STRUCTURE OF ATOM

OBJECTIVES:

After studying this unit, student will be able to

1) Know about the discovery of electron, proton and neutron and their characteristics;
2) Describe Thomson, Rutherford and Bohr atomic models;
3) Understand the important features of the quantum mechanical model of atom;
4) Understand nature of electromagnetic radiation and Planck's quantum theory;
5) Explain the photoelectric effect and describe features of atomic spectra;
6) State the de Broglie relation and Heisenberg uncertainty principle;
7) Define an atomic orbital in terms of quantum numbers;
8) State Aufbau principle, Pauli exclusion principle and Hund's rule of maximum multiplicity; and
9) Write the electronic configurations of atoms.

❖ Sub - atomic particles or fundamental particles :

Up to 19th century atom was considered as the ultimate particles of matter. This first theory about the structure of matter was given by John Dalton and the theory was called as Dalton's theory.

According to Dalton's atomic theory atom was indivisible, but by a species of experiment evidences, it is confirmed that atom is divisible and consists of sub-atomic particles such as electron, proton and neutron. These sub-atomic particles are regarded as fundamental particles.

❖ Fundamental particles of an atom :

(1) Electron

(2) Proton

(3) Neutron

❖ Discovery of Electron :

J.J. Thomson discovered electron by discharge tube experiment.

The charge on the electron was discovered by Millikan by oil drop method.

Charge on electron: 1.602×10^{-19} C.

∴ Electron is considered as unit negative charged particle.

The charged to mass ratio (*e*/m ratio) of electron = $1.76 \times 10^{11}\ Ckg^{-1}$

The *e*/m ratio for electron is same for all irrespective of **the gas used in the Discharge Tube**

Mass of an electron = $\frac{e}{e/m} = \frac{charge\ on\ electron}{charge\ to\ mass\ of\ electron}$

$$= \frac{1.602\ x\ 10^{-19}\ C}{1.758820\ x\ 10^{11}\ Ckg^{-1}}$$

∴ Mass of an electron = 9.1094×10^{-31} kg

∴ Mass of electron is considered as Zero Mass.

Symbol of electron = ${}_{-1}^{0}e$

❖ Discovery of Proton :

Proton was discovered by Goldstein by discharge tube experiment using perforated Cathode.

The charge to mass ratio ($e/$m) of proton = **9.58 x $10^7 Ckg^{-1}$**

The $e/$m ratio of proton depends upon the nature of the gas used in discharge tube. It is maximum for hydrogen.

Charge on proton = 1.602 x 10^{-19} C.

Proton is considered as unit positive charge particle.

Mass of proton, $\frac{e}{e/m} = \frac{1.602 \ x \ 10^{-19} \ C}{9.58 \ x \ 10^7 \ Ckg^{-1}}$

∴ Mass of proton = 1.67 x 10^{-27} kg (one unit mass)

Mass of proton is 1837 times the mass of electron

i.e. Electron has mass equal to $\frac{1}{1837}$ of mass of hydrogen atom.

Symbol of proton = ${}^1_1H$

❖ Discovery of Neutron :

Neutron was discovered by James Chadwick in 1932, after Rutherford's atomic model.

Mass of neutron = 1.673 x 10^{-27} kg

(equal to the mass of hydrogen i.e. unit mass)

Charge on Neutron = No charge.

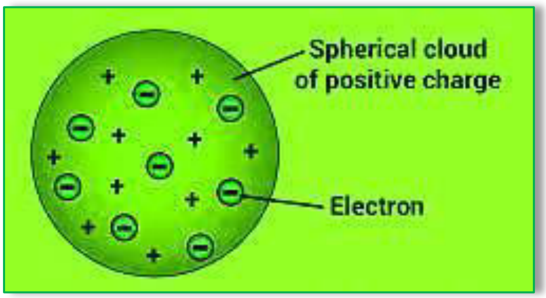
Symbol of Neutron = ${}^1_0n$

After the discovery of electron and proton, scientists started thinking of arranging these particles in an atom. The first model was prepared by J.J Thomson, known as Thomson's atomic model.

❖ Thomson's atomic model :

According to Thomson's atomic model, atom is sphere of about 10^{-8} cm (10^{-10} m) radius in which positive charge is uniformly distributed and negatively charge electrons are embeded into it to give stable electrostatic arrangements.

This model is also known as Plum Pudding model or water melon model.



❖ **Drawbacks of Thomson's atomic model** :

1) Mass of the atom is considered to be evenly spread over the atom.

2) It does not reflect the movement of electron.

❖ **Rutherford's atomic model**:

By performing $\propto$ - ray scattering experiment (i.e. Bombardment of $\propto$ - particle on gold foil), Rutherford proposed a model, called as Rutherford's atomic model also known as planetary model of an atom.

According to Rutherford,

1)The entire mass of an atom is concentrated in a region of the centre, known as Nucleus.

2) The volume occupied by nucleus is very small as compared to the total volume of the atom.

3)Radius of nucleus is of the order of 10^{-15} m

4) Radius of an atom is of the order of 10^{-10} m

5) The positive charge of the nucleus is due to protons.

6) The mass of the nucleus is due to protons and other neutral particles, having mass nearly equal to mass of the proton.

7) This neutral particle is Neutron, discovered by Chadwick in 1932.

8) Protons and Neutrons present in the nucleus are collectively called Nucleons.

9)The nucleus is surrounded by negatively charged electrons which balance the positive charge on the nucleus .So the atom is electrically neutral.

10) The electron are not stationary but revolves around the nucleus at very high speed. The electrons and the nucleus are held together by electrostatic forces of attraction.

11) The radius of the nucleus is proportional to the cube root of the no. of nucleons within it.

$\therefore$ R $\propto A^{1/3}$

R $= R_0 A^{1/3}$ cm

Where,

R_0= 1.33 x 10^{-13}(a constant)

A = mass no. (P + n)

R = radius of the nucleus

❖ **<u>Drawbacks of the Rutherford's atomic model</u> :**

1) It fails to explain stability of atom.

2) It fails to explain position, energy and distribution of electron around the nucleus.

3) It fails to explain atomic spectrum.

Atomic Number (Z) and Atomic Mass No. (A)

Atomic Number (Z) = Total no. of protons present in the nucleus .

= Total no. of electrons present in an atom.

Atomic mass No.(A) = No. of Protons (Z) + No. of Neutrons

A = Z + N

A = Total number of Nucleons

$^{A}_{Z}X$ X = symbol of an element

A = Mass Number

Z = Atomic Number

$\because$ A = Z + N

$\therefore$ N = A – Z

Q.1 How many protons, neutrons and electrons are present in $^{17}_{8}O$.

$\Longrightarrow$ At.No.(Z) = 8 , Mass No. (A) = 17

No. of protons = No of electrons = Z = 8

No. of neutrons (N) = A – Z = 17 – 8

= 9

IMPORTANT TERMS

<u>1) Isotopes :</u>

Atoms of same element having same atomic number but different mass number.

Examples :

(1) Isotopes of Hydrogen : $^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$

(2) Isotopes of Carbon : $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$

2) Isobars :

Atoms having same mass number but different atomic number are called Isobars.

