2 Structure of Atom

Facts that Matter

• Discovery of Electron—Discharge Tube Experiment

In 1879, William Crooks studied the conduction of electricity through gases at low pressure. He performed the experiment in a discharge tube which is a cylindrical hard glass tube about 60 cm in length. It is sealed at both the ends and fitted with two metal electrodes as shown in Fig. 2.1.

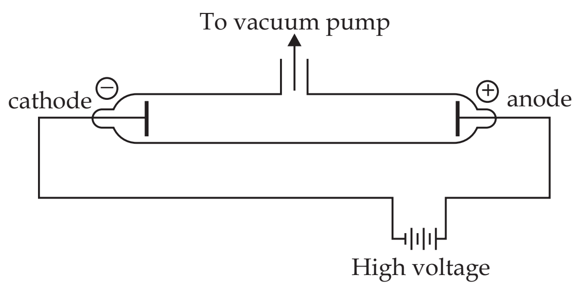

Fig. 2.1 *A cathode ray discharge tube.*

The electrical discharge through the gases could be observed only at very low pressures and at very high voltages.

The pressure of different gases could be adjusted by evacuation. When sufficiently high voltage is applied across the electrodes, current starts flowing through a stream of particles moving in the tube from the negative electrode (cathode) to the positive electrode (anode). These were called cathode rays or cathode ray particles.

• Properties of Cathode Rays

(*i*) Cathode rays travel in straight line.

(*ii*) Cathode rays start from cathode and move towards the anode.

(*iii*) These rays themselves are not visible but their behaviour can be observed with the help of certain kind of materials (fluorescent or phosphorescent) which glow when hit by them.

(*iv*) Cathode rays consist of negatively charged particles. When electric field is applied on the cathode rays with the help of a pair of metal plates, these are found to be deflected towards the positive plate indicating the presence of negative charge.

(*v*) The characteristics of cathode rays do not depend upon the material of electrodes and the nature of gas present in the cathode ray tube.

• Determination of Charge/Mass (e/m) Ratio for Electrons

J. J. Thomson for the first time experimentally determined charge/mass ratio called e/m ratio for the electrons. For this, he subjected the beam of electrons released in the discharge tube as cathode rays to influence the electric and magnetic fields. These were acting perpendicular to one another as well as to the path followed by electrons.

According to Thomson, the amount of deviation of the particles from their path in presence of electrical and magnetic field depends upon following factors:

(*i*) Greater the magnitude of the charge on the particle, greater is the interaction with the electric or magnetic field and thus greater is the deflection.

(*ii*) The mass of the particle—lighter the particle, greater the deflection.

(*iii*) The deflection of electrons from their original path increases with the increase in the voltage across the electrodes or strength of the magnetic field.

By carrying out accurate measurements on the amount of deflections observed by the electrons on the electric field strength or magnetic field strength, Thomson was able to determine the value of

$$e/m_e = 1.758820 \times 10^{11} \text{ C kg}^{-1}$$

where m_e = Mass of the electron in kg

e = magnitude of charge on the electron in coulomb (C).

• Charge on the Electron

R.A. Millikan devised a method known as oil drop experiment to determine the charge on the electrons.

$$\text{Charge of an electron } (e) = -1.6022 \times 10^{-19} \text{ C}$$

$$\text{The mass of electron } (m_e) = \frac{e}{e/m_e} = \frac{1.6022 \times 10^{-19}\,\text{C}}{1.758820 \times 10^{11}\,\text{C kg}^{-1}} = 9.1094 \times 10^{-31} \text{ kg}$$

• Discovery of Proton—Anode Rays

In 1886, Goldstein modified the discharge tube by using a perforated cathode. On reducing the pressure, he observed a new type of luminous rays passing through the holes or perforations of the cathode and moving in a direction opposite to the cathode rays. These rays were named as positive rays or anode rays or as canal rays.

Anode rays are not emitted from the anode but from a space between anode and cathode.

• Properties of Anode Rays

(*i*) The value of positive charge (e) on the particles constituting anode rays depends upon the nature of the gas in the discharge tube.

(*ii*) The charge to mass ratio of the particles is found to depend on the gas from which these originate.

(*iii*) Some of the positively charged particles carry a multiple of the fundamental unit of electrical charge.

(*iv*) The behaviour of these particles in the magnetic or electric field is opposite to that observed for electron or cathode rays.

• Proton

The smallest and lightest positive ion was obtained from hydrogen and was called proton.

$$\text{Mass of proton} = 1.676 \times 10^{-27} \text{ kg}$$

$$\text{Charge on a proton} = (+)\ 1.602 \times 10^{-19} \text{ C}$$

- **Neutron**

It is a neutral particle. It was discovered by Chadwick (1932).
By the bombardment of thin sheets of beryllium with fast moving α-particles he observed that highly penetrating rays consist of neutral particles which were named neutrons.

- **Thomson Model of Atom**

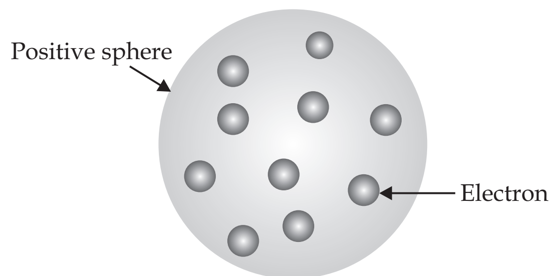

Fig. 2.2 *Thomson model of atom*

(*i*) J. J. Thomson proposed that an atom may be regarded as a sphere of approximate radius 10^{-8} cm carrying positive charge due to protons and in which negatively charged electrons are embedded.
(*ii*) In this model, the atom is visualized as a pudding or cake of positive charge with electrons embedded into it.
(*iii*) The mass of atom is considered to be evenly spread over the atom according to this model.

Drawback of Thomson Model of Atom

This model was able to explain the overall neutrality of the atom, it could not satisfactorily, explain the results of scattering experiments carried out by Rutherford in 1911.

- **Rutherford's α-particle Scattering Experiment**

Rutherford in 1911, performed some scattering experiments in which he bombarded thin foils of metals like gold, silver, platinum or copper with a beam of fast moving α-particles. The thin gold foil had a circular fluorescent zinc sulphide screen around it. Whenever α-particles struck the screen, a tiny flash of light was produced at that point.
From these experiments, he made the following observations:

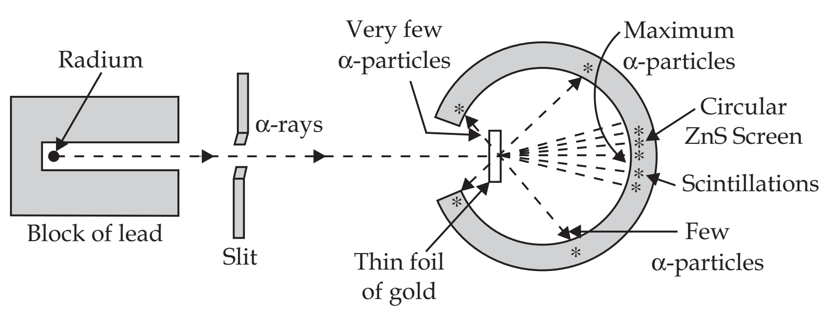

Fig. 2.3 *Rutherford's scattering experiment.*

(*i*) Most of the α-particles passed through the foil without undergoing any deflection.
(*ii*) A few α-particles underwent deflection through small angles.
(*iii*) Very few mere deflected back *i.e.*, through an angle of nearly 180°.

