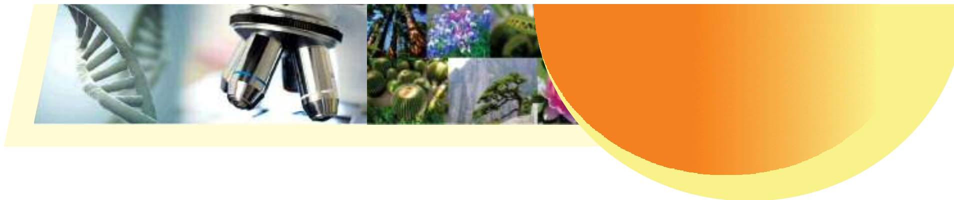
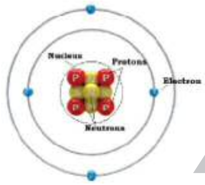




Chapter - 2 Atomic Structure

(1) Atom (Given by Dalton) : Matter is made up of extremely small particles which are indivisible in nature. It consists of subatomic particles electron, proton and neutrons knows as FUNDAMENTAL PARTICLES.

(2) Electron (Named by Stoney) : Discovered by Cathode Ray experiment [In CROOK'S TUBES]. A long glass tube with two metal electrodes. At every low pressure when high voltage is applied a flow is produced due to flow of – ve charge particle [KNOWN AS ELECTRON], cathode rays. Cathode rays have – ve change, travel in straight lines have electric and magnetic field have heating effect more penetrating effect.

Charge on e^- was found by **OIL DROP** experiment [MILLIKAN].

(3) Proton (Discovered by GOLDSTEIN in **ANODE RAY EXPT** : in a perforated cathode tube with gas at low pressure high voltage was passed b/w electrode rays from cathode produced green FLOURESCENCE on ZnS wall. These were called as ANODE RAYS. They travel in straight line, with + ve charge, get deflected in electric and magnetic field.

(4) Neutron : Fundamental particle which carries **no** charge but has mass equal to N atom or PROTON. Discovered by **JAMES CHADWICK.**

Name of Constant	UNIT	Electron	Proton	Neutron
Mass	amu	0.000546	1.00728	1.008665
	Kg	9.109×10^{-31}	1.673×10^{-27}	1.675×10^{-27}
Charge	Coloumbs	1.602×10^{-19}	$+ 1.602 \times 10^{-19}$	Zero
	esu	$- 4.8 \times 10^{-10}$	$+ 4.8 \times 10^{-10}$	Zero
	Relative	– 1	+ 1	Zero

(5) J.J. Thomson : The sphere of + ve change nucleus model of atom is balanced by coloumbic force of attraction of e^-. Like a Raisin Pudding Model.

(6) Ruther Ford (Discovery of nucleus) : $\left[{}^{4}_{2}He\right]\alpha$ particles (+ ve charge) bombared on gold foil.

(a) 99.9% passed without deflection : Most space inside the atom is empty.

(b) Only few deflected therefore mass of atom centrally placed called nucleus

(c) very few deflected back therefore mass of atom contains + ve change particles [Protons].

(d) Atom is electrically neutral hence – ve change particles placed outside the nucleus and have very less mass.

Limitations : No distribution and energies of e^- considered, could not explain e^- does not fall into the nucleus or not; no details of line spectra of H atom.

(7) (a) Atomic number (Z) : The no. of protons or electron in a neutral atom

or

No. of protons in an atom (or ion).

(b) At mass no. (A) : Total no. of protons and neutron in an atom

${}^{A}_{Z}X$ $\quad$ $A = \text{No. of } (P + n)$ $\quad$ $A - Z = \text{No. of neutrons}$

(8) (a) Isotopes : Atoms of same element with different mass no. $\left[{}^{14}_{7}N, {}^{15}_{7}Al\right]$

(b) Isobars : Atoms of different element with same mass no. $\left[{}^{40}_{20}Ca; {}^{40}_{18}Ar\right]$

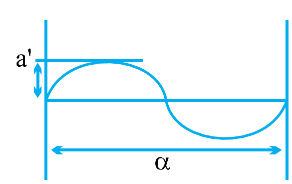
(c) Isotones : Atoms of different element with same no. of neutron $\left[{}^{23}_{11}Na; {}^{24}_{12}Mg\right]$

(d) Isoelectronic : Atoms, molecules or ions with same no. of e^- [Ne; O^{2-}].