Examples :

(1) ${}^{40}_{18}Ar$, ${}^{40}_{19}K$, ${}^{40}_{20}Ca$, ${}^{14}_{6}C$, ${}^{14}_{7}N$

3) Isotones :

Atoms of different elements having same no. of neutrons in their nuclei are called as Isotones.

Examples :

(1) ${}^{14}_{6}C$, ${}^{15}_{7}N$, ${}^{16}_{8}O$ (8 neutrons each)

4) Isoelectronic species :

Species (atoms or ions) containing same number of electrons are called Isoelectronic.

O_2^-, F^-, Na^+, Mg^{2+}, Al^{3+}, Ne are isoelectronic because each of these contain 10 electrons.

5) Electromagnetic radiation :

Radiation is a form of energy which can be transferred from one point to another in space .Radiation, which is associated with electric and magnetic field, are called as electronic radiation.

Examples :

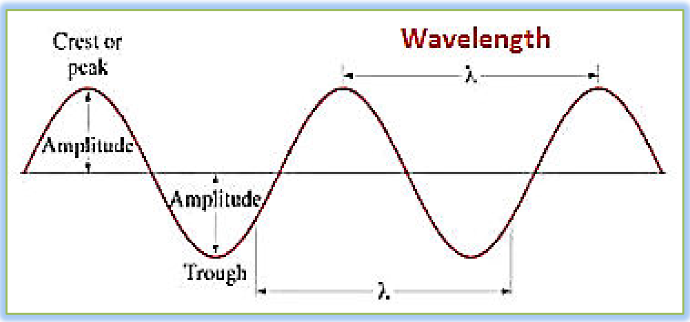
Cosmic rays, γ- rays, x-rays, ultra – violet rays, Infra red, micro waves, radio waves.

- ❖ **Dual Nature of electromagnetic radiation :**

(1) Wave nature of radiation.

(2) Particle nature of radiation.

- ❖ **Characteristics of waves :**



1) **Wavelength**:

It is denoted by lambda (λ). It is the distance between two adjacent crest or trough.

Unit:Angstrom (A^{0})

S.I. Unit : metre (m)s

$1A^{0} = 10^{-8}$ cm $= 10^{-10}$ m

1 nanometre (1nm) $= 10^{-9}$ m

1 picometre(1pm) $= 10^{-10}$cm $=10^{-12}$ m

1 micrometre (1 μm) $= 10^{-6}$m

2) **Wave number ($\bar{\nu}$)** :

The number of wave travelled per unit length is called Wave number.

Wave number is reciprocal of wavelength

$\therefore \quad \bar{\nu} = {}^{1}/_{\lambda}$

Unit of Wave number $= cm^{-1}$ or m^{-1}
S.I. Unit : m^{-1}

3) **Frequency (ν)** :

Number of waves passing through a given point in one second is called Frequency.

S.I. Unit : Hertz (Hz) or s^{-1}

4) **Amplitude (A)** :

Height of Crest or depth of Trough of a wave is called Amplitude.

5) **Velocity (C)** :

Distance travelled by a wave in one second.

C = 3 x 10^{8} m/s

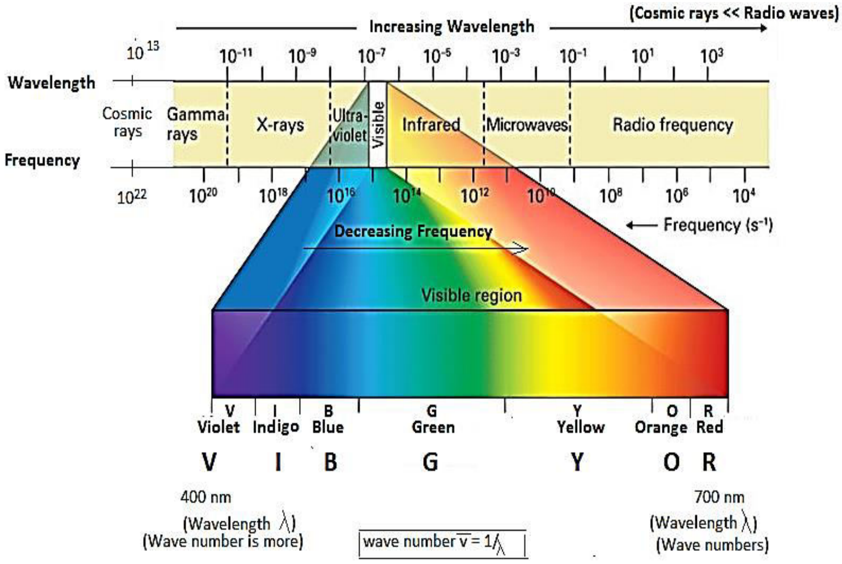
❖ **Relationship between wavelength(λ), wave number($\bar{\nu}$), Frequency (ν) and velocity (C)**

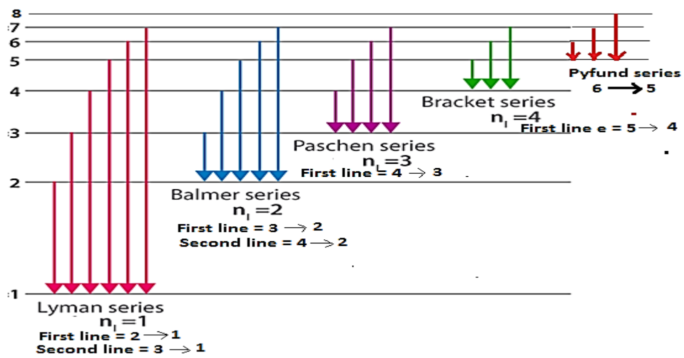
$\because \quad C = \nu\lambda$

$\therefore \quad \nu = \frac{C}{\lambda}$ and $\frac{1}{\lambda} = \bar{\nu}$

$\therefore \quad \nu = C\bar{\nu}$

❖ Electromagnetic spectrum :





Spectra lines of Hydrogen atom

The arrangement of different types of electromagnetic radiation in the order of increasing wavelength (or decreasing frequency) is called Electromagnetic spectrum.

Out of the various colours in the visible range, violet colour has minimum wavelength (400 nm) or maximum frequency (7.5 x 10^{14}) Hz .

Red colour has maximum wavelength (700 nm) or minimum frequency (4.0 x 10^{14}) Hz.

❖ Numericals on Electromagnetic radiation :

Q1) Visible spectrum has wavelength extends from 400 nm (violet) to 750 nm (red). Express these wavelength in terms of frequencies Hz.

$\Longrightarrow$ Given : Wavelength , λ_1 = 400 nm = 400 x 10^{-9} (for violet rays)

λ_2 = 750 nm = 750 x 10^{-9} (for red light)

($\because$ 1nm = 10^{-9})

To find : Frequencies, ν_1 = ? ν_2 = ?

Formula : $\nu = \frac{C}{\lambda}$

Calculation :

$$\nu_1 = \frac{3.0\ x\ 10^8\ ms^{-1}}{400\ x\ 10^{-9}\ m} = 7.50 \text{ x } 10^{14} \text{ Hz}$$

$$\nu_2 = \frac{3.0\ x\ 10^8\ ms^{-1}}{750\ x\ 10^{-9}\ m} = 4.00 \text{ x } 10^{14} \text{ Hz}$$

Q2) Calculate the frequency and wave number of yellow light emitted from a lamp having wavelength of 580 nm.