From these observations, Rutherford drew the following conclusions:

(*i*) Since most of the α-particles passed through the foil without undergoing any deflection, there must be sufficient empty space within the atom.
(*ii*) A small fraction of α-particles was deflected by small angles. The positive charge has to be concentrated in a very small volume that repelled and deflected a few positively charged α-particles. This very small portion of the atom was called nucleus.
(*iii*) The volume of nucleus is very small as compared to total volume of atom.

• Rutherford's Nuclear Model of an Atom

(*i*) The positive charge and most of the mass of the atom was densely concentrated in an extremely small region. This very small portion of the atom was called nucleus by Rutherford.
(*ii*) The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called orbits.
(*iii*) Electrons and nucleus are held together by electrostatic forces of attraction.

• Atomic Number

The number of protons present in the nucleus is equal to the atomic number (z). For example, the number of protons in the hydrogen nucleus is 1, in sodium atom it is 11, therefore, their atomic numbers are 1 and 11. In order to keep the electrical neutrality, the number of electrons in an atom is equal to the number of protons (atomic number, z). For example, number of electrons in hydrogen atom and sodium atom are 1 and 11 respectively.

Atomic Number (z) = Number of protons in the nucleus of an atom.
= Number of electrons in a neutral atom.

• Mass Number

Number of protons and neutrons present in the nucleus are collectively known as nucleons. The total number of nucleons is termed as mass number (A) of the atom.

Mass Number (A) = Number of protons (p) + Number of neutrons (n).

• Isotopes

Atoms with identical atomic number but different atomic mass number are known as Isotopes.

Isotopes of Hydrogen:

Name of the isotope	*Formula of the isotope*	*Mass number (A)*	*Number of protons (p)*	*Number of neutrons (n)*
Protium	${}_1^1H(H)$	1	1	0
Deuterium	${}_1^2H(D)$	2	1	1
Tritium	${}_1^3H(T)$	3	1	2

These three isotopes are shown in the figure below:

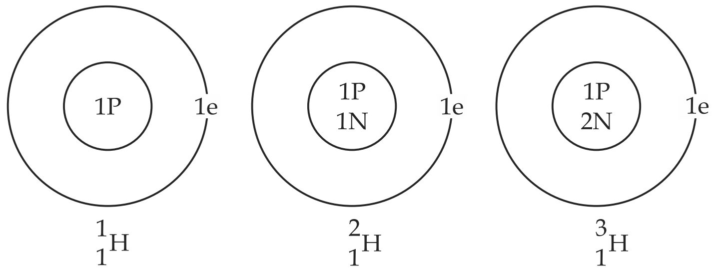

Fig. 2.4. *The three isotopes of hydrogen.*

Isotopes of Chlorine: There are two isotopes of chlorine with mass numbers 35 and 37.

Formula of the isotope	*Mass number (A)*	*Number of protons (p)*	*Number of neutrons (n)*
$^{35}_{17}Cl$	35	17	18
$^{37}_{17}Cl$	37	17	20

The two isotopes differ in their number of neutrons, having 18 and 20 neutrons, respectively.

Isotopes of some common elements

Element	*Isotopes*	*Element*	*Isotopes*	*Element*	*Isotopes*
Carbon (*C*)	$^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$	Oxygen (*O*)	$^{16}_{8}O$, $^{17}_{8}O$, $^{18}_{8}O$	Uranium	$^{235}_{92}U$, $^{238}_{92}U$, $^{239}_{92}U$
Nitrogen (*N*)	$^{14}_{7}N$, $^{15}_{7}N$	Chlorine (*Cl*)	$^{35}_{17}Cl$, $^{37}_{17}Cl$	Sulphur	$^{32}_{16}S$, $^{33}_{16}S$, $^{34}_{16}S$, $^{36}_{16}S$

Characteristics of Isotopes

(*i*) Since the isotopes of an element have the same atomic number, but different mass number, the nuclei of isotopes contain the same number of protons, but different number of neutrons.

(*ii*) Since, the isotopes differ in their atomic masses, all the properties of the isotopes depending upon the mass are different.

(*iii*) Since, the chemical properties are mainly determined by the number of protons in the nucleus, and the number of electrons in the atom, the different isotopes of an element exhibit similar chemical properties. For example, all the isotopes of carbon on burning give carbon dioxide.

• Isobars

Isobars are the atoms with same mass number but different atomic number, for example, $^{14}_{6}C$ and $^{14}_{7}N$.

Another example is $^{40}_{18}Ar$, $^{40}_{19}K$ and $^{40}_{20}C$ are typical isobars.

Each of these have same mass number but different atomic number.

• Drawbacks of Rutherford Model

(*i*) When a body is moving in an orbit, it achieves acceleration. Thus, an electron moving around nucleus in an orbit is under acceleration.

According to Maxwell's electromagnetic theory, charged particles when accelerated must emit electromagnetic radiations. Therefore, an electron in an orbit will emit radiations, the energy carried by radiation comes from electronic motion. Its path will become closer to nucleus and ultimately should spiral into nucleus within 10^{-8} s. But actually this does not happen.

Thus, Rutherford's model cannot explain the stability of atom if the motion of electrons is described on the basis of classical mechanics and electromagnetic theory.

(*ii*) Rutherford's model does not give any idea about distribution of electrons around the nucleus and about their energies.

• Developments Leading to the Bohr's Model of Atom

Two developments played a major role in the formulation of Bohr's model of atom. These were:

(*i*) Dual character of the electromagnetic radiation which means that radiations possess both wave like and particle like properties.

(*ii*) Experimental results regarding atomic spectra which can be explained only by assuming quantized electronic energy levels in atoms.

• Nature of Electromagnetic Radiation (Electromagnetic Wave Theory)

This theory was put forward by James Clark Maxwell in 1864. The main points of this theory are as follows:

(*i*) The energy is emitted from any source (like the heated rod or the filament of a bulb through which electric current is passed) continuously in the form of radiations and is called the radiant energy.

(*ii*) The radiations consist of electric and magnetic fields oscillating perpendicular to each other and both perpendicular to the direction of propagation of the radiation.

(*iii*) The radiations possess wave character and travel with the velocity of light 3×10^8 m/sec.

(*iv*) These waves do not require any material medium for propagation. For example, rays from the sun reach us through space which is a non-material medium.

• Characteristics of a Wave

Wavelength: It is defined as the distance between any two consecutive crests or troughs. It is represented by λ and its S.I. unit is metre.

$$1 \text{ Å} = 10^{-10} \text{ m.}$$

Frequency: Frequency of a wave is defined as the number of waves passing through a point in one second. It is represented by ν (*nu*) and is expressed in Hertz (Hz).

$$1 \text{ Hz} = 1 \text{ cycle/sec.}$$

Velocity: Velocity of a wave is defined as the linear distance travelled by the wave in one second.

It is represented by *c* and is expressed in cm/sec or m/sec.

Amplitude: Amplitude of a wave is the height of the crest or the depth of the through. It is represented by '*a*' and is expressed in the units of length.

Wave Number: It is defined as the number of waves present in 1 metre length. Evidently it will be equal to the reciprocal of the wavelength. It is represented by $\bar{\nu}$ (read as nu bar).

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

Electromagnetic Spectrum: When electromagnetic radiations are arranged in order of their increasing wavelengths or decreasing frequencies, the complete spectrum obtained is called electromagnetic spectrum.