(9) Electromagnetic radiations : Energy emitted from any source (in forms of waves) in which electric and magnetic fields oscillated perpendicular to each other and travelling with a velocity to light is known as EM radiation.

(10) Characteristics of waves :

(a) Wavelength : the distance of one crest and one trough in a wave. Denoted by ‘λ’.



(b) Frequency : no. of waves passing through a given point in one second.

Denoted by υ. $\left[\begin{array}{l} \upsilon = \frac{1}{t} \Rightarrow \sec^{-1} \text{ or Hz} \\ t = \text{Time period} \end{array}\right]$

(c) Amplitude : The height of crest or depth of a trough denoted by 'a'.

(d) Wave no. : No. of waves per unit length denoted by $\bar{\upsilon}$

$$\bar{\upsilon} = \frac{1}{\lambda} = \text{cm}^{-1} \ (\text{or m}^{-1})$$

(e) Velocity : Linear distance travelled by a wave in one second.

$$\text{velocity of light } c = \frac{\text{Distance}}{\text{Time}} = \lambda \times \frac{1}{t} = \upsilon\lambda$$

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

(10) Energywise order for EM radiation.

cosmic < γ rays < X rays < uv < VIBGYOR < IR < Microwaves < Radiowaves

$\longrightarrow$ $\longrightarrow$ $\longrightarrow$ $\longrightarrow$ $\longrightarrow$

λ Inc. υ Dec. $\bar{\upsilon}$ Dec. Energy Dec.

(11) Proton : A packet or particle of light energy is knows as **Photon.**

(12) Planck's quantum theory : The energy emitted from a source or absorbed by a source is not continuous but discontinuous in form of small packet of energy [for light each packet is known as photon]

Energy [emitted or absorbed] by a photon α υ (frequency)

$E \propto \upsilon$ or $E = h\upsilon$ [h = Planck's constant]

$= 6.626 \times 10^{-34}$ J sec

If 'n' photons are emitted $E = nh\upsilon$

(13) Photo electric effect : The phenomenon of ejection of electrons from a metal surface when a light radiation of suitable frequency falls on metal surface.

$$h\upsilon = h\upsilon_0 + \frac{1}{2}mv^2$$

$h\upsilon \Rightarrow$ Light radiation falling upon metal surface.

$h\upsilon_0 \Rightarrow$ Energy used for work function or [Energy for removing e^- from metal]

$\frac{1}{2}mv^2$ = K. energy by which e^- is emitted from metal surface.

(a) Large atoms have less work function.

$\therefore$ e^- emitted with more velocity.

(b) Small atoms have more work function.

$\therefore$ e^- emitted with less velocity.

Debroglie equation : All material particles possess both matter should also exhibit wave like properties. Wave character as well as wave character as well as wave character.

$$\lambda \propto \frac{1}{\text{Momentum}}$$

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

For microscopic particles mass is very less therefore 'λ' more and more wave character.

For macroscopic particles mass is large λ is less therefore more particle character.

Dual behaviour $\begin{bmatrix} e^- \text{ behaves as a particle as well as wave} \\ \text{Photon behaves as wave as well as particle} \end{bmatrix}$

Heisenberg's Uncertainty Principle

It is impossible simultaneously to determine the exact position and exact velocity of a subatomic particle.

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

Δx = uncertainty in position

Δv = uncertainty in momentum

For microscopic (mass very less) certainty in position is less therefore Δx is more Δv is less.

For macroscopic (large mass) certainty in position is more Δx is less Δv is more.

Bohr's theory for H [H like one e^- systems He^+; Li^{2+}]

(1) e^- revolving round the nucleus in circular path [stationey state; SHELL] with a definite angular momentum $\frac{nh}{2\pi}$ [n no. of shell of e^-] and with definite energy $\left[\frac{-2\pi^2 me^4 z^2}{n^2h^2}\right] \Rightarrow -2.18 \times 10^{-18}\frac{Z^2}{n^2}$ J/Atom.