⟹ Given : Wavelength (λ) = 580 nm = 580 x 10^{-9} m

To find : Frequency (ν) = ?   Wave number($\bar{\nu}$) = ?

Formulae : $\nu = \frac{C}{\lambda}$

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

Frequency (ν) = $\frac{C}{\lambda} = \frac{3.0\ x\ 10^8\ ms^{-1}}{580\ x\ 10^{-9}\ m}$ = 5.17 x $10^{14} s^{-1}$

Wave number ($\bar{\nu}$) = $\frac{1}{\lambda} = \frac{1}{580\ x\ 10^{-9}\ m}$ = 1.72 x $10^6 m^{-1}$

❖ **Particle Nature of Electromagnetic radiation**

(Plank's Quantum Theory of Radiation)

In the year 1900, Max Plank proposed his quantum theory to explain black body radiation.

❖ **Black body radiation :**

An ideal body which emits and absorb radiations of all wavelengths or frequency is called Black Body and the radiations emitted by black body is called Black Body Radiation.

According to Plank's Quantum Theory, the energy of electromagnetic radiation depends upon the frequency. Plank gave the name (Quantum) to the smallest quantity of energy that can be emitted or absorbed. The energy (E) of each quantum of radiation is directly proportional to its frequency (ν).

∴ E $\propto \nu$

∴ E = hν(h = Plank's Constant = 6.625 x 10^{-34})

∴ $E = \frac{hC}{\lambda}$ $(\nu = \frac{1}{\lambda})$

Shorter the wavelength, larger the frequency ν and higher the energy. So, Blue light, which has shorter wavelength (400 nm) than red light (750 nm) has higher energy.

❖ **Photoelectric effect :**

This effect was observed by J.J Thomson. When a light of suitable frequency called Threshold frequency, strikes the surface of a metal, electrons are ejected from the metal. This phenomenon is known as Photoelectric effect and the ejected electrons are called as Photoelectron.

The number of ejected electrons (Photoelectron) depends upon the intensity or brightness of incident light but does not depend upon its frequency.

The kinetic energy of the photo – electrons is directly proportional to the frequency of the intensity. Therefore, Blue light eject electrons faster than red light.

For the ejection of photo-electrons, the frequency of light (ν) must be greater than threshold frequency (ν_0). If frequency of light (ν) is less than threshold frequency (ν_0), there is no ejection of electron. Threshold frequency is not the same for all metals.

❖ **Einstein's Photoelectric Equation :**

$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

where,

$h\nu_0$ is the threshold energy, the energy required to overcome the attractive forces on the electron in the metal.

$h\nu_0$ is also called work function (W_0).

❖ **Numerical on photoelectric effect :**

Q1) Calculate the work function , if in a photo-electric effect , the energy of the photo striking a metallic surface is 5.6 x 10^{-19} J and the kinetic energy of the ejected electrons is 12.0 x 10^{-20} J .

a) 6.4 x 10^{-19} J | b) 6.8 x 10^{-19} J

c) 4.4 x 10^{-19} J | d) 6.4 x 10^{-20} J

⟹ Formula : $h\nu = h\nu_0 + \frac{1}{2}mv^2$

∴ work function = $h\nu_0 = h\nu$ – K.E.

= 5.6 x 10^{-19} – 1.2 x 10^{-19}

∴ work function = 4.4 x 10^{-19} J

Q2) The threshold wavelength for the ejection of electron from metal 'X' is 330 nm. Calculate the work function for photoelectric emission from metal X.

a) 1.2×10^{-18} J　　b) 1.2×10^{-20} J

c) 6×10^{-19} J　　d) 6×10^{-12} J

$\Rightarrow$ Work function, $h\nu_0 = \frac{hC}{\lambda_0}$

$$= \frac{6.6 \times 10^{-34} \times 3 \times 10^{8}}{330 \times 10^{-9}}$$

$= 6 \times 10^{-19}$ J

❖ **Atomic Spectra :**

Atoms of different elements emit electromagnetic radiations of definite frequencies, when excited by heating,passing current or electric discharge. Arrangement of these radiations in decreasing order of frequencies is called atomic spectrum.

❖ **Continuous Spectrum :**

It is obtained by passing sunlight (white light) through a prism. The light is dispersed into continuous spectra of colours from violet to red (like rainbow). It contains radiations of all the frequencies.

❖ **Discontinuous Spectra or line Spectrum :**

It is an atomic spectrum of an element which consists of number of bright lines, separated by dark line.

❖ **Absorption Spectrum :**

When continuous electromagnetic radiation (like white light) is allowed to passed through a gas or a solution and transmitted light is analysed, a spectrum is obtained in which dark lines are obtained over continuous spectrum, called as Absorption Spectrum.

❖ Emission Spectrum :

It is obtained by passing radiations from the atoms through prism.

It has few bright lines against a dark background.

❖ Hydrogen Spectrum :

When electric discharge is passed through gaseous hydrogen, it emits radiation. The recorded spectrum of this emitted radiation is called hydrogen emission spectrum. The entire spectrum consists of five series of lines. The wavelength of all these series can be expressed by a single formula given by Rydberg.

Rydberg formula :

$$\bar{\nu} = \frac{1}{\lambda} = Z^2 R_H \left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right) cm^{-1} \qquad \begin{Bmatrix} For\ He^+ , Z = 2 \\ For\ Li^{2+} , Z = 3 \end{Bmatrix}$$

Where,

$\bar{\nu}$ = wave number

λ = wavelength

R_H = Rydberg Constant for Hydrogen = 109678 cm^{-1}

n_1, n_2 = Electronic levels involved in transition

Spectral lines	n_1	n_2	Region
Lyman	1	2 , 3 , 4 etc	Ultra - violet
Balmer	2	3 , 4 , 5 etc	Visible
Paschen	3	4 , 5 , 6 etc	Infra - red
Bracket	4	5 , 6 , 7 etc	Infra - red
Pfund	5	6 , 7 , 8 etc	Infra - red

❖ Numerical on Rydberg equation :

Q1 What is the wavelength of the radiation emitted during the transition from the orbit of n = 2 to that of n = 1 in the hydrogen atom?

$\Longrightarrow$ $\bar{\nu} = 109677 \left[\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right] cm^{-1}$

Here, $n_1 = 1,\ n_2 = 2$

$\therefore \quad \bar{\nu} = 109677 \left[\frac{1}{1^2} - \frac{1}{2^2}\right]$

$= 109677 \left[\frac{1}{1} - \frac{1}{4}\right]$

$\bar{\nu} = 109677 \text{ x } \frac{3}{4}$

$\bar{\nu} = 82257.75\ cm^{-1}$

Now, wavelength, $\lambda = \frac{1}{\nu}$

$\lambda = \frac{1}{82257.75}$

$\lambda = 1.216 \text{ x } 10^{-5}$ cm

$= 1.216 \text{ x } 10^{-5} \text{ x } 10^{7}$ nm

$\lambda = 121.6$ nm

Q2 In the Rydberg equation, a spectral line corresponds to n_1 = 3 and n_2 = 5, Calculate the wavelength and frequency of this spectral lines. To which spectral lines, this line belongs?