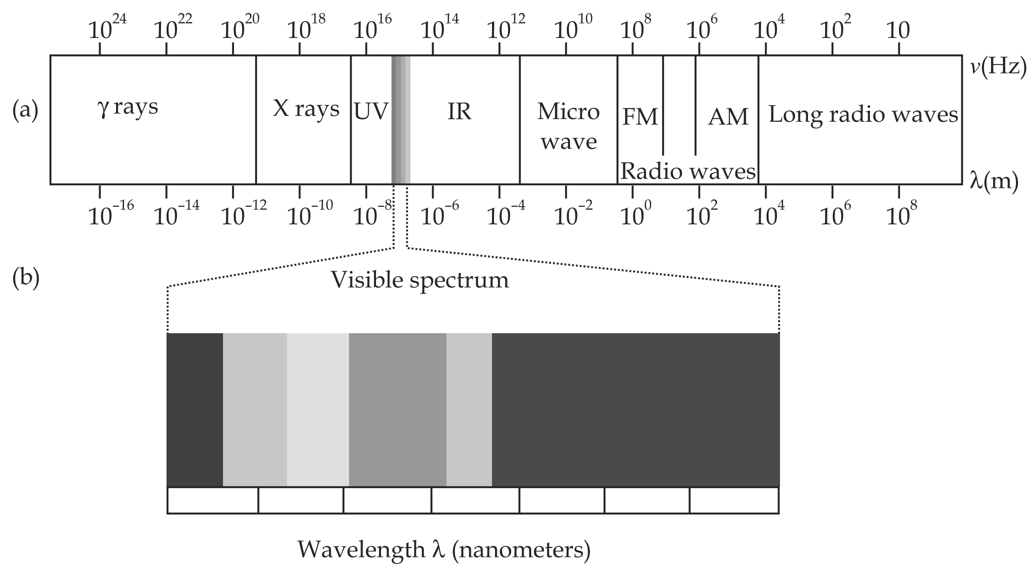

Fig. 2.5 *(a) The spectrum of electromagnetic radiation. (b) Visible spectrum. The visible region is only a small part of the entire spectrum.*

• Limitations of Electromagnetic Wave Theory

Electromagnetic wave theory was successful in explaining properties of light such as interference, diffraction etc; but it could not explain the following:

(*i*) The phenomenon of black body radiation.
(*ii*) The photoelectric effect.
(*iii*) The variation of heat capacity of solids as a function of temperature.
(*iv*) The line spectra of atoms with reference to hydrogen.

• Black Body Radiation

The ideal body, which emits and absorbs all frequencies is called a black body and the radiation emitted by such a body is called black body radiation. The exact frequency distribution of the emitted radiation from a black body depends only on its temperature.

At a given temperature, intensity of radiation emitted increases with decrease of wavelength, reaches a maximum value at a given wavelength and then starts decreasing with further decrease of wavelength as shown in Fig 2.6.

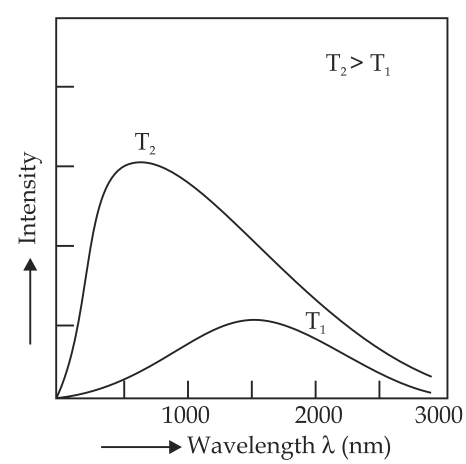

Fig. 2.6 *Wavelength-intensity relationship.*

• Planck's Quantum Theory

To explain the phenomenon of 'Black body radiation' and photoelectric effect, Max Planck in 1900, put forward a theory known as Planck's Quantum Theory.

This theory was further extended by Einstein in 1905. The main points of this theory was as follows:

(*i*) The radiant energy emitted or absorbed in the form of small packets of energy. Each such packets of energy is called a quantum.

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

$$E \propto \nu$$

or $$E = h\nu$$

where h is a proportionality constant, called Planck's constant. Its value is equal to 6.626×10^{-34} Jsec.

• Photoelectric Effect

Hertz, in 1887, discovered that when a beam of light of certain frequency strikes the surface of some metals, electrons are emitted or ejected from the metal surface. The phenomenon is called photoelectric effect.

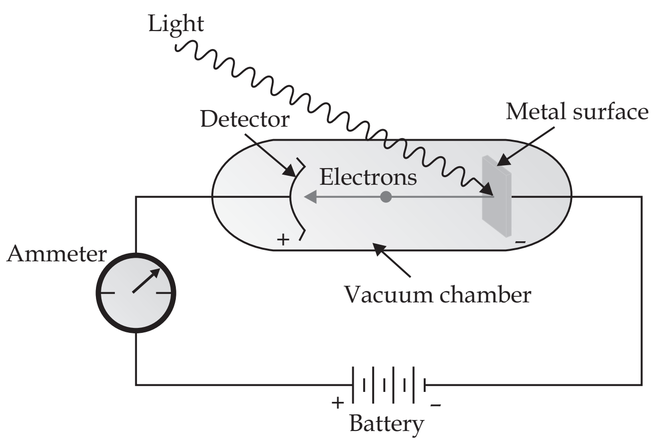

Fig. 2.7 *Equipment for studying the photoelectric effect. Light of a particular frequency strikes a clean metal surface inside a vacuum chamber. Electrons are ejected from the metal and are counted by a detector that measures their kinetic energy.*

Observations in Photoelectric Effect

(*i*) Only photons of light of certain minimum frequency called threshold frequency (ν_0) can cause the photoelectric effect. The value of ν_0 is different for different metals.

(*ii*) The kinetic energy of the electrons which are emitted is directly proportional to the frequency of the striking photons and is quite independent of their intensity.

(*iii*) The number of electrons that are ejected per second from the metal surface depends upon the intensity of the striking photons or radiations and not upon their frequency.

Explanation of Photoelectric Effect

Einstein in (1905) was able to give an explanation of the different points of the photoelectric effect using Planck's quantum theory as under:

(*i*) Photoelectrons are ejected only when the incident light has a certain minimum frequency (threshold frequency v_0)

(*ii*) If the frequency of the incident light (v) is more than the threshold frequency (v_0), the excess energy ($hv - hv_0$) is imparted to the electron as kinetic energy.

K.E. of the ejected electron

$$\frac{1}{2}\text{mv}^2 = hv - hv_0$$

or

$$hv = hv_0 + \frac{1}{2}mv^2$$

$$hv = W_0 + \frac{1}{2}mv^2$$

hence, the greater is the frequency of the incident light, the greater is the kinetic energy of the emitted electron.

(*iii*) On increasing the intensity of light, more electrons are ejected but the energies of the electrons are not altered.

• Dual Behaviour of Electromagnetic Radiation

From the study of behaviour of light, scientists came to the conclusion that light and other electromagnetic radiations have dual nature. These are wave nature as well as particle nature. Whenever radiation interacts with matter, it displays particle like properties in contrast to the wavelike properties (interference and diffraction) which it exhibits when it propagates. Some microscopic particles, like electrons, also exhibit this wave-particle duality.