(2) As n increases, Z Decreases Energy of e^- becomes ≤ less – ve [Due to less force of Proton attraction]

As n Decreases Z increases Energy of e^- becomes More – ve [Due to more force of attraction by protons]

(3) In infinity shell e^- has zero force of attraction therefore zero energy.

(4) Electron energy only changes by definite values $\Delta E = E_f - E_i$.

$$\frac{hc}{\lambda} = 2.18 \times 10^{-18}\left[\frac{1}{n_i^2} - \frac{1}{n_f^2}\right] Z^2 \text{ J/Atom}$$

$$\frac{1}{\lambda} = \frac{2.18 \times 10^{-18}}{\lambda}\left[\frac{1}{n_i^2} - \frac{1}{n_f^2}\right] Z^2$$

$$\left[\begin{array}{l} \text{If } n_i > n_f \text{ energy emitted by } e^- \\ \quad n_i > n_f \text{ energy absorbed by } e^- \end{array}\right]$$

$$= 109678\left[\frac{1}{n_i^2} - \frac{1}{n_f^2}\right] Z^2 \text{ cm}^{-1}$$

Rydberg constant

Hydrogen spectrum : When e^- in hydrogen atom is provided energy it gets excited to higher shell from ground state, it comes back to ground state by emitting energy in definite values.

"Quanta" : The emission of light energy is known as emission spectra. It corresponds to each atom depending upon which energy shell e^- is excited. It is **discontinuous** spectra as 'λ' of light radiations do not merge with each other like is VIBGYOR (Continous Spectra).

When e^- falls from any excited state to

$$\frac{1}{\lambda} = 1{,}09{,}678\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] Z^2$$

(a) I[st] energy level $n_f = 1$, $n_i = 2, 3, 4, ...$ [Lyman series] (UV)

(b) When e^- to final state $n = 2$, $n_i = 3, 4, 5, ...$ [Balmer series] (VIBGYOR)

(c) When e^- to falls to final state 3 $n_i = 4, 5, 6$ [Paschem series] IR.

(d) When e^- to falls to final state 4 $n_i = 5, 6, 7$ [Bracket series] IR.

(e) When e^- to falls to final state 5 $n_i = 6, 7, 8$ [Pfund series] IR.

Quantum numbers : The no. which **completely** define the **state** of e^-.

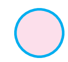
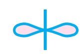
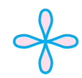
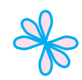
(a) Position and energy of e^- (b) Path of e^- (*c*) Orienation in space (d) Rotational motion of e^-. These are described by four quantum numbers.

(1) Principal Q. No. : It describes the distance of e^- from **nucleus '*n*'** *i.e.*, defines the **shell** no. It is denoted by '*n*'.

$n \quad = 1, 2, 3, 4, 5, ...$

K, L, M, N, O .....

(2) Azimutha Q. No. : It defines the path of e^- decided by angular momentum of e^-. Each angular momentum value corresponds to one subshell. The no. of subshells in a shell is 0 to $n - 1$.

n	l (0 to n^{-1})				
1	0	$l = 0$	'*s*'	subshell	
2	0, 1	$l = 1$	'*p*'	subshell	
3	0, 1, 2	$l = 2$	'*d*'	subshell	
4,	0, 1, 2, 3	$l = 3$	'*f*'	subshell	

All subshells are wave functions for locating e^-.

In the same shell energy wise S < P < *d* < *f*.

(3) Magnetic Q. No. : It gives the no. of magnetic orientations an e^- can have in a subshell. The no. of magnetic orientation an e^- can have in a subshell $\Rightarrow -l$ to 0 to $+l$.

(4) Spin Q. No. : An e^- is continuously spinning on its own axis. This Q. No. describes e^- can have clockwise spin motion $\left(+\frac{1}{2}\text{value}\right)$ or e^- can have anticlockwise spin motion $\left(-\frac{1}{2}\right)$. An orbital can have mximum two e^- one with clockwise and other with anticlockwise spin.

Aufbau principle

(a) e^- are filled in increasing order of energy of subshell.

(b) As '$n + l$' value inc. energy of e^- increases in that subshell.

(c) For two subshells with some '$n + l$' value. As '*n*' value increases energy of e^- increases.