$\Longrightarrow$ According to Rydberg equation,

$\frac{1}{\lambda} = R_H \left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right) cm^{-1}$

$\frac{1}{\lambda} = 109677 \left(\frac{1}{3^2} - \frac{1}{5^2}\right)$

$\frac{1}{\lambda} = 109677 \text{ x } \frac{16}{225}\ cm^{-1}$

$\therefore \quad \lambda = \frac{225}{109677\ x\ 16}$

$\therefore \quad \lambda = 12.82 \text{ x } 10^{-5}$ cm $= 1282 \text{ x } 10^{-9}$ m $=$ 1282 nm

Now,

$\text{v} = \frac{c}{\lambda} = \frac{3.0\ x\ 10^{8}\ ms^{-1}}{1282\ x\ 10^{-9}\ m}$

$\text{v} = 2.34 \text{ x } 10^{14} s^{-1}$

Since, this line corresponds to $n_1 = 3$, it belongs to Paschen series.

Q3 Calculate the longest wavelength transition in the Paschen series of He^+

$\Longrightarrow$ $$\bar{\nu} = Z^2 R_H \left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right) cm^{-1}$$

For He, Z = 2, for Paschen series, $n_1 = 3$

For longest wavelength, $n_2 = 4$

$$\therefore \quad \bar{\nu} = \frac{1}{\lambda} = 109678 \text{ x } (2)^2 \text{ x } \left[\frac{1}{3^2} - \frac{1}{4^2}\right]$$

$$= 109678 \text{ x } 4 \text{ x } \frac{7}{144}$$

$$\therefore \quad \lambda = 4689 A^0$$

❖ ***Calculation of total number of emission lines when an electron jumps from n_2 level to n_1 level are given by the relation,***

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

Ex: Calculate total number of spectral lines in Balmer series from 5^{th} energy level.

$\Longrightarrow$ No. of spectral lines = $\frac{(5-2)[(5-2)+1]}{2}$

$$= \frac{3 \times 4}{2} = 6$$

$$\left\{\begin{array}{r} 5 \to 4 , 5 \to 3 , 5 \to 2 = 3 \; lines \\ 4 \to 3 , 4 \to 2 \; = 2 \; lines \\ 3 \to 2 = 1 \; lines \\ = 6 \; lines \end{array}\right\}$$

$\therefore$ No. of spectral lines = 6

❖ **<u>Bohr's Atomic Model :</u>**

Niel Bohr developed a model for hydrogen atom and hydrogen like one electron species like He^+ , Li^{2+} , Be^{3+} , etc .

Important Postulates of BOHR'S MODEL of an atom:

1) Electron revolves around the nucleus in fixed circular orbits of definite energy.

2) As long as an electron revolve in such orbits, it does not radiate any energy.

3) Electrons revolves only in thoseorbits whose angular momentum (mvr) is an integral multiple of $\frac{h}{2\pi}$ i.e. mvr = n $\frac{h}{2\pi}$

4) Ordinarily, an electron continues to move in a lowest energy state, called ground state and when electron absorb energy and jumps to higher state, it is called excited state.

5) The emission or absorption of energy in the form of radiation can only occur, when electron jumps from one stationary state to another.

6) Energy is absorbed when electron jumps from inner to outer orbit and energy is emitted when electron moves from outer to inner orbit.

7) The energy change (ΔE) in an electron jump is given by, $\Delta E = E_2 - E_1 = h\nu$

Energy of electron in n^{th} orbit for H atom,

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

Where,

z=atomic number of atom

m = mass of electron = 9.10×10^{-31} Kg

e =charge on the electron = 1.602×10^{-19}C

h =Plank's Constant = 6.625×10^{-34} $kgm^2 s^{-1}$

Important Formulae for energy of electron in the n^{th} orbit

1) $E_n = -\frac{13.6\, Z^2}{n^2}$ eV

2) $E_n = -\frac{13.6}{n^2}$ J $atom^{-1}$

$$\left[\begin{array}{r}\text{Ground state energy of H atom} = -13.6\text{ eV}\\ \text{Ground state energy of He}^{+}\text{ion }(Z=2)\\ = -13.6 \times 4 = -54.4\text{ eV}\end{array}\right]$$

3) $E_n = -2.18 \times 10^{-18}\left(\frac{Z^2}{n^2}\right)$ J

4) $E_n = -\frac{2.18 \times 10^{-11} Z^2}{n^2}$ ergs

Q) For H like atom, $E_n = Z^2 \times 1311.8$ kJ/mol

Ex: For He^+, Z = 2

$\therefore$ $E_n = (2)^2 \times 1311.8 = 5247.2$ kJ/mol

For Li^{2+}, Z = 3

$\therefore$ $E_n = (3)^2 \times 1311.8 = 11806.2$ kJ/mol

❖ Numerical on Bohr's orbit :

1) Energy of third orbit of Bohr's atom is

a) -13.6 eV   b) -3.4 eV   c) -1.5 eV   d) -6.8 eV

⟹ $E_n = -\frac{13.6\, Z^2}{n^2}$ eV

For Hydrogen, Z = 1 and for 3^{rd} orbit, n = 3

∴ $E_3 = -\frac{13.6}{3^2} = -1.5$ eV $= -2.42 \times 10^{-19}$ J (1eV = 1.602×10^{-19} J)

2) Energy of first orbit of Bohr's atom is

a) -13.6 J   b) 52.9 J   c) 2.18×10^{-18} J   d) -2.18×10^{-18} J

⟹ Energy of first orbit,

$$E_1 = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2}\right) \text{J}$$

$$= -2.18 \times 10^{-18} \left(\frac{1^2}{1^2}\right) \text{J}$$

$$= -2.18 \times 10^{-18} \text{ J}$$

3) **For second orbit, n = 2**

⟹ $$E_2 = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2}\right) \text{J}$$

$$= -2.18 \times 10^{-18} \left(\frac{1^2}{2^2}\right) \text{J}$$

$$= -0.545 \times 10^{-19} \text{ J}$$

The radius of the first orbit of hydrogen atom is called <u>Bohr's radius</u>. (Bohr's orbit)

Radius of Bohr's Orbit (n^{th} orbit) :

$$r = \frac{n^2h^2}{4\pi^2 mZe^2} \qquad \{Greater\ the\ value\ of\ 'n',\ Greater\ is\ the\ radius\}$$

$$r_n = 0.529 \times \frac{n^2}{Z} A^0 = 52.9 \times n^2 \text{ pm}$$

$$r_n = 5.29 \times 10^{-11} \text{ m} \quad (\text{for } n = 1, \text{ for H-atom})$$

For Hydrogen like species, (He^+ , Li^{2+} , Be^{3+})

$$r_n = \frac{52.9(n^2)}{Z} \text{ pm}$$

❖ Numerical / MCQ's on Bohr's radius / energy.