• Spectrum

When a ray of white light is passed through a prism the wave with shorter wavelength bends more than the one with a longer wavelength. Since ordinary white light consists of waves with all the wavelengths in the visible range, array of white light is spread out into a series of coloured bands called spectrum. The light of red colour which has longest wavelength is deviated the least while the violet light, which has shortest wavelength is deviated the most.

Continuous Spectrum

When a ray of white light is analysed by passing through a prism it is observed that it splits up into seven different wide bands of colours from violet to red (like rainbow). These colours are so continuous that each of them merges into the next. Hence, the spectrum is called continuous spectrum.

Emission Spectra

Emission Spectra is noticed when the radiations emitted from a source are passed through a prism and then received on the photographic plate. Radiations can be emitted in a number of ways such as:

(*i*) from sun or glowing electric bulb.

(*ii*) by passing electric discharge through a gas at low pressure.

(*iii*) by heating a substance to high temperature.

Line Spectra

When the vapours of some volatile substance are allowed to fall on the flame of a Bunsen burner and then analysed with the help of a spectroscope. Some specific coloured lines appear on the photographic plate which are different for different substances. For example, sodium or its salts emit yellow light while potassium or its salts give out violet light.

Absorption Spectra

When white light is passed through the vapours of a substance and the transmitted light is then allowed to strike a prism, dark lines appear in the otherwise continuous spectrum. The dark lines indicate that the radiations corresponding to them were absorbed by the substance from the white light. This spectrum is called absorption spectrum.

Dark lines appear exactly at the same positions where the lines in the emission spectra appear.

• Line Spectrum of Hydrogen

When electric discharge is passed through hydrogen gas enclosed in discharge tube under low pressure and the emitted light is analysed by a spectroscope, the spectrum consists of a large number of lines which are grouped into different series. The complete spectrum is known as hydrogen spectrum.

On the basis of experimental observations, Johannes Rydberg noted that all series of lines in the hydrogen spectrum could be described by the following expression:

$$\bar{v} = 109,677\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]\text{cm}^{-1}$$

where $n_1 = 1, 2$ ....

$n_2 = n_1 + 1, n_1 + 2$ ....

The value 109, 677 cm^{-1} is called the Rydberg constant for hydrogen.

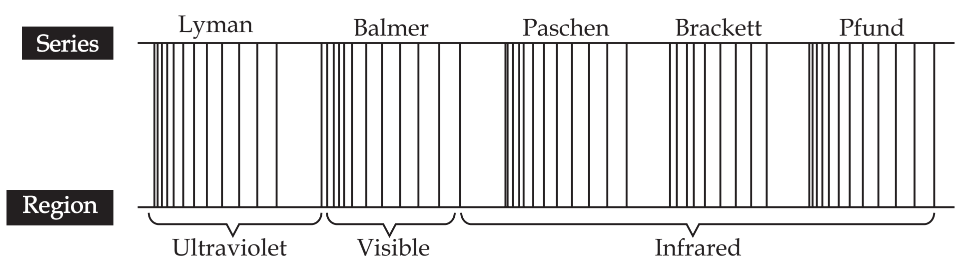

Fig. 2.8 *Atomic spectrum of hydrogen.*

Rydberg in 1890, and has given a simple theoretical equation for the calculation of wavelengths and wave numbers of the spectral lines in different series of hydrogen spectrum. The equation is known as **Rydberg formula (or equation).**

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

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* This relation is valid for hydrogen atom only. For other species,

$$\frac{1}{\lambda} = \bar{v} = R \times Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where Z is the atomic number of the species.

Here R_H = constant, called Rydberg constant for hydrogen and n_1, n_2 are integers ($n_2 > n_1$)

For any particular series, the value of n_1 is constant while that of n_2 changes. For example,

For *Lyman series*, $n_1 = 1$, $n_2 = 2, 3, 4, 5.....$
For *Balmer series*, $n_1 = 2$, $n_2 = 3, 4, 5, 6.....$
For *Paschen series*, $n_1 = 3$, $n_2 = 4, 5, 6, 7.....$
For *Brackett series*, $n_1 = 4$, $n_2 = 5, 6, 7, 8.....$
For *Pfund series*, $n_1 = 5$, $n_2 = 6, 7, 8, 9.....$

Thus, by substituting the values of n_1 and n_2 in the above equation, wavelengths and wave number of different spectral lines can be calculated. When $n_1 = 2$, the expression given above is called **Balmer's formula.**

• Bohr's Model of Atom

Niels Bohr in 1913, proposed a new model of atom on the basis of Planck's Quantum Theory. The main points of this model are as follows:

(*i*) In an atom, the electrons revolve around the nucleus in certain definite circular paths called orbits.

(*ii*) Each orbit is associated with definite energy and therefore these are known as energy levels or energy shells. These are numbered as 1, 2, 3, 4 ..... or *K, L, M, N*.....

(*iii*) Only those energy orbits are permitted for the electron in which angular momentum of the electron is a whole number multiple of $\frac{h}{2\pi}$

Angular momentum of electron (mvr) = $\frac{nh}{2\pi}$ (n = 1, 2, 3, 4 ..... etc).

m = mass of the electron.
v = tangential velocity of the revolving electron.
r = radius of the orbit.
h = Planck's constant.

n is an integer.

(*iv*) As long as electron is present in a particular orbit, it neither absorbs nor loses energy and its energy, therefore, remains constant.

(*v*) When energy is supplied to an electron, it absorbs energy only in fixed amounts as quanta and jumps to higher energy state away from the nucleus known as excited state. The excited state is unstable, the electron may jump back to the lower energy state and in doing so, it emits the same amount of energy. ($\Delta E = E_2 - E_1$).

• Achievements of Bohr's Theory

1. Bohr's theory has explained the stability of an atom.
2. Bohr's theory has helped in calculating the energy of electron in hydrogen atom and one electron species. The mathematical expression for the energy in the nth orbit is,

$$E_n = -\frac{2\pi^2 m_e e^4 Z^2}{n^2 R^2}$$

By substituting the values of,

m_e (mass of electron)
e (charge of electron)
R (Rydberg constant)
Z (Atomic number)

The value comes out to be,

$$E_n = -\frac{2.178 \times 10^{-18} \times Z^2}{n^2} \text{ J per atom}$$

$$= -\frac{1312 \times Z^2}{n^2} \text{ kJ mol}^{-1}$$

For hydrogen electron, $Z = 1$

$$E_n = -\frac{2.178 \times 10^{-18}}{n^2} \text{ J per atom}$$

$$= -\frac{1312}{n^2} \text{ kJ mol}^{-1}$$

3. Bohr's theory has explained the atomic spectrum of hydrogen atom.

- **Limitations of Bohr's Model**

(*i*) The theory could not explain the atomic spectra of the atoms containing more than one electron or multielectron atoms.

(*ii*) Bohr's theory failed to explain the fine structure of the spectral lines.

(*iii*) Bohr's theory could not offer any satisfactory explanation of Zeeman effect and Stark effect.

(*iv*) Bohr's theory failed to explain the ability of atoms to form molecule formed by chemical bonds.

(*v*) It was not in accordance with the Heisenberg's uncertainty principle.

- **Dual Behaviour of Matter (de Broglie Equation)**

de Broglie in 1924, proposed that matter, like radiation, should also exhibit dual behaviour i.e., both particle like and wave like properties. This means that like photons, electrons also have momentum as well as wavelength.

From this analogy, de Broglie gave the following relation between wavelength (λ) and momentum (p) of a material particle.