Pauli's principle

No two e^- can have same set of 4 quantum nos. If two e^- are present in same shell, subshell, orbital they will have different spin value.

Hund's rule

The pairing of e^- in degenerate orbitals (different orbitals with same energy) will get paired only once they have been singly occupied.

Important Points

The filling of e^- in subshells follows this order. (As per Aufbau principle)

(A) $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$

(B) Half filled and completely filled subshells have more **stability** than incompletely filled subshells.

(24) Cr = [Ar] $4s^2\ 3d^4$ changes to Cr = [Ar] $4s^1\ 3d^5$

(29) Cu = [Ar] $4s^2\ 3d^9$ changes to Cu = [Ar] $4s^1\ 3d^{10}$

(C) As the shell no. inc. size of subshell increases *e.g.*, size of ($2s > 1s$) : ($3p > 2p$); ($4d > 3d$)

(D) The region is an orbital for no probability of finding the e^- is known as **Nodal plane** (or Node).

The no. of [Spherical nodes or radical nodes] = $n - l - 1$.

(E) **ψ(psi)**	**ψ^2(psi square)**
A wave function for locating an e^-	The square of wave function where the probability of finding the e^- is maximum. [Each value of ψ^2 is a region and **defines** one orbital]

(F) **Orbit**	**Orbital**
(1) A definite distance from the nucleus for finding the e^- [e^- as a particle]	(1) A probability region for locating the e^- around the nuclues. It is a wave function [e^- as a wave]
(2) It has definite size and e^- in this orbit has definite energy.	(2) It does not define definite size. But only a boundry region diagram of a wave for locating the e^-.

STRUCTURE OF ATOM

A. Fundamental particles of an atom, Electromagnetic Spectrum & Bohr's Theory

1-Mark Questions

Q. 1. Give the difference in the origin of cathode and anode rays.

Q. 2. Which of the following will not show deflection from the path on passing through an electric field ? Proton, electron and neutron.

Q. 3. Neutron or proton of same wavelength, which will have more velocity ?

Q. 4. Chlorophyll absorbs radiations of wavelength 700 nm. What is the frequency of this radiation ? [**Ans.**4.2 × 10^s Hz]

Q. 5. If the length of the crest of a wave is 4 pm. Write the wavelength of this wave. [**Ans.**8 pm]

Q. 6. Arrange the following in the increasing order of their e/m values.

electron, proton and neutron.

Q. 7. A radiation emitted from a hot iron is photon or quantum ?

Q. 8. Give two examples from everyday life where cathode rays tube used.

Q. 9. The line spectrum of an element is known as fingerprints of its atom. Comment

Q. 10. An element lias atomic number 30 and mass number 66, what will be the number of protons and neutrons in this atom?

[**Ans.**protons = 30 & neutrons = 36]

Q. 11. Calculate the mass and charge of one mole of electrons. [**Ans.**5.48 × 10^{-4} g]

Q. 12. Calculate the charge of one mole of electrons. [**Ans.**9.6488 × 10^4 C]

Q. 13. What is the value of the Bohr's radius for the first orbit of hydrogen atom ?

Q. 14. Distinguish between a photon and a quantum,

Q. 15. What type of metals are used in photoelectric cell ? Give one example.

[**Ans.**With large size less work function.]

Q. 16. Write the number of electrons present in 1 mol. of N^{3-} ion.

Q. 17. Write the difference between atomic mass and atomic mass number ?

Q. 18. Which series of lines of the hydrogen spectrum lie in the visible region ?

Q. 19. Cs show maximum photoelectric effect, why ?

Q. 20. CI and CI^- have same atomic number or mass number or both. Explain.

2-Mark questions

Q. 1. How could Rutherford in his experiment explain Nucleus ?

Q. 2. Define black body and black body radiations.

Q. 3. Give the essential postulates of Bohr's model of an atom. Howdid it explain ?

(i) the stability of the atom ?

(ii) origin of the spectral lines in H-atom ?