Q1) If 'r' is the radius of first orbit the radius of n^{th} orbit of H-atom is given by

a) $r\, n^2$ b) $r\, n$ c) r/n d) $r^2\, n^2$

⟹ Radius of n^{th} orbit of H-atom = $r_0 n^2$

Where,

r_0 = radius of first orbit

Q2) According to Bohr theory, which of the following transitions in the hydrogen atom will give rise to the least energetic photon?

a) n= 6 to n = 1 b) n = 5 to n = 4 c) n = 6 to n = 5 d) n = 5 to n = 3

$\Rightarrow$ $\Delta E = \left[\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right]$ where, $n_2 > n_1$

$\therefore$ n = 6 to n = 5 will give least energetic photon.

Q3) The energy of second Bohr orbit of the hydrogen atom is – 328 kJmol^{-1}, hence the energy of fourth Bohr orbit would be

a) – 41 kJmol^{-1} b) – 82 kJmol^{-1} c) – 164 kJmol^{-1} d) – 1312 kJmol^{-1}

$\Rightarrow$ $E_n = -K\left(\frac{Z}{n}\right)^2$

$Z = 1$ for hydrogen , n = 2

$E_2 = -\frac{K \times 1}{4}$

$-328 = -\frac{K \times 1}{4}$

$\therefore K = 4 \times 328$

$E_4 = -\frac{K \times 1}{(4)^2} = -\frac{K}{16}$

$= -\frac{4 \times 328}{16}$

$E_4 = -82$ kJmol^{-1}

Q4) The radius of hydrogen atom in the ground state is 0.53 A^0. The radius of Li^{2+} ion (At.no.= 3) in a similar state is

a) 0.53 A^0 b) 1.06 A^0 c) 0.17 A^0 d) 0.265 A^0

$\Rightarrow$ Due to ground state, state of hydrogen atom (n = 1) , Radius of hydrogen atom (r) = 0.53 A^0 , Atomic No. (3) of Li = 3

Radius of Li^{2+} ion = $r \times \frac{n^2}{Z}$

$= 0.53 \times \frac{(1)^2}{3}$

$= 0.17\ A^0$

Q5) Calculate the radius and energy associated with the first orbit of He^+.

$\Rightarrow$ He^+ is a hydrogen like species having Z = 2

$\therefore$ $r_n = \frac{52.9(n^2)}{Z}$ pm

$\therefore \quad r_1 = \frac{52.9(1^2)}{2} = 26.45$ pm

$\therefore \quad E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2}\right)$ J

$\therefore \quad E_n = -2.18 \times 10^{-18} \left(\frac{2^2}{1^2}\right)$ J

$E_1 = -8.72 \times 10^{-18}$ J

❖ **De-Broglie's Equation :**

$\lambda = \frac{h}{mv}$, λ = wavelength of a particle

$\lambda = \frac{h}{p}$ m = mass

v = velocity

h = Plank's Constant

❖ **Heisenberg's Uncertainity Principle :**

Statement:

" It is not possible to determine simultaneously the position and momentum of an electron in an atom "

$\Delta x . \Delta p \geq \frac{h}{4\pi}$ or $\Delta x . (m. \Delta v) \geq \frac{h}{4\pi}$

Where,

Δx = Uncertainity in position

Δp = Uncertainity in momentum

As the mass of the particle increase, the uncertainty decreases.

The minimum uncertainty in the simultaneously determination of position and velocity $= \frac{h}{4\pi m}$

❖ **Numericals on Uncertainty Principle :**

Q1) If the uncertainty in the position of an electron is 0.33 pm, what will be Uncertainty in its velocity ?

$\Longrightarrow$ Given : $\Delta x = 0.33 \times 10^{-12}$ m

$\because \quad \Delta x . \Delta p \geq \frac{h}{4\pi}$ or $\Delta x . (m. \Delta v) \geq \frac{h}{4\pi}$

$\therefore \quad 0.33 \times 10^{-12} \times 9.1 \times 10^{-31} \times \Delta v = \frac{h}{4\pi}$

$\therefore \quad \Delta v = \frac{6.6 \times 10^{-34}}{4 \times 3.14 \times 0.33 \times 10^{-12} \times 9.1 \times 10^{-31}}$

$\therefore \quad \Delta v = 1.75 \times 10^{8}$ m/sec

Q2) If the Uncertainty in position of a moving particle is zero, then find out Δp.

$\Rightarrow \quad \because \quad \Delta x \,.\, \Delta p \geq \frac{h}{4\pi}$ or $\Delta x \,.\, (m.\, \Delta v) \geq \frac{h}{4\pi}$

$\therefore \quad \Delta p \geq \frac{h}{4\pi \times 0}$

$\therefore \quad \Delta p \geq \infty$

Q3) The Uncertainty in position and velocity of a particle are 10^{-10} m and 5.27×10^{-24} ms^{-1} of the particle.

$\Rightarrow$ According to Heisenberg's Uncertainty Principle,

$\because \quad \Delta x . m\Delta v = \frac{h}{4\pi}$

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

$\therefore \quad m = \frac{6.625 \times 10^{-34}}{4 \times 3.14 \times 10^{-10} \times 5.27 \times 10^{-24}}$

$m = 0.099$ kg

Q4) Calculate the Uncertainty in velocity of a cricket ball of mass 150 g of the uncertainty in its position is of the order of $1 A^{0}$.

$\Rightarrow$ Given : m = 150 g = 0.150 kg

$\lambda = 6.6 \times 10^{-34}$ $kgm^2 s^{-1}$

$1 A^{0} = 10^{-10}$

$\because \quad \Delta x . m\Delta v = \frac{h}{4\pi}$

$\therefore \quad \Delta v = \frac{h}{4\pi \,.\, \Delta x \,.\, m} = \frac{6.6 \times 10^{-34}}{4 \times 3.143 \times 10^{-10} \times 0.150} = 3.499 \times 10^{-24}$ ms^{-1}

❖ SCHRODINGER WAVE EQUATION :

Schrodinger described the wave motion of the electron in 3 – D space around the nucleus, by a mathematical equation known as Schrodinger Wave Equation.

Schrodinger Wave Equation is

$$\frac{\partial^2 \Psi}{dx^2} + \frac{\partial^2 \Psi}{dy^2} + \frac{\partial^2 \Psi}{dz^2} + \frac{8\pi^2 m}{h^2} (E - V)\ \psi = 0$$

Where,

Ψ = wave function of an electron,represents the amplitude of the electron wave

E = Total energy of the electron

V = Potential energy of the electron

m = mass of the electron

h = Plank's Constant

x, y and z = Coordinates of the position of electron .

Ψ^2 = Square of wave function gives the probability of finding an electron at various places in a given region around the nucleus and this region is called orbital.

1) The probability of finding electron at the nucleus is zero.

2) The probability of finding electron is maximum at a distance of $0.53A^{0}$ ($0.529A^{0}$) or 52.9 pm from the nucleus. It is called Bohr's first orbit.

❖ **Quantum Numbers :**

The set of four numbers which give a complete information about energy, shape, orientation and also spin of electron in an atom are called Quantum Numbers.

There are four quantum numbers.

(i) Principal quantum number (n)

(ii) Azimuthal or subsidiary quantum number (l)

(iii) Magnetic quantum number (m)

(iv) Spin quantum number (s)

(i) Principal quantumnumber (n) :

Principal quantum number indicates the principal shell or main energy level to which the electron belongs.