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

where m = mass of the particle

v = velocity of particle

P = momentum of the particle

This relationship has been verified by an experiment.

- **Heisenberg's Uncertainty Principle**

It states that, "It is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron".

Mathematically, it can be given as,

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

or $$\Delta x \times \Delta (mv_x) \geq \frac{h}{4\pi}$$

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

where Δx is the uncertainty in position and ΔP_x (or ΔV_x) is the uncertainty in momentum (or velocity) of the particle and h is Planck's constant.

- **Significance of Uncertainty Principle**

(*i*) It rules out existence of definite paths or trajectories of electrons and other similar particles.

(*ii*) The effect of Heisenberg's uncertainty principle is significant only for microscopic objects and is negligible for macroscopic objects.

- **Reasons for the Failure of Bohr Model**

(*i*) The wave character of the electron is not considered in Bohr Model.

(*ii*) According to Bohr Model an orbit is a clearly defined path and this path can completely be defined only if both the position and the velocity of the electron are known exactly at the same time. This is not possible according to the Heisenberg's uncertainty principle.

- **Quantum Mechanical Model of Atom**

Quantum mechanics: Quantum mechanics is a theoretical science that deals with the study of the motions of the microscopic objects that have both observable wave like and particle like properties.

Important Features of Quantum Mechanical Model of Atom

(*i*) The energy of electrons in atom is quantized *i.e.,* can only have certain values.

(*ii*) The existence of quantized electronic energy level is a direct result of the wave like properties of electrons.

(*iii*) Both the exact position and exact velocity of an electron in an atom cannot be determined simultaneously.

(*iv*) An atomic orbital has wave function ψ. There are many orbitals in an atom. Electron occupy an atomic orbital which has definite energy. An orbital cannot have more than two electrons. The orbitals are filled in increasing order of energy. All the information about the electron in an atom is stored in orbital wave function ψ.

(*v*) The probability of finding electron at a point within an atom is proportional to square of orbital wave function *i.e.,* $|\psi^2|$ at that point. It is known as probability density and is always positive.

From the value of ψ^2 at different points within atom, it is possible to predict the region around the nucleus where electron most probably will be found.

- **Quantum Numbers**

Atomic orbitals can be specified by giving their corresponding energies and angular momentums which are quantized (*i.e.,* they have specific values). The quantized values can be expressed in terms of quantum number. These are used to get complete information about electron *i.e.*, its location, energy, spin etc.

Principal Quantum Number (*n*)

It is the most important quantum number since it tells the principal energy level or shell to which the electron belongs. It is denoted by the letter '*n*' and can have any integral value except zero, *i.e.,* n = 1, 2, 3, 4 ..... etc.

The various principal energy shells are also designated by the letters, *K, L, M, N, O, P* ..... etc. **Starting from the nucleus**.

The principal quantum number gives us the following information:

(*i*) It gives the average distance of the electron from the nucleus.

(*ii*) It completely determines the energy of the electron in hydrogen atom and hydrogen like particles.

(*iii*) The maximum number of electrons present in any principal shell is given by $2n^2$ where n is the number of the principal shell.

Azimuthal or Subsidiary or Orbital Angular Quantum Number (l)

It is found that the spectra of the elements contain not only the main lines but there are many fine lines also present. This number helps to explain the fine lines of the spectrum.

The azimuthal quantum number gives the following information:

(*i*) The number of subshells present in the main shell.

(*ii*) The angular momentum of the electron present in any subshell.

(*iii*) The relative energies of various subshells.

(*iv*) The shapes of the various subshells present within the same principal shell.

This quantum number is denoted by the letter 'l'. For a given value of n, it can have any value ranging from 0 to $n - 1$. For example,

For the 1st shell (k), $n = 1$, l can have only one value *i.e.,* $l = 0$

For $n = 2$, the possible value of l can be 0 and 1.

Subshells corresponding to different values of l are represented by the following symbols:

value of l	0	1	2	3	4	5	.....
Notation for subshell	*s*	*p*	*d*	*f*	*g*	*h*	.....

Magnetic Orbital Quantum Number (m or m_l)

The magnetic orbital quantum number determines the number of preferred orientations of the electrons present in a subshell. Since each orientation corresponds to an orbital, therefore, the magnetic orbital quantum number determines the number of orbitals present in any subshell.

The magnetic quantum number is denoted by letter m or m_l and for a given value of l, it can have all the values ranging from $- l$ to $+ l$ including zero.

Thus, for energy value of l, m has $2l + 1$ values.

For example,

For $l = 0$ (*s*-subshell), m_l can have only one value *i.e.,* $m_l = 0$.

This means that *s*-subshell has only one orientation in space. In other words, *s*-subshell has only one orbital called *s*-orbital.

Spin Quantum Number (S or m_s)

This quantum number helps to explain the magnetic properties of the substances. A spinning electron behaves like a micromagnet with a definite magnetic moment. If an orbital contains two electrons, the two magnetic moments oppose and cancel each other.

• Shapes of *s*-orbitals

s-orbital is present in the *s*-subshell. For this subshell, $l = 0$ and $m_l = 0$. Thus, s-orbital with only one orientation has a spherical shape with uniform electron density along all the three axes.

The probability of 1s electron is found to be maximum near the nucleus and decreases with the increase in the distance from the nucleus. In 2s electron, the probability is also maximum near the nucleus and decreases to zero probability. The spherical empty shell for 2s electron is called nodal surface or simply node.

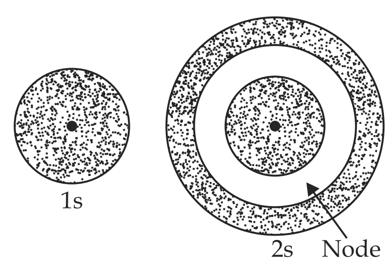
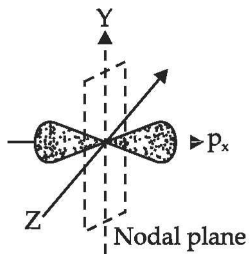


Fig. 2.9 *Shapes of 1s and 2s orbitals.*

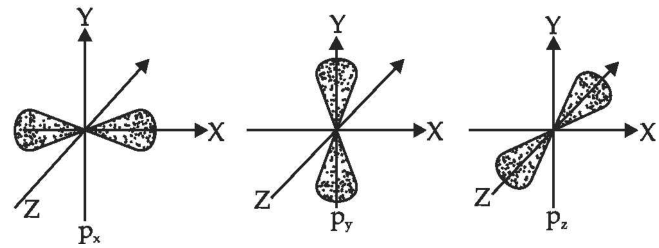
• Shapes of *p*-orbitals

p-orbitals are present in the *p*-subshell for which l = 1 and m_l can have three possible orientations – 1, 0, + 1.

Thus, there are three orbitals in the *p*-subshell which are designated as p_x, p_y and p_z orbitals depending upon the axis along which they are directed. The general shape of a *p*-orbital is dumb-bell consisting of two portions known as lobes. Moreover, there is a plane passing through the nucleus along which finding of the electron density is almost nil. This is known as nodal plane as shown in the fig.



General shape of *p*-orbital



Shapes of $2p_x$, $2p_y$ and $2p_z$ orbitals.

Fig. 2.10

From the dumb-bell pictures, it is quite obvious that unlike *s*-orbital, a *p*-orbital is directional in nature and hence it influences the shapes of the molecules in the formation of which it participates.