Q. 3. What is quantisation ? How quantisation of energy was introduced in Bohr's model ?

Q. 4. An element with mass number 81 contain 31.7% more neutron as compared to protons.

Assign the atomic symbol. **[Ans.** ^{81}**Br]**

Q. 5. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+spectrum?

[**Ans.** $n_1 = 1$ and $n_2 = 2$]

Q. 6. What transition of Li^{2+} spectrum will have the same wavelength as that of the second line of Balmer series in He^+spectrum ?

[**Ans.** $n_2 = 6$ to $n_1 = 3$]

Q. 7. Calculate the energy required for the process

$$He^+(g) \longrightarrow He^{2+}(g) + e^-$$

The ionization energy for the H atom in the ground state is 2.18×10^{-18} J atom^{-1}] [**Ans.** 8.72×10^{-18} J]

Q. 8. Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen. [**Ans.** 1.523×10^6 m^{-1}]

Q. 9. To which orbit the electron in H atom will jump on absorbing 12.1 eV energy ? [**Ans.** 3rd orbit]

Q. 10. Calculate the energy associated with the first orbit of He^+. What is the radius of this orbit. [**Ans.** – 54.38 eV, 0.2645 Å]

Q. 11. What is the distance of separation between 3rd and 4th orbit of H-atom? [**Ans.** 3.703 Å]

Q. 12. The energy of electron in the first Bohr's obit is – 13.6 eV. Calculate the energy of electron in the first excited state. [**Ans.** – 3.4 eV]

Q. 13. Calculate the number of protons emitted in 10 hours by a 60 W sodium lamp emitting radiations of wavelength 6000 Å.

Q. 14. Which one has a higher energy, a photon of violet light with wavelength 4000 Å or a proton of red light with wavelength 7000 Å ?

[Given. $h = 6.62 \times 10^{-34}$ J sec.]

Q. 15. A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of protons emitted per second by the bulb.

[**Ans.** $2.012 \times 10^{20}\ s^{-1}$]

Q. 16. What is the maximum number of emission lines when the excited electron of a H atom in $n = 4$ drops to the ground state ? [**Ans.** 6]

Q. 17. Which has more energy, light radiation of wavelength 400 Pm or light radiation of frequency 10^{15} Hz ?

Q. 18. Find the energy of electron in 4th shell of Li^{2+} ion.

Q. 19. What is the wave number of an electron with shortest wavelength radiation in Lyman spectrum of He^{+} ion.

Q. 20. Write short note on :

(a) Continuous and discontinuous spectrum.

(b) Absorbtion and emission spectrum.

3-Mark Questions

Q. 1. How were cathode rays discovered ? With the help of suitable experiments show that :

(i) Cathode rays tavel in straight lines,

(ii) Cathode rays consists of material particles,

(iii) Cathode rays consists of negatively charged particles.

Q. 2. Write the complete symbol for :

(i) the nucleus with atomic number 56 and mass number 138.

(ii) the nucleus with atomic number 26 and mass number 55.

(iii) the nucleus with atomic number 4 and mass number 9.

Q. 3.(i) The energy associated with the first orbit in the hydrogen atom is -2.18×10^{-18} J atom^{-1}. What is the energy associated with the fourth orbit ?

(ii) Calculate the radius of Bohr's third orbit for hydrogen atom.

[**Ans.**– 1.36×10^{-19} J atom^{-1}.4.761 nm]

Q. 4. A bulb emits light of wave length 4500Å. The bulb is rated as 150 watt and 8% of the energy is emitted as light. How many photons are emitted by the bulb per second ? [H.O.T.S1 [**Ans.** $n = 27.2 \times 10^{18}$]

Q. 5. When light with a wavelength of 400 nm falls on the surface of sodium, electrons with a kinetic energy of 1.05×10^{5} J mol^{-1} are emitted.

(a) What is the minimum energy needed to remove on electron from sodium ?

(b) What is the maximum wavelength of light that will cause a photoelectron to be emitted ?

[**Ans.** $a = 3.2255 \times 10^{19}$ J, $b = 616$ nm]

Q. 6. Compare the frequency of light radiations emitted when electron falls from 5th shell to the 2nd shell in Li^{2+} ion and electron falls from 4th shell to the 1st shell in He^{+} ion.

5- Mark Questions

Q. 1.(a) Define Photoelectric effect ? Mention its one practical application in daily life,

(b) Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (ν_o) and work function (W_o) of the metal.