It gives an idea of the size of the shell and energy of the orbit. It is denoted by 'n'. It has all positive integral values such as 1,2,3,4…. and the shells or orbits are represented by letters K,L,M,N etc. With an increase of 'n' , the distance from the nucleus , size of the shell , and energy increase .

Maximum no. of electron in a shell is given by $2n^2$.

Value of n	1	2	3	4	5	6
Designation of shell	K	L	M	N	O	P
Maximum no. of e^-	2	8	18	32	50	72

(ii) Azimuthal Quantum Number (l):

This quantum number is also known as subsidiary quantum number represented by (l).

It represents the subshell to which the electron belongs. It also determines the shape of the electron cloud. It also determines the orbital angular momentum of the electron using the relation $\sqrt{l(l+1)}\frac{h}{2\pi}$ or $\sqrt{l(l+1)\hbar}$

The number of subshells present in the main shell is equal to 'n'.

The value of azimuthal quantum number 'l' depends upon the value of principal quantum number 'n' by the relation , l = 0 to (n = 1)

Number of election in a subshell = 2 (2l + 1)

Value of l	Subshell to which electron belongs	No. of electron in subshell
0	s	2
1	p	6
2	d	10
3	f	14

Maximum no. of electrons in the subshell = 2 (2l + 1)

= 4 l + 2

Shell	Value of 'n'	Permissible Value of 'l'	Possible subshell	No. of e^- in subshell	Total no. of e^- in main shell
K	1	0	1s	2	2
L	2	0 1	2s 2p	2 6	8
M	3	0 1 2	3s 3p 3d	2 6 10	18
N	4	0 1 2 3	4s 4p 4d 4f	2 6 10 14	32

(iii) Magnetic Quantum Number (m) :

This quantum number describes the behavior of electron in a magnetic field. Under the influence of external magnetic field, the electron in the given subshell orient themselves in a certain preferred regions of space around the nucleus. These are called orbitals.

These quantum number gives the number of orbitals in a given subshell. It is denoted by m . It has values from - l to+ l through zero.

m = - l,…..-2,-1,0,+1,+2,….. + l

So for a given value of , m will have ($2l$ + 1) values.

Total no. of orbitals in a principal shell 'n' is equal to n^2

Value of ‘n’	Value of ‘l’	Value of ‘m’ m = (2l + 1)	Subshell	No. of subshell	No. of orbitals (n^2)
1	0	0	1s	1	1
2	0 1	0 -1 , 0 , +1	2s 2p	1 3	4
3	0 1 2	0 -1 , 0 , +1 -2 , -1 , 0 , +1 , +2	3s 3p 3d	1 3 5	9

For s subshell, $l = 0$ .So, m = 0

It shows that s – subshell contains only one orbital because we have only one value of m orbital contained in s-subshell, is called s-orbital.

For p-subshell, $l = 1$ So, mwill have three values (m = 1 , m = 0 , m = -1). So p-subshell contains three p-orbitals. They are p_x , p_y and p_z orbitals. In the absence of magnetic field, these orbitals have different orientations along x , y and z axis but are equivalent in energy and therefore these orbitals are called as degenerated orbitals .

For d-subshell , $l = 2$. So, mwill have five values [a/c to formula {m = (2l + 1)}] i.e.(m = +2 , m = +1 , m = 0 , m = -1 , m = -2) . So, d subshell possesses five orientations in space i.e. d-subshell contains five d-orbital .These are d_{xy} , d_{yz} = d_{zx} , $d_{x^2-y^2}$ and d_{z^2} . In the absence of magnetic field, these orbitals are equivalent in energy and are called as degenerated orbitals.

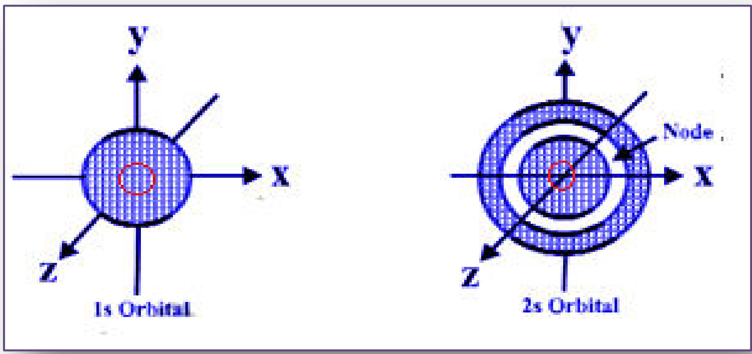
(iv) Spin Quantum Number (s) :

This quantum number describes the spin orientation of the electron. It is represented by s. Since, the electron can spin in only two ways – clockwise or anti clockwise. The probability of rotation in one direction is only 1/2 .Therefore, the spin quantum number can have only two values i.e. + 1/2 and – 1/2. For clockwise spin value is +1/2 and for anti - clockwise spin value is -1/2

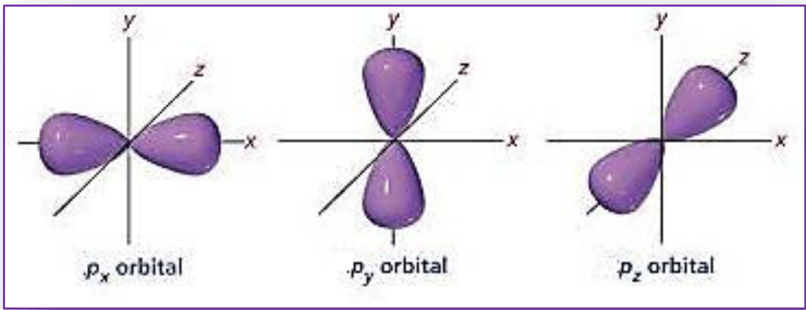
The spin of electron is also represented by arrow. A vertical arrow pointing upwards (↑) represents clockwise spin and vertical arrow pointing downwards (↓) represents anti-clockwise spin.

❖ **Shapes of Orbitals :**

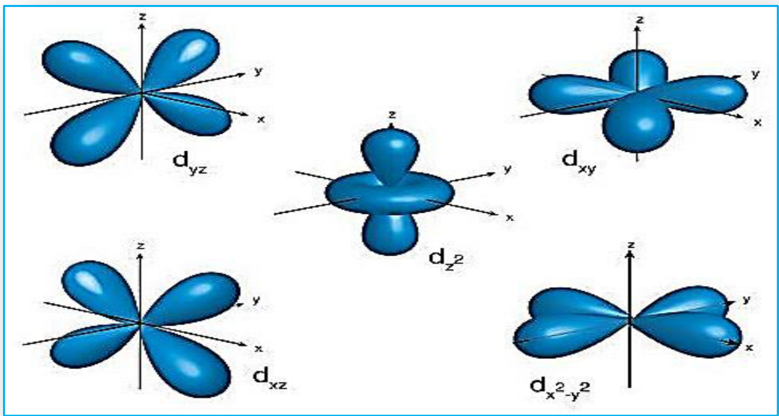
(1) Shape of S- Orbital ⇒ Spherical



(2) Shape of P – Orbital :



(3) Shape of d – Orbital :



❖ Node and Nodal plane :

Spherical or radial node :

A spherical surface within an orbital where the probability of finding the electron is zero, is called Spherical or radial node.

