
(v)	3	2	3	+1/2

Which of the following sets of quantum numbers is not possible ?
(1) (i) and (iii) (2) (ii), (iii) and (iv)
(3) (i), (ii), (iii) and (iv) (4) (ii), (iv) and (v)

38. The measurement of the electron position is associated with an uncertainity in momentum which is equal to 1×10^{-18} g cm s^{-1}. The Uncertainity in electron velocity is
(mass of an electron is 9×10^{-28} g)
(1) 1×10^{9} cm s^{-1} (2) 1×10^{6} cm s^{-1}
(3) 1×10^{5} cm s^{-1} (4) 1×10^{11} cm s^{-1}

39. If the uncertainity in position and momentum are equal, then uncertainity in velocity is :
(1) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ (2) $\sqrt{\frac{h}{2\pi}}$
(3) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$ (4) $\sqrt{\frac{h}{\pi}}$

40. The isoelectronic pair is :
(1) Cl_2O, ICl_2^- (2) ICl_2^-, ClO_2
(3) IF_2^+, I_3^- (4) ClO_2^-, ClF_2^+

41. α – particles can be detected using :
(1) thin aluminium sheet
(2) barium suplhate
(3) zinc sulphide screen
(4) gold foil

42. The most probable radius (in pm) for finding the electron in He^+ is :
(1) 0.0 (2) 52.9
(3) 26.5 (4) 105.8

43. The de-broglie wavelength associated with a ball of mass 1 kg having kinetic energy 0.5 J is.
(1) 6.626×10^{-34} m (2) 13.20×10^{-34}m
(3) 10.38×10^{-21} m (4) 6.626×10^{-34} Å

44. The uncertainties in the velocities of two particles, A and B are 0.05 and 0.02 ms^{-1}, respectively. The mass of B is five times of that of the mass of A. What is the ratio of uncertainties $\left(\frac{\Delta x_A}{\Delta x_B}\right)$
(1) 2 (2) 0.25
(3) 4 (4) 1

45. What is the packet of energy then :
(1) Electron (2) Photon
(3) Positron (4) Proton

46. Heisenberg uncertainty principle can be explained as :
(1) $\Delta x \geq \frac{\Delta P \times h}{4\pi}$ (2) $\Delta x \times \Delta P \geq \frac{h}{4\pi}$
(3) $\Delta x + \Delta P \geq \frac{h}{\pi}$ (4) $\Delta P \geq \frac{\pi h}{\Delta x}$

47. Neutron is discovered by:
(1) Chadwick (2) Rutherford
(3) Yukawa (4) Dalton

48. If the wavelength of photon is 2.2×10^{-11} m, $h=6.6 \times 10^{-34}$ J-s, then momentum of photon is :
(1) 3×10^{-23} kg ms^{-1}
(2) 3.33×10^{22} kg ms^{-1}
(3) 1.4252×10^{-44} kg ms^{-1}
(4) 6.89×10^{43} kg ms^{-1}

49. Atoms with same atomic number and different mass numbers are called :
(1) isobars (2) isomers
(3) isotones (4) isotopes

50. The shape of the orbital with the value of $l = 2$ and m = 0 is :
(1) spherical (2) dumb-bell
(3) trigonal planar (4) square palnar

51. The correct one for d-orbitral is :
(1) $(n-1)\,d^{1-9}\,ns^1$
(2) $(n-1)\,d^{1-10}\,ns^{1-2}$
(3) $(n-1)\,d^{1-5}$
(4) $(n-1)\,d^{1-10}\,ns^2$

52. The value of amu is which of the following?
(1) 1.57×10^{-24} (2)1.66×10^{-124} kg
(3) 1.99×10^{-23} kg (4) 1.66×10^{-27} kg

53. Which one of the following has unit positive charge and 1 amu mass ?
(1) Electron (2) Neutron
(3) proton (4) None of these

54. The radius of hydrogen atom is 0.53 Å. The radius of $_3Li^{+2}$ is of :
(1) 1.27 Å (2) 0.17 Å
(3) 0.57 Å (4) 0.99 Å

55. The most probable velocity (In cm/s) of hydrogen molecule at 27°C will be
(1) 19.3×10^4 (2) 17.8×10^4
(3) 24.93×10^9 (4) $17.8\ 10^8$

EXERCISE # 3

Questions | Assertion and Reason

Directions : Each of these questions contains an Assertion followed by reason. Read them carefully and answer the question on the basis of following options. You have to select the one that best describes the two statements.