[**Ans.** $\nu_o = 4.41 \times 10^{14}$ s^{-1} $W_o = 2.92 \times 10^{-19}$ J]

Q. 2.(a) The electronic energy in Bohr's orbit is negative .How will you account for it

(b) The ionisation energy of hydrogen atom is 13.6 eV. What will be the energy of the first orbit of He^{+} and Li^{2+} ions ?

[**Ans.** E_1 of He^{+} = – 54.4 eV, E_1 of Li^{2+} = – 122.4 eV]

Q. 3.(a) Define the following terms :

(i) Threshold frequency (ii) Work function.

(b) The work function for Cs atom is 1 .9 eV. Find threshold wavelength (λ_0)and threshold fequency (ν_0) of this light radiation. If Cs metal is irradiated with a radiation of wavelength 500 nm find kinetic energy and velocity of emitted electron.

Dual Nature of Matter, de Broglie Equation,

Heisenberg's uncertainty principle and wave mechanical model

1 - Mark Questions

Q. 1. Mention the physical significance of ψ and ψ^2.

Q. 2. Why did Heisenberg replaces the concept of definite orbit by the concept of probability ? (e^- with very less mass more as wave)

Q. 3. What is uncertain in uncertainty principle ?

Q. 4. Can a moving cricket ball have a wave character ? Justify your answer.

Q. 5. Heisenberg uncertainty principle has no significane in our everyday life. Explain.

Q. 6. Out of the *d* orbitals which does not have four lobes ?

Q. 7. Write the Schrodinger wave equation.

Q. 8. Why uncertainty in position is more when uncertainty in velocity is less for an electron ?

2-Mark Questions

Q. 1. Calculate the mass of the photon with wavelength of 3 .6 Å.

[**Ans.** 6.135×10^{-29} kg]

Q. 2. Calculate the mass of the photon with wavelength of 5 Pm.

Q. 3. On the basis of uncertainty principle show that an electron cannot exist with in atomic nucleus. *(Given : Nuclear radius = 10^{-15} m)*

*[**Hint :** Taking 10^{-15} m as Δx, the Δv comes much higher than the velocity of light and hence is not possible]*

Q. 4. Explain why the uncertainty principle is significant only from the motion of subatomic paricles and is negligible for macroscopic particles.

Q. 5. List two differences between orbit and orbital .

Q. 6. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit

Q.7. Comment on "Bohr's model is against the Heisenberg uncertainty principle".

Q. 8. What are the similiarties and difference in 2*s* and 2*px* orbitals and 1s and 2*s* orbitals ?

Q. 9. Give three differences between 3*p* and 4*p* subshell.

3-Mark Questions

Q. 1. Calculate the number of waves made by Bohr electron in one complete revolution in its third orbit. **[H.O.T.S]** [**Ans.** 3]

Q. 2. What should be the ratio of velocities of CH_4 and O_2 molecules so that they are associated with de Broglie waves of equal wavelength ? [**Ans.** 2]

Q. 3. Calculate the wavelength of an electron that has been accelerated in a particle accelerator through a potential difference of 1 kv.

[Given 1eV = 1.6 × 10^{-19} J] **[H.O.T.S]** [**Ans.** 3.87 × 10^{-7} m]

Q. 4.(i) Discuss the similiarities and differences between a 1s and 2s orbital.

(ii) Draw the shape of dz^2.

Q. 5. Calculate the wavelength of a tennis ball of mass 60 gm moving with a velocity of 10 m per second. [**Ans.** 10^{-3} metre]

Q. 6. Calculate the wavelength of 1000 kg rocket moving with a velocity of 3000 Kilometre per hour. [**Ans.** 7.9512 × 10^{-40} m]

Q. 7. Calculate the uncertainity in the velocity of a cricket ball of mass 150 g, if uncertainity in its position is of the order of 1 Å.

[**Ans.** 3.5 × 10^{-24} m s^1]

Q. 8.(a) What is de-Broglie wavelength for an electron with light velocity?

(b) What is the angular monemtum of electron in 5th shell.

Q. 9. Two particles A and B have wavelength $\lambda_A = 5 \times 10^{-10}$ m and $\lambda_B = 10 \times 10^{10}$ m. Find their frequency, wave number and energies. Which has more penetrating power and why ?