• Shapes of *d*-orbitals

d-orbitals are present in *d*-subshell for which l = 2 and m_l = –2, –1, 0, +1 and +2. This means that there are five orientations leading to five different orbitals.

These five orientations are designated as d_{zy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$ and d_{z^2}. However, they have the same energy *i.e.*, are in degeneracy state and are known as degenerate orbitals. The first three orbitals have clover leaf shape and lie in different planes which are *xy, yz* and *zx* planes respectively. The $d_{x^2-y^2}$ orbital is also clover leaf shaped but its lobes are directed along the *X* and *Y*-axis.

• Aufbau Principle

The principle states: In the ground state of the atoms, the orbitals are filled in order of their increasing energies.

In other words, electrons first occupy the lowest energy orbital available to them and enter into higher energy orbitals only after the lower energy orbitals are filled.

The order in which the energies of the orbitals increase and hence the order in which the orbitals are filled is as follows:

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, 5*f*, 6*d*, 7*p* .....

The order may be remembered by using the method given in fig. 2.11.

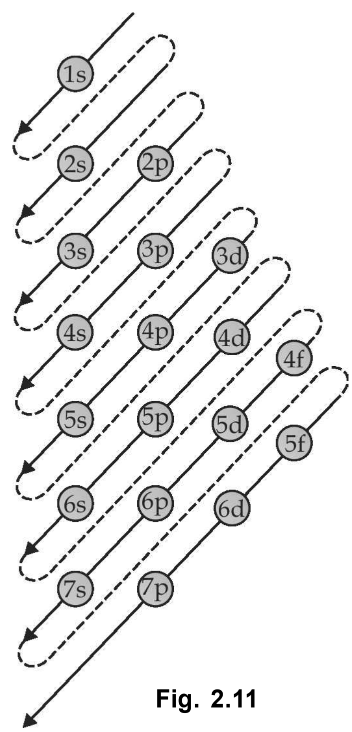


Fig. 2.11

- **Pauli Exclusion Principle**

According to this principle, no two electrons in an atom can have the same set of four quantum numbers.
Pauli exclusion principle can also be stated as: Only two electrons may exist in the same orbital and these electrons must have opposite spins.

- **Hund's Rule of Maximum Multiplicity**

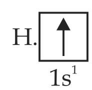
It states that: pairing of electrons in the orbitals belonging to the same subshell (*p, d* or *f*) does not take place until each orbital belonging to that subshell has got one electron each *i.e.,* it is singly occupied.

- **Electronic Configuration of Atoms**

The distribution of electrons into orbitals of an atom is called its electronic configuration. The electronic configuration of different atoms can be represented in two ways.
For example:

(*i*) $s^a\ p^b\ d^c$ ..... notation
(*ii*) Orbital diagram

For example: The electronic configuration of hydrogen—$1s^1$.



- **Causes of Stability of Completely Filled and Half Filled Subshells**

The completely filled and half filled subshells are stable due to the following reasons:

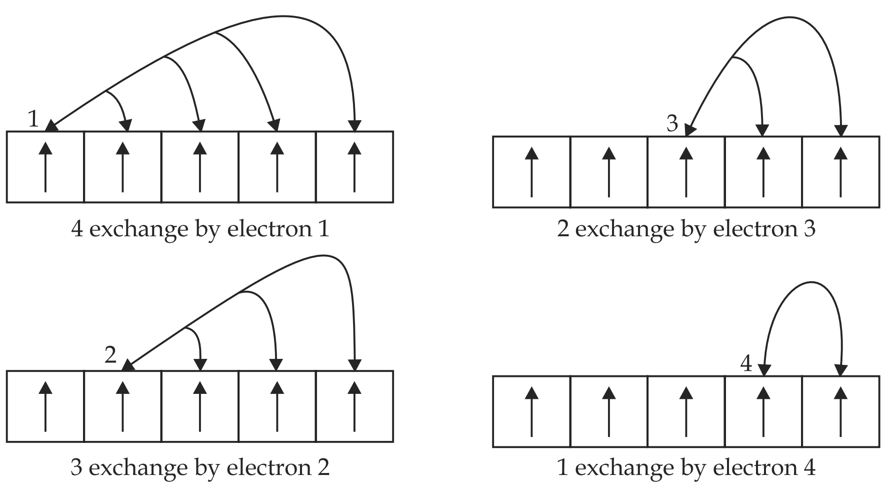


Fig. 2.12 *Possible exchange for a d^5 configuration.*

1. **Symmetrical distribution of electrons:** The completely filled or half filled subshells have symmetrical distribution of electrons in them and are therefore more stable.
2. The stabilizing effect arises whenever two or more electrons with same spin are present in the degenrate orbitals of a subshell. These electrons tend to exchange their positions

and the energy released due to their exchange is called exchange energy. The number of exchanges that can takes place is maximum when the subshell is either half filled or completely filled.

As a result the exchange energy is maximum and so is the stability.