(1) If both assertion and reason are true and reason is the correct explanation of assertion.

(2) If both assertion and reason are true but reason is not the correct explanation of assertion.

(3) If Assertion is true but reason is false.

(4) If both assertion and reason are false.

1. **Assertion :** 2p orbital do not have any spehrical node.
Reason : The number of nodes in p-orbitals is given by (n–2) where n is the principal quantum number.

2. **Assertion :** The radii of corresponding orbitals in all H-like particles are equal.
Reason : All H-like particles contain more than one electron.

3. **Assertion :** The number of radial nodes in 3s and 4p orbitals is are equal.
Reason : The number of radial nodes in any orbital depends upon the values of 'n' and 'l' which are different for 3s and 4p orbitals.

4. **Assertion :** Electrons are ejected from a certain metal when either blue or violet light strikes the metal surface. However, only violet light causes ejection from second metal.
Reason : The electrons in the first metal require less energy for ejection.

5. **Assertion :** Hydrogen has one electron in its orbit but it produces several spectral lines.
Reason : There are many excited energy levels available.

6. **Assertion :** The energy of an electron is largely determined by its principal quantum number.
Reason : The principal quantum number (n) is a measure of the most probable distance of finding the electron around the nucleus.

7. **Assertion :** The 19th electron in potassium atom enters into 4s-orbital and not the 3d-orbital.
Reason : (n + 1) rule is followed for determining the orbital of the lowest energy state.

8. **Assertion :** The free gaseous Cr atom has six unpaired electrons.
Reason : Half-filled s-orbital has greater stability.

9. **Assertion :** The atoms of different elements having same mass number but different atomic number are known as isobars.
Reason : The sum of protons and neutrons in isobars is always different.

10. **Assertion :** K and Cs are used in photo–electric cells.
Reason : K and Cs emit electrons on exposure to light.

11. **Assertion :** A beam of electrons deflects more than a beam of α – particles in an electric field.
Reason : Electrons possess negative charge while α – particles possess positive charge.

12. **Assertion :** In Lyman of H–spectra, the maximum wavelength of lines is 121.65 nm.
Reason : Wavelength is maximum if there is transition from the very next level.

EXERCISE # 4

Questions | **Previous Year (NEET)**

1. Maximum number of electron in a subshell of an atom in determined by the following: **[AIPMT 2009]**

(1) $4l + 2$ (2) $2l + 1$
(3) $4l - 2$ (4) $2n^2$

2. The energy absorbed by each molecule (A_2) of a substance is 4.4×10^{-19} J and both energy per molecule is 4.0×10^{-19} J. The kinetic energy of the molecule per atom will be : **[AIPMT 2009]**

(1) 2.0×10^{-20} J (2) 2.2×10^{-19} J
(3) 2.0×10^{-19} J (4) 4.0×10^{-20} J

3. Which of the following is not permissible arrangement of electrons in an atom? **[AIPMT 2009]**

(1) n = 4, l = 0, m = 0, s = – 1/2
(2) n = 5, l = 3, m = 0, s = + 1/2
(3) n = 3, l = 3, m = 0, s = – 1/2
(4) n = 3, l = 2, m = – 2, s = – 1/2

4. The de–Broglie wavelength of helium atom at room temperature is : **[AIIMS 2009]**

(1) 6.6×10^{-34} m (2) 4.39×10^{-10} m
(3) 7.34×10^{-11} m (4) 2.335×10^{-20} m

5. n and ℓ for some electrons are given. Which of the following is expected to have least energy ?

(1) n = 3, ℓ = 2 (2) n = 3, ℓ = 0
(3) n = 2, ℓ = 1 (4) n = 4, ℓ = 0

6. Gamma rays are : **[AFMC 2009]**

(1) high energy electrons
(2) low energy electrons
(3) high energy electro-magnetic waves
(4) high energy positrons

7. A 0.66 kg ball is moving with a speed of 100 m/s. The associated wavelength will be:

($h = 6.6 \times 10^{-34}$ Js) **[AIPMT 2010]**

(1) 6.6×10^{-32} m (2) 6.6×10^{-34} m
(3) 1.0×10^{-35} m (4) 1.0×10^{-32} m

8. The total number of atomic orbitals in fourth energy level of an atom is : **[AIPMT 2011]**

(1) 8 (2) 16 (3) 32 (4) 4

9. The energies E_1 and E_2 of two radiations are 25 eV and 50 eV respectively. The relation between their wavelengths i.e. λ_1 and λ_2 will be : **[AIPMT 2011]**