Q. 10.(a)Which has max. uncertainty regarding position and why ?

Electron, proton and neutron.

(b) Find the number of waves associated with a light radiation of time period 5 ns.

Q. 11. If an electron in He^+ has angular momentum of $5h/2\pi$. Find its energy and wavelength associated with it. Find the kinetic energy of this electron.

5-Mark Questions

Q. 1.(a) State de Broglie equation. Write its significance.

(b) A beam of helium atoms moves witha velocity of 2.0×10^3 m s^{-1}. Find the wavelength of the particle constituting the beam

($h = 6.626 \times 10^{-34}$ J s) [**Ans.** 49.9 pm]

Q. 2.(a) State Heisenbergs uncertainty principle. Give its mathematical expression. Also give its significance.

(b) Calculate the uncertainity in the position of a dust particle with mass equal to 1 mg if the uncertainity in its velocity is $5.5 \times 10^{-20}ms^{-1}$.

[**Ans.** 9.55×10^{10} m]

Q. 3.(a) Cricket ball, a tennis ball and a proton which has more uncertainity in velocity and which follows Heisenberg uncertainity principle maximum.

(b) Whet is same in de-Broglie and Heisenberg principle, which is diffent from Bohr theory for structure of atom.

(c) Why Schrondinger wave only defines electron path uncertain but definite energy.

Quantum Numbers, Electronic Configuration and Nodes

1 Mark Questions

Q. 1. What is the lowest value of n that allows g orbitals to exist ?

Q. 2. What are the four quantum numbers of 19^{th} electron of copper ?

(**Ans.** Atomic number = 29)

Q. 3. Which quantum number is not obtained from solution of Schrondinger wave equation ?

Q. 4. How many electrons will be present in the sub-shells having m_s, value of – 1/2 for $n = 4$?

Q. 5. Write the electronic configuration of Ni^{2+}. (At. no. of Ni = 28)

Q. 6. How many radial and angular nodes are present in 2*p* orbital.

[**Ans.** Radial nodes = 0, Angular nodes = 1]

Q. 7. Which of the following orbitals are possible ?

1*p*, 2*s*, 2*p* and 3*f*.

Q. 8. Write the name of non-directional orbital.

Q. 9. Write the name of quantum numbers which determines the orientation of orbital and shape of orbitals.

Q. 10. Write the name of quantum number which determines the shape of orbital.

Q. 11. Using *s*, *p*, *d* notations, describe the orbital with the following quantum numbers :

(a) $n = 4$, $1 = 2$ (b) $n = 1$, $l = 0$. [**Ans.** (a) 4*d* (b) 1*s*]

Q. 12. How many orbitals are presenting subshell ?

Q. 13. How many total electrons can be filled in all orbitals with $(n + l) = 5$? [**Ans.** 18 electrons ($4p^6\,3d^{10}\,5s^2$)]

Q. 14. Name the dipositive ion represented by the electronic configuration : $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$. [$Ca^{2+}$]

Q. 15. Is it correct to say that every atom with even atomic number has all electron paired ?

Q. 16. Cr in ground state has how many unpaired electrons.

(Cr, Atomic number = 24).

Q. 17. Which has more energy of electron 4*p* or 5*s* ?

Q. 18. Nitrogen has correct configuration of $1s^2$, $2s^2$, $2p_x^{\,1}$,$2p_y^{\,1}$, $2p_z^{\,1}$ is described by which principle ?

Q. 19. Whal are degenerate orbitals ?

2-Mark Questions

Q. 1. On the basis of Pauli's exclusion principle show that the maximum number of electrons in the M -shell ($n = 3$) of any individual atom is 18.

Q. 2. Designate each subshell with $n = 4$.

Q. 3. List the possible values for all the quantum numbers for the following subshell.

(a) $2p$ (b) $4f$

Q. 4. Write down the elctronic configuration of Fe^{3+} and Ni^{2+}. How many unpaired electrons are present ? (Atomic number, Fe = 26, Ni = 28)

Q. 5. Out of principal, angular, magnetic and spin quantum number, which quantum number determines the ?

(a) Shape of the orbital

(b) Number of orbitals in an orbit

(c) Size of the orbital

(d) Spin orientation of the electron.