TABLE 2.1. Electronic configurations of elements in the ground state

Atomic Number	*Element*	*Electronic Configuration*	*Atomic Number*	*Element*	*Electronic Configuration*
1.	H	$1s^1$	12.	Mg	$[Ne]^{10}\ 3s^2$
2.	He	$1s^2$	13.	Al	$[Ne]^{10}\ 3s^2\ 3p^1$
3.	Li	$[He]^2\ 2s^1$	14.	Si	$[Ne]^{10}\ 3s^2\ 3p^2$
4.	Be	$[He]^2\ 2s^2$	15.	P	$[Ne]^{10}\ 3s^2\ 3p^3$
5.	B	$[He]^2\ 2s^2\ 2p^1$	16.	S	$[Ne]^{10}\ 3s^2\ 3p^4$
6.	C	$[He]^2\ 2s^2\ 2p^2$	17.	Cl	$[Ne]^{10}\ 3s^2\ 3p^5$
7.	N	$[He]^2\ 2s^2\ 2p^3$	18.	Ar	$[Ne]^{10}\ 3s^2\ 3p^6$
8.	O	$[He]^2\ 2s^2\ 2p^4$			or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
9.	F	$[He]^2\ 2s^2\ 2p^5$	19.	K	$[Ar]^{18}\ 4s^1$
10.	Ne	$[He]^2\ 2s^2\ 2p^6$	20.	Ca	$[Ar]^{18}\ 4s^2$
11.	Na	$[Ne]^{10}\ 3s^1$	21.	Sc	$[Ar]^{18}\ 3d^1\ 4s^2$
22.	Ti	$[Ar]^{18}\ 3d^2\ 4s^2$	55.	Cs	$[Xe]^{54}\ 6s^1$
23.	V	$[Ar]^{18}\ 3d^3\ 4s^2$	56.	Ba	$[Xe]^{54}\ 6s^2$
*24.	Cr	$[Ar]^{18}\ 3d^5\ 4s^1$	*57.	La	$[Xe]^{54}\ 5d^1\ 6s^2$
25.	Mn	$[Ar]^{18}\ 3d^5\ 4s^2$	*58.	Ce	$[Xe]^{54}\ 4f^1\ 5d^1\ 6s^2$
26.	Fe	$[Ar]^{18}\ 3d^6\ 4s^2$	59.	Pr	$[Xe]^{54}\ 4f^3\ 5d^0\ 6s^2$
27.	Co	$[Ar]^{18}\ 3d^7\ 4s^2$	60.	Nd	$[Xe]^{54}\ 4f^4\ 6s^2$
28.	Ni	$[Ar]^{18}\ 3d^8\ 4s^2$	61.	Pm	$[Xe]^{54}\ 4f^5\ 6s^2$
*29.	Cu	$[Ar]^{18}\ 3d^{10}\ 4s^1$	62.	Sm	$[Xe]^{54}\ 4f^6\ 6s^2$
30.	Zn	$[Ar]^{18}\ 3d^{10}\ 4s^2$	63.	Eu	$[Xe]^{54}\ 4f^7\ 6s^2$
31.	Ga	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^1$	*64.	Gd	$[Xe]^{54}\ 4f^7\ 5d^1\ 6s^2$
32.	Ge	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^2$	65.	Tb	$[Xe]^{54}\ 4f^9\ 6s^2$
33.	As	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^3$	66.	Dy	$[Xe]^{54}\ 4f^{10}\ 6s^2$
34.	Se	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^4$	67.	Ho	$[Xe]^{54}\ 4f^{11}\ 6s^2$
35.	Br	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^5$	68.	Er	$[Xe]^{54}\ 4f^{12}\ 6s^2$
36.	Kr	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^6$	69.	Tm	$[Xe]^{54}\ 4f^{13}\ 6s^2$
		or $1s^2$, $2s^2\ 2p^6$,	70.	Yb	$[Xe]^{54}\ 4f^{14}\ 6s^2$
		$3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^6$	71.	Lu	$[Xe]^{54}\ 4f^{14}\ 5d^1\ 6s^2$
37.	Rb	$[Kr]^{36}\ 5s^1$	72.	Hf	$[Xe]^{54}\ 4f^{14}\ 5d^2\ 6s^2$
38.	Sr	$[Kr]^{36}\ 5s^2$	73.	Ta	$[Xe]^{54}\ 4f^{14}\ 5d^3\ 6s^2$
39.	Y	$[Kr]^{36}\ 4d^1\ 5s^2$	74.	W	$[Xe]^{54}\ 4f^{14}\ 5d^4\ 6s^2$
40.	Zr	$[Kr]^{36}\ 4d^2\ 5s^2$	75.	Re	$[Xe]^{54}\ 4f^{14}\ 5d^5\ 6s^2$
*41.	Nb	$[Kr]^{36}\ 4d^4\ 5s^1$	76.	Os	$[Xe]^{54}\ 4f^{14}\ 5d^6\ 6s^2$
*42.	Mo	$[Kr]^{36}\ 4d^5\ 5s^1$	77.	Ir	$[Xe]^{54}\ 4f^{14}\ 5d^7\ 6s^2$
*43.	Tc	$[Kr]^{36}\ 4d^5\ 5s^2$	*78.	Pt	$[Xe]^{54}\ 4f^{14}\ 5d^9\ 6s^1$

*44.	Ru	$[kr]^{36}\ 4d^7\ 5s^1$
*45.	Rh	$[Kr]^{36}\ 4d^8\ 5s^1$
*46.	Pd	$[Kr]^{36}\ 4d^{10}\ 5s^0$
*47.	Ag	$[Kr]^{36}\ 4d^{10}\ 5s^1$
48.	Cd	$[Kr]^{36}\ 4d^{10}\ 5s^2$
49.	In	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^1$
50.	Sn	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^2$
51.	Sb	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^3$
52.	Te	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^4$
53.	I	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^5$
54.	Xe	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^6$ or $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^6\ 4d^{10}$, $5s^2\ 5p^6$
89.	Ac	$[Rn]^{86}\ 6d^1\ 7s^2$
90.	Th	$[Rn]^{86}\ 6d^2\ 7s^2$
91.	Pa	$[Rn]^{86}\ 5f^2\ 6d^1\ 7s^2$
92.	U	$[Rn]^{86}\ 5f^3\ 6d^1\ 7s^2$
93.	Np	$[Rn]^{86}\ 5f^4\ 6d^2\ 7s^2$
94.	Pu	$[Rn]^{86}\ 5f^6\ 7s^2$
95.	Am	$[Rn]^{86}\ 5f^7\ 7s^2$
96.	Cm	$[Rn]^{86}\ 5f^7\ 6d^3\ 7s^2$
97.	Bk	$[Rn]^{86}\ 5f^9\ 7s^2$
98.	Cf	$[Rn]^{86}\ 5f^{10}\ 7s^2$
99.	Es	$[Rn]^{86}\ 5f^{11}\ 7s^2$
100.	Fm	$[Rn]^{86}\ 5f^{12}\ 7s^2$
*101.	Md	$[Rn]^{86}\ 5f^{13}\ 7s^2$
102.	No	$[Rn]^{86}\ 5f^{14}\ 7s^2$
103.	Lr	$[Rn]^{86}\ 5f^{14}\ 6d^1\ 7s^2$
104.	Rf or Ku or Unq Rutherfordium or Kurchatovium (Unnilquadium)	$[Rn]^{86}\ 5f^{14}\ 6d^2\ 7s^2$
*79.	Au	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^1$
80.	Hg	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2$
81.	Tl	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^1$
82.	Pb	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^2$
83.	Bi	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^3$
84.	Po	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$
85.	At	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^5$
86.	Rn	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$ or $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^6\ 4d^{10}\ 4f^{14}$, $5s^2\ 5p^6\ 5d^{10}$, $6s^2\ 6p^6$
87.	Fr	$[Rn]^{86}\ 7s^1$
88.	Ra	$[Rn]^{86}\ 7s^2$
105.	Ha or Db or Unp Hahnium or Dubnium (Unnilpentium)	$[Rn]^{86}\ 5f^{14}\ 6d^3\ 7s^2$
106.	Sg or Unh Seaborgium (Unnilhexium)	$[Rn]^{86}\ 5f^{14}\ 6d^4\ 7s^2$
107.	Ns or Bh or Uns Neilsbohrium or Bohrium (Unnilseptium)	$[Rn]^{86}\ 5f^{14}\ 6d^s\ 7s^2$
108.	Hs or Uno Hassnium (Unniloctium)	$[Rn]^{86}\ 5f^{14}\ 6d^6\ 7s^2$
109.	Mt or Une Meitnerium (Unnilennium)	$[Rn]^{86}\ 5f^{14}\ 6d^7\ 7s^2$
110.	Uun (Ununnilium) or Ekaplatinum	$[Rn]^{86}\ 5f^{14}\ 6d^8\ 7s^2$
111.	Uuu (Unununium) or Ekagold	$[Rn]^{86}\ 5f^{14}\ 6d^9\ 7s^2$
112.	Uub (Ununbium) or Ekamercury	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2$
113.	Uuq (Ununquadium)	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^2$
114.	Uup (Ununpentium)	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^3$

NCERT TEXTBOOK QUESTIONS SOLVED

Q1. *(i) Calculate the number of electrons which will together weigh one gram.*

(ii) Calculate the mass and charge of one mole of electrons.