(1) $\lambda_1 = \lambda_2$ (2) $\lambda_1 = 2\lambda_2$
(3) $\lambda_1 = 4\lambda_2$ (4) $\lambda_1 = \frac{1}{2}\lambda_2$

10. If n = 6, the correct sequence for filling of electrons will be : **[AIPMT 2011]**

(1) ns → (n – 2)f → (n – 1)d → np
(2) ns → (n – 1)d → (n – 2)f → np
(3) ns → (n – 2)f → np → (n – 1)d
(4) ns → np(n – 1)d → (n – 2)f

11. According to the Bohr Theory, which of the following transitions in the hydrogen atom will give rise to the least energetic photon ? **[AIPMT 2011]**

(1) n = 6 to n = 1 (2) n = 5 to n = 4
(3) n = 6 to n = 5 (4) n = 5 to n = 3

12. Maximum number of electrons in a subshell with : I = 3 and n = 4 is : **[AIPMT 2012]**

(1) 14 (2) 16
(3) 10 (4) 12

13. The correct set of four quantum numbers for the valence elecron of rubidium atom (Z=37) is : **[AIPMT 2012]**

(1) 5, 1, + 1/2 (2) 6, 0, 0 + 1/2
(3) 5, 0, 0 + 1/2 (4) 5, 1, 0 + 1/2

14. The orbital angular momentum of a p-electron is given as : **[AIPMT 2012]**

(1) $\frac{h}{\sqrt{2}\pi}$ (2) $\sqrt{3}\frac{h}{2\pi}$
(3) $\sqrt{\frac{3}{2}}\frac{h}{\pi}$ (4) $\sqrt{6}.\frac{h}{2\pi}$

15. The value of Planck's constant is 6.63×10^{-34} Js$\left(\frac{Z^2}{n^2}\right)$. The speed of light is 3×10^{17} nm s^{-1}. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 6×10^{15} s^{-1}? **[NEET 2013]**

(1) 25 (2) 50 (3) 75 (4) 10

16. Based on equation $E = -2.178 \times 10^{-18} J\left(\frac{Z^2}{n^2}\right)$, certain conclusions are written. Which of them is not correct? **[NEET 2013]**

(1) Larger the value of n, the larger is the orbit radius.

(2) Equation can be used to calculate the change in energy when the electron changes orbit.

(3) For n = 1, the electron has a more negative energy than it does for n = 6 which mean that the electron is more loosely bound in the smallest allowed orbit.

(4) The negative sign in equation simply means that the energy or electron bound to the nucleus is lower than it would be if the electrons were at the infinite distance from the nucleus.

17. What is the maximum numbers of electrons that can be associated with the following set of quantum numbers ? $n = 3$, $l = 1$ and $m = -1$. **[NEET 2013]**

(1) 6 (2) 4
(3) 2 (4) 10

18. Be^{2+} is isoelectronic with which of the following ions? **[NEET 2014]**

(1) H^+ (2) Li^+
(3) Na^+ (4) Mg^{2+}

19. Calculate the energy in joule corresponding to light of wavelength 45 nm. (Plank's constant, $h = 6.63\times10^{-34}$ J s, speed of light, $c = 3 \times 10^8$ ms^{-1}) **[NEET 2014]**

(1) 6.67×10^{15} (2) 6.67×10^{11}
(3) 4.42×10^{-15} (4) 6.67×10^{-18}

20. What is the maximum numbers of orbitals that can be identified with the following set of quantum numbers ? $n = 3$, $l = 1$ and $m = 0$. **[NEET 2014]**

(1) 1 (2) 2
(3) 3 (4) 4

21. The angular momentum of electron in 'd' orbital is equal to **[NEET 2015]**

(1) $2\sqrt{3}$ h (2) 0 h
(3) $\sqrt{6}$ h (4) $\sqrt{2}$ h

22. The number of d-electron in Fe^{2+} (Z = 26) is not equal to the number of electrons in which one of the following? **[NEET 2015]**

(1) d-electrons in Fe (Z = 26)
(2) p-electrons in Ne (Z = 10)
(3) s-electrons in Mg (Z = 12)
(4) p-electrons in Cl (Z = 17)

23. Which is the correct order of increasing energy of the listed orbitals in the atom of titanium? (At. no. Z = 22) **[NEET 2015]**