Number of Spherical or radial nodes in an orbital = ($n - l - 1$)

i.e. Number of nodes in orbital = ($n - l - 1$)

For Ex:

1) For 1s orbital, ($n = 1$, $l = 0$)

∴ No. of spherical / radial node = ($1 - 0 - 1$) = zero

2) For 2s orbital, ($n = 2$, $l = 0$)

∴ No. of spherical / radial node = ($2 - 0 - 1$) = one

3) For 2p orbital, ($n = 2$, $l = 1$)

∴ No. of spherical / radial node = ($2 - 1 - 1$) = zero

4) For 3p orbital, ($n = 3$, $l = 1$)

∴ No. of spherical / radial node = ($3 - 1 - 1$) = one

❖ Nodal Plane (Angular or Non spherical node) :

A plane passing through the nucleus on which the probability of finding an electron is zero is called a Nodal Plane. (or Angular or Non spherical node)

Total no. of nodal plane = Value of Azimuthal Quantum Number 'l'

(angular or Non spherical node in an orbital)

Orbital	Value of 'l'	No. of nodal planes (No. of angular node)
s	0	0 (No nodal plane)
p	1	One
d	2	Two
f	3	Three

❖ Pauli's Exclusion Principle :

Statement :

"No two electron in an atom (orbital) can have all the four quantum numbers same. Out of four, three quantum numbers may be same but the fourth quantum number will be different having opposite spin."

Consider helium atom, which has two electrons in 1s orbitals.

The four quantum numbers for two electrons in 1s orbital are as follows.

Electron Number	Quantum Number				Spt of values of quantum No.
	n	l	m	n	
1^{st} Electron	1	0	0	$+\frac{1}{2}$	$(1 , 0 , 0 , +\frac{1}{2})$
2^{nd} Electron	1	0	0	$-\frac{1}{2}$	$(1 , 0 , 0 , -\frac{1}{2})$

So,

Maximum number of electrons in an orbital = 02 (two)

Two electrons in an orbital can be represented by $\boxed{\uparrow\downarrow}$ or $\boxed{\downarrow\uparrow}$

❖ Hund's Rule of Maximum Multiplicity :

Statement :

"When several orbitals of equivalent energy are available , the electron are first filled singly with parallel spin and then with pairs with anti-parallel spin.

For example, the filling of four electron in p-orbitals can be represented as

$\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$ and not as $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow}\ \boxed{\uparrow}$

1s 2s $2p_x$ $2p_y$ $2p_z$ 1s 2s $2p_x$ $2p_y$ $2p_z$

Ex : In case of oxygen , having atomic no. 8 . According to Hund's Rule of maximum multiplicity, filling of electrons into orbitals will be as follows –

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

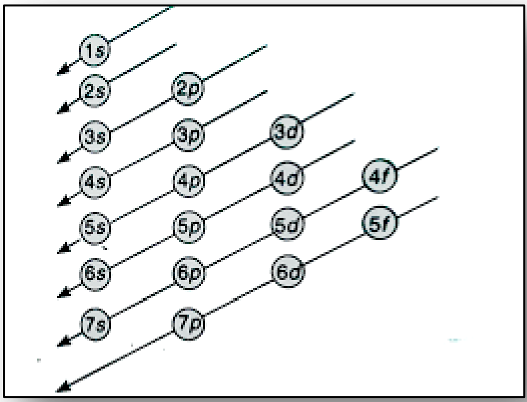
1s 2s $2p_x$ $2p_y$ $2p_z$

❖ Aufbau Principle :

Statement :

"The lower energy orbitals are first filled and then higher energy orbitals are filled successively."

The order of energy of different orbitals in an atom is given below.



1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d and so on.

❖ <u>**Electrons enter the orbitals according to (n + l) rule or Bohr Bury's rule**</u> **:**

The lower the sum (n + l) for an orbital, the lower is its energy. If two orbitals have the same (n + l) values, then the orbital with the lower values, thenof 'n' is of lower energy. This is called the (n + l) rule or Bohr – Bury's rule .

$\Longrightarrow$ The lower the sum (n + l) for an orbital, the lower is its energy. If two orbitals have the same (n + l) values, then the orbital with the lower value of 'n' is of lower energy. This is called the (n + l) rule or Bohr – Bury's rule.

Orbital energy		Principal Quantum number (n)	Azimuthal Quantun number (l)	(n + l)
1 s	↓ **Increases**	n = 1	l = 0	1 + 0 = 1
2 s		n = 2	l = 0	2 + 0 = 2
2 p		n = 2	l = 1	2 + 1 = 3, n = 2 (lower)
3s		n = 3	l = 0	3 + 0 = 3, n = 3 (higher)
3 p		n = 3	l = 1	3 + 1 = 4, n = 3 (lower)
4s		n = 4	l = 0	4 + 0 = 4, n = 4 (higher)
3 d		n = 3	l = 2	3 + 2 = 5, n = 3 (lower)
4 p		n = 4	l = 1	4 + 1 = 5, n = 4(higher)

❖ **Write the electronic configuration of following elements :**

(1) Lithium (Z = 3) **(2) Carbon (Z = 6)** **(3) Silicon (Z = 14)**
(4) Chlorine (Z = 17) **(5) Calcium (Z = 20)**

$\Longrightarrow$ (1) Li = $1s^2, 2s^1$ (2) C = $1s^2, 2s^2, 2p^2$ (3) Si = $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

(4) Cl = $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

(5) Ca = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$

❖ **Write the condensed orbital notation of electronic configuration of following elements :**

(1) Fluorine (Z = 9) **(2) Scandium (Z = 21)** **(3) Cobalt (Z = 27)**

(4) Zinc (Z = 30)

⟹ (1) Fluorine (Z = 9) = [He] $2s^2, 2p^5$

(2) Scandium (Z = 21) = [Ar] $4s^2, 3d^1$

(3) Cobalt (Z = 27) = [Ar] $4s^2, 3d^7$

(4) Zinc (Z = 30) = [Ar] $4s^2, 3d^{10}$

❖ **Anomalous behavior of Copper and Chromium (Exceptional Electronic Configuration of Cu and Cr)**

⟹ **Copper (Cu) has atomic no. 29**

It's excepted electronic configuration = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^9$

The d-orbital being degenerate, it acquires more stability when it is half filled ($3d^5$) or fully filled ($3d^{10}$).

The energy difference between 3d and 4s orbital is very low.

Due to inter electronic repulsion forces, one 4s electron enters into 3d orbital completely filled and 4s orbital half filled , which gives extra stability and the electronic configuration of Cu becomes , $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$

⟹ **Chromium, Cr has atomic no. 24**

It's expected electronic configuration = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^4$

The 3d orbital is less stable as it is not half filled.

Due to inter electronic repulsion forces, one 4s electron enters into 3d – orbital. This makes 4s and 3d orbitals half filled , which gives extra stability and the actual electronic configuration of Cr becomes , $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^5$

MCQ's On ATOMIC STRUCTURE

1) The hydride ion is isoelectronic with …

(a) H^+ (b) He^+ (c) He (d) Be

2) Line spectra is characteristic of ..