Q. 6. What is the Hund 's rule of maximum multiplicity ? Explain with suitable example.

Q. 7. Explain why :

(a) The three electron present in 2p subshell of nitrogen remain unpaired

(b) Cr has configuration $3d^5\ 4s^1$ and not $3d^4\ 4s^2$.

Q. 8. (a) $4p^1$ and $4p^2$ electron will have same (i) n, l, m or (ii) n, l, s or (iii) $n, l, m, s,$ or n, 1

(b) Nitrogen has 7 proton, 7 electron and 7 neutrons. Calculate the number of electron protons and neutrons in N^{3-} ion.

Q. 9. Which has more energy of electron :

(a) Last electron of Cl^- or last electron of O^{2-}.

(b) $n = 4, l = 3$ or $n = 5, l = 2$.

3-Mark Questions

Q. 1.(i) An atomic orbital has $n = 2$. What are the possible values of l and m_1 ?

(ii) List the quantum numbers (m_1 and l) of electrons for $3d$ orbital.

(iii) Which of the following orbitals are possible ?

$2d$, $1s$, $2p$ and $3f$.

Q. 2.(a) Write the maximum nunber of electron in a subshell with $l = 3$ and $n = 4$.

(b) Write the maximum number of electron that can be associated with the following set of quantum numbers ?

$[n = 3, l = 1 \text{ and } m_1 = -1]$

(c) Write the maximum nunber of electron that can be accomadated in an atom in which the highest principal quantum number value is 4.

Q. 3.(i) Write the electronic configurations of the following ions :

(a) H^- (b)Na^+ (c)O^{2-} (d) F^-

(ii) What are the atomic numbers of elements whose outermost electrons are represented by (a) $3s^1$(b) $2p^3$ and (c) $3p^5$?

(iii) Which atoms are indicated by the following configurations ?

(a) [He] $2s^1$ (b) [Ne] $3s^2$ $3p^3$ (c) [Ar] $4s^2$ $3d^1$.

Q. 4. Calculate:

(a) Total number of spherical nodes in $3p$ orbital.

(b) Total number of Nodal planes in $3p$ orbital.

(c) What is nodal planes in $3d$ orbital,

5-Mark Questions

Q. 1.(a) Write short notes on:

(i) Aufbau principle (ii) Pauli's principle (iii) Hund's rule.

(b) Write the electronic configuration of the following ions :

(i) Fe^{3+} (ii) Cu^+ [Atomic number of Fe and Cu are 26 & 29]

Q. 2.(a) Draw the shapes of the following orbitals.

(i) $3dxy$ (ii) dz^2

(b) What is the total number of orbitals associated with the principal quantum number $n = 3$?

(c) Using s, p, d, f notations, describe the orbital with the following quantum numbers.

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

Q .3. Explain the following :

(i) Energy of electron is not decided by : n, l ,m and s.

(ii) Maximum number of electron with – 1/2 spin for $n = 3$ is 6,9,12 or none.

(iii) Maximum number of electron can be present for $n + l = 4$.

(iv) $3f$ subshell is not possible.

(v) Maximum number of electrons in a subshell is :

$(2l + 1)$ or $(4l + 1)$ or n^2

Q. 4. (a) A neutral atom has 2K, 8L & 15 M electrons. Find the total numbers of electrons in s, p, d and f subshell.

(b) How many unpaired electrons are present in the following ions :

Al^{+}, Cr^{2+}, Co^{3+} and Mn^{2+}

(Atomic number : Al=13, Cr = 24, Co = 27 & Mn = 25)

(c) One electron is present in 4f subshell. What is the sum of $n + l + m_1 + m_s$ values assuming 'f' subshell follows – 3 to + 3 order of filling electron.

Q. 5. Write the :

(a) $n + l$ value for 14th electron in an atom.

(b) Increasing order of filling electron in $4f$, $5p$ and $6d$ subshells.

(c) 'm' and 'l' value for last electron of Mg atom.

(Mg atomic number is 12)

(d) Subshell in which last electron is present in Ga.

(Atomic number is 31)

(e) Sum of spin of all the electron in atomic number 14.