(1) 4s 3s 3p 3d (2) 3s 3p 3d 4s
(3) 3s 3p 4s 3d (4) 3s 4s 3p 3d

24. Two electrons occupying the same orbital are distinguished by **[NEET 2016]**

(1) azimuthal quantum number
(2) spin quantum number
(3) principal quantum number
(4) magnetic quantum number

25. Which of the following pairs of d-orbitals will have electron density along the axes? **[NEET 2016]**

(1) d_{z^2}, d_{xz} (2) d_{xz}, d_{yz}
(3) $d_{z^2}, d_{x^2-y^2}$ (4) $d_{xy}, d_{x^2-y^2}$

26. How many electrons can fit in the orbital for which $n = 3$ and $l = 1$? **[NEET 2016]**

(1) 2 (2) 6
(3) 10 (4) 14

27. Which one is the wrong statement? **[NEET 2017]**

(1) The uncertainty principal is $\Delta E \times \Delta t \geq \frac{h}{4\pi}$

(2) Half-filled and fully filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.

(3) The energy of 2s-orbital is less than the energy of 2p-orbital in case of hydrogen like atoms.

(4) de-Broglie's wavelength is given by $\lambda = \frac{h}{mv}$ where m = mass of the particles, v = group velocity of the particles.

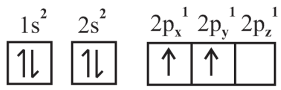
28. Which one is a wrong statement?

[NEET 2018]

(1) Total orbital angular momentum of electron in s-orbital is equal to zero.

(2) An orbital is designed by three quantum numbers while an electron in an atom is designed by four quantum numbers.

(3) The electronic configuration of N atom is



(4) The value of m for d_{z^2} is zero.

29. In hydrogen atom, the de Broglie wavelength of an electron in the second Bohr orbit is [Given that Bohr radius, a_0 = 5.29 pm] **[NEET 2019]**

(1) 211.6 pm (2) 211.6 π pm

(3) 52.9 π pm (4) 105.8 pm

30. Orbital having 3 angular nodes and 3 total nodes is

(1) 5*p* (2) 3*d*

(3) 4*f* (4) 6*d*

31. Which of the following series of transitions in the spectrum of hydrogen atom falls in visible region?

[NEET 2019]

(1) Brackett series (2) Lyman series

(3) Balmer series (4) Paschen series

32. 4*d*, 5*p*,5*f* and 6*p* orbitals are arranged in the order of decreasing energy. The correct option is

[NEET 2019]

(1) 5*f* > 6*p* >4*d* > 5*p*

(2) 5*f* > 6*p* >5*p* > 4*d*

(3) 6*p* > 5*p* > 5*p* > 4*d*

(4) 6*p* > 5*f* > 4*d* > 5*p*

ANSWER KEY

EXERCISE-1

Qus.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	3	2	4	2	1	4	2	1	3	1	1	2	4	1	2	1	3	1	2	1
Qus.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	2	4	1	2	1	1	3	2	3	2	3	2	2	1	3	3	3	2	3	2
Qus.	**41**	**42**	**43**	**44**	**45**	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**	**56**	**57**	**58**	**59**	**60**
Ans.	1	3	3	4	3	1	1	1	2	3	4	2	4	3	1	4	2	3	1	3
Qus.	**61**	**62**	**63**	**64**	**65**	**66**	**67**	**68**	**69**	**70**	**71**	**72**	**73**	**74**	**75**	**76**	**77**	**78**	**79**	**80**
Ans.	3	3	4	1	2	1	3	1	1	2	1	1	3	3	4	4	4	2	4	1
Qus.	**81**	**82**	**83**	**84**	**85**															
Ans.	3	3	4	3	1															

EXERCISE-2

Qus.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	4	3	3	2	4	1	3	4	4	3	2	1	1	3	2	1	3	1	4	4
Qus.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	2	2	1	3	4	4	1	3	1	1	3	3	4	2	3	3	4	1	1	4
Qus.	**41**	**42**	**43**	**44**	**45**	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**					
Ans.	3	3	1	1	2	2	1	3	4	2	2	4	3	2	2					

EXERCISE-3

Qus.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**
Ans.	1	4	1	1	1	1	1	3	3	3	2	1

EXERCISE-4

Qus.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	1	1	3	3	3	3	3	2	2	1	3	1	3	1	3	3	3	2	4	1
Qus.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**								
Ans.	3	4	3	2	3	1	3	3	2	3	3	2								