(a) Atoms (b) Molecules (c) Radicals (d) Compounds

3) In Hydrogen spectrum, which of the following lies in the wavelength range 350 – 700 nm ?

(a) Balmer series (b) Lyman series

(c) Bracket series (d) Paschen series

4) Splitting of spectral lines when atoms are subjected to strong electric field is called..

(a) Stark effect (b) Zeeman effect

(c) Photoelectric effect (d) Decay

5) When a strong magnetic field is applied to a source of spectrum, each spectral line gets split up into a number of separate lines . This phenomenon is called …

(a) Zeeman effect (b) Stark effect

(c) Compton effect (d) Quantum effect

6) The first emission line in the atomic spectrum of hydrogen in the Balmer series appears at …

(a) $\frac{9\,R}{400} cm^{-1}$ (b) $\frac{7\,R}{144}\, cm^{-1}$ (c) $\frac{3\,R}{4}\, cm^{-1}$ (d) $\frac{5\,R}{36}\, cm^{-1}$

7) Bracket series are produced when the electrons from the outer orbits jump to …

(a) 2^{nd} orbit (b) 3^{rd} orbit (c) 4^{th} orbit (d) 5^{th} orbit

8) In the emission spectrum of hydrogen, which electronic transition causes third line of the Balmer series ?

(a) Fourth Bohr orbit to second one (b) Fifth Bohr orbit to first one

(c) Fifth Bohr orbit to second one (d) Fourth Bohr orbit to first one

9) Number of spectral lines possible when an electron falls from fifth orbit to ground state in hydrogen atom is..

(a) 4 (b) 15 (c) 10 (d) 21

10) The orbital angular momentum of an electron in 2s orbital is..

(a) $+\frac{1}{2} \cdot \frac{h}{2\pi}$ (b) zero (c) $\frac{h}{2\pi}$ (d) $\sqrt{2} \cdot \frac{h}{2\pi}$

11) In a Bohr model of an atom , when an electron jumps from n = 3 to n = 1, how much energy will be emitted ?

(a) 2.15 x 10^{-11} ergs (b) 2.389 x 10^{-12} ergs

(c) 0.239 x 10^{-10} ergs (d) 0.1936 x 10^{-10} ergs

12) **The value of the energy for the first excited state of hydrogen atom will be ..**

(a) – 13.6 eV (b) – 3.40 eV (c) – 1.51 eV (d) – 0.85 eV

13) **The Bohr's orbit radius for the H-atom (Z = 1) is approximately 0.53A°. The radius for the first excited state orbit is ..**

(a) 0.13 A° (b) 1.06 A° (c) 4.77A° (d) 2.12A°

14) **The ratio of energy of photon of wavelength 3000 A° and 6000A° is …**

(a) 3 : 1 (b) 2 : 1 (c) 1 : 2 (d) 1 : 3

15) **The energy of the first electron in helium will be ..**

(a) 13.6 eV (b) -54 eV (c) -5.44 eV (d) zero

16) **The energies E_1 and E_2 of two radiations are 25 eV and 50 eV respectively. The relation between their wavelength i.e. λ_1 and λ_2 will be …**

(a) $\lambda_1 = \lambda_2$ (b) $\lambda_1 = 2\,\lambda_2$ (c) $\lambda_1 = 4\,\lambda_2$ (d) $\lambda_1 = \frac{1}{2}\,\lambda_2$

17) **According to the Bohr theory, which of the following transitions in the hydrogen atom will be give rise to the least energetic photon?**

(a) n = 6 to n = 1 (b) n = 5 to n = 4 (c) n = 6 to n = 5 (d) n = 5 to n= 3

18) **The values of Plank's Constant is 6.63 x 10^{-3}. The speed of light is 3 x 10^{17} nms^{-1}. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 6 x10^{15}s^{-1} ?**

(a) 25 (b) 50 (c) 75 (d) 10

19) **Energy of an electron is given by E = - 2.178 x 10^{-18} J $\left(\frac{Z^2}{n^2}\right)$ .Wavelength of light required to excite an electron in a hydrogen atom from level n = 1 to n = 2 , will be (h = 6.62 x 10^{-34}Js and C = 3 x 10^{8} ms^{-1})**

(a) 1.214 x 10^{-7} m (b) 2.816 x 10^{-7} m

(c) 6.500 x 10^{-7}m (d) 8.500x10^{-7}m

20) **Calculate the energy in joule corresponding to light of wavelength 45 nm? (Plank's Constant = 6.63 x 10^{-34}Js and C =3 x 10^{8} ms^{-1})**

(a) 6.67 x 10^{15} (b) 6.67 x10^{11}

(c) 4.42 x10^{-15} (d) 4.42 x10^{-18}

21) **For a p-electron, orbital angular moment is..**

(a) $\sqrt{2}$h (b) h (c) $\sqrt{6}$ h (d) 2 h

22) **If the radius of 2nd Bohr orbit of hydrogen atom is r_2, the radius of third Bohr orbit will be ..**

(a) 4/9 r^2 (b) 4 r^2 (c) 9/4 r^2 (d) 9 r^2

23) **The ionization energy of the ground state of hydrogen atom is 2.18 x 10^{-18}J. The energy of an electron in its second orbit will be ..**

(a) -1.09 x 10^{-18} J (b) - 2.18 x 10^{-18} J (c) – 4.36x 10^{-18}J (d) -5.45 x 10^{-18}J

24) **Which of the following set of quantum numbers is not possible for 4p electron?**

(a) n = 4 , l = 1 , m = -1 , s = +$\frac{1}{2}$ (b) n = 4 , l = 1 , m = 0 , s = +$\frac{1}{2}$

(c) n = 4 , l = 1 , m = 2 , s = +$\frac{1}{2}$ (d) n = 4 , l = 1 , m = -1 , s = -$\frac{1}{2}$

25) **For how many orbitals the quantum number n = 3 , l = 2, m = +2 are possible ?**

(a) 1 (b) 2 (c) 3 (d) 4

26) **The number of nodal planes for 'd' orbital equal to …**

(a) 1 (b) 2 (c) 3 (d) 0

27) **The correct set of four quantum numbers for the valence electrons of rubidium atom (Z = 37) is …**

(a) 5 , 1 , 1 , +$\frac{1}{2}$ (b) 5 , 0 , 1 , +$\frac{1}{2}$ (c) 5 , 0 , 0 , +$\frac{1}{2}$ (d) 5 , 1 , 0 , +$\frac{1}{2}$

28) **Which of the following set of quantum numbers belongs to highest energy ?**

(a)n = 4 , l = 0 , m = 0 , s = +$\frac{1}{2}$ (b) n = 3 , l = 0 , m = 0 , s = +$\frac{1}{2}$

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

ANSWER KEY

01) - c	02) – a	03) – a	04) – a	05) – a	06) – d
07) – c	08) – c	09) – c	10) – b	11) – d	12) – b
13) – d	14) - b	15) – b	16) – b	17) – c	18) – b
19) – a	20) – d	21) – a	22) – c	23) – d	24) – c
25) – a	26) – b	27) – c	28) – d		