Ans. (*i*)

Mass of an electron = 9.1×10^{-28} g

9.1×10^{-28} g is the mass of = 1 electron

1.0 g is the mass of = $\dfrac{1}{9.1 \times 10^{-28}}$ = **1.098×10^{27} electrons**

(*ii*)

One mole of electrons = 6.022×10^{23} electrons

Mass of 1 electron = 9.1×10^{-31} kg

Mass of 6.022×10^{23} electrons = $(9.1 \times 10^{-31}$ kg$) \times (6.022 \times 10^{23})$

= **5.48×10^{-7} kg**

Charge on one electron = 1.602×10^{-19} coulomb

Charge on one mole electrons = $1.602 \times 10^{-19} \times 6.022 \times 10^{23}$

= **9.65×10^{4} coulomb.**

Q2. *(i) Calculate the total number of electrons present in one mole of methane.*

(ii) Find (a) the total number and (b) the total mass of neutrons in 7 mg of ^{14}C. (Assume that mass of a neutron = 1.675×10^{-27} kg.)

(iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3 at STP. Will the answer change if the temperature and pressure are changed?

Ans. (*i*) One mole of methane (CH_4) has molecules = 6.022×10^{23}

No. of electrons present in one molecule of CH_4 = 6 + 4 = 10

No. of electrons present in 6.022×10^{23} molecules of CH_4 = $6.022 \times 10^{23} \times 10$

= **6.022×10^{24} electrons**

(*ii*) **Step I.** *Calculation of total number of carbon atoms*

Gram atomic mass of carbon (*C* – 14) = 14g = 14×10^{3} mg

14×10^{3} mg of carbon (*C* – 14) have atoms = 6.022×10^{23}

7 mg of carbon (*C* – 14) have atoms = $\dfrac{6.022 \times 10^{23}}{(14 \times 10^{3}\,\text{mg})} \times (7\text{ mg})$ = **3.011×10^{20}**

Step II. *Calculation of total number and total mass of neutrons*

No. of neutrons present in one atom (*C* – 14) of carbon = 14 – 6 = 8

No. of neutrons present in 3.011×10^{20} atoms (*C* – 14) of carbon = $3.011 \times 10^{20} \times 8$

= **2.408×10^{21} neutrons**

Mass of one neutron = 1.675×10^{-27} kg

Mass of 2.408×10^{21} neutrons = $(1.675 \times 10^{-27}$ kg$) \times 2.408 \times 10^{21}$

= **4.033×10^{-6} kg.**

(*iii*) **Step I.** *Calculation of total number of NH_3 molecules*

Gram molecular mass of ammonia (NH_3) = 17g = 17×10^{3} mg

17×10^{3} mg of NH_3 have molecules = 6.022×10^{23}

34 mg of NH_3 have molecules = $\dfrac{6.022 \times 10^{23}}{(17 \times 10^{3}\,\text{mg})} \times (34\text{ mg})$

= **1.2044×10^{20} molecules.**

Step II. *Calculation of total number and mass of protons*

No. of protons present in one molecule of NH_3 = 7 + 3 = 10

No. of protons present in 12.044×10^{20} molecules of NH_3 = $12.044 \times 10^{20} \times 10$

$= \mathbf{1.2044 \times 10^{22}}$ **protons**

Mass of one proton $= 1.67 \times 10^{-27}$ kg

Mass of 1.2044×10^{22} protons $= (1.67 \times 10^{-27}\text{ kg}) \times 1.2044 \times 10^{22}$

$= \mathbf{2.01 \times 10^{-5}}$ **kg.**

No, the answer will not change upon changing the temperature and pressure because only the number of protons and mass of protons are involved.

Q3. *How many protons and neutrons are present in the following nuclei?*

(i) ${}^{13}_{6}C$ *(ii)* ${}^{16}_{8}O$ *(iii)* ${}^{24}_{12}Mg$ *(iv)* ${}^{56}_{26}Fe$ *(v)* ${}^{88}_{38}Sr$

Ans. (*i*) ${}^{13}_{6}C$; Atomic no. (Z) = 6 Mass no. (A) = 13

No. of protons (p) = **6** No. of neutrons (n) = 13 – 6 = **7**

(*ii*) ${}^{16}_{8}O$; Atomic no. (Z) = 8 Mass no. (A) = 16

No. of protons (p) = **8** No. of neutrons (n) = 16 – 8 = **8**

(*iii*) ${}^{24}_{12}Mg$; Atomic no. (Z) = 12 Mass no. (A) = 24

No. of protons (p) = **12** No. of neutrons (n) = 24 – 12 = **12**

(*iv*) ${}^{56}_{26}Fe$; Atomic no (Z) = 26 Mass no. (A) = 56

No. of protons (p) = **26** No. of neutrons (n) = 56 – 26 = **30**

(*v*) ${}^{88}_{38}Sr$; Atomic no (Z) = 38 Mass no. (A) = 88

No. of protons (p) = **38** No. of neutrons (n) = **50.**

Q4. *Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)*

(i) Z = 17, A = 35 (ii) Z = 92, A = 233 (iii) Z = 4, A = 9.

Ans. (*i*) ${}^{35}_{17}X$ (*ii*) ${}^{233}_{92}X$ (*iii*) ${}^{9}_{4}X$

(X = Cl) (X = U) (X = Be)

Q5. *Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency (ν) and wave number ($\bar{\nu}$) of yellow light.*

Ans. Step I. *Calculation of frequency of yellow light*

We know that $\nu = \dfrac{c}{\lambda}$

$c = 3 \times 10^{8}\text{ m s}^{-1}$; $\lambda = 580\text{ nm} = 580 \times 10^{-9}\text{ m}$

$$\therefore \quad \nu = \frac{(3 \times 10^{8}\,\text{ms}^{-1})}{(580 \times 10^{-9}\,\text{m})} = \mathbf{5.17 \times 10^{14}\ s^{-1}}$$

Step II. *Calculation of wave number of yellow light*

$$\text{Wave number } (\bar{\nu}) = \frac{1}{\lambda} = \frac{1}{(580 \times 10^{-9}\,\text{m})} = \mathbf{1.724 \times 10^{6}\ m^{-1}.}$$

Q6. *Calculate the energy of each of the photons which*

(i) correspond to light of frequency 3×10^{15} Hz

(ii) have wavelength of 0.50 Å.

Ans. (*i*) Energy of photon (E) = $h\nu$

$$h = 6.626 \times 10^{-34} \text{ J s}; \nu = 3 \times 10^{15} \text{ Hz} = 3 \times 10^{15} \text{ s}^{-1}$$

$$E = (6.626 \times 10^{-34} \text{ J s}) \times (3 \times 10^{15} \text{ s}^{-1}) = \mathbf{1.986 \times 10^{-18}\ J}$$

(*ii*) Energy of photon (E) = $h\nu = \frac{hc}{\lambda}$

$$h = 6.626 \times 10^{-34} \text{ J s}; c = 3 \times 10^{8} \text{ ms}^{-1};$$

$$\lambda = 0.50 \text{ Å} = 0.5 \times 10^{-10} \text{ m}.$$

$$E = \frac{(6.626 \times 10^{-34} \text{ J s}) \times (3 \times 10^{8} \text{ ms}^{-1})}{(0.5 \times 10^{-10} \text{ m})} = \mathbf{3.98 \times 10^{-15}\ J.}$$

Q7. *Calculate the wavelength, frequency and wave number of light wave whose period 2.0×10^{-10} s.*

Ans.

$$\text{Frequency } (\nu) = \frac{1}{\text{Period}} = \frac{1}{(2.0 \times 10^{-10} \text{ s})} = \mathbf{5.0 \times 10^{9}\ s^{-1}}$$

$$\text{Wavelength } (\lambda) = \frac{c}{\nu} = \frac{(3 \times 10^{8} \text{ ms}^{-1})}{(5 \times 10^{9} \text{ s}^{-1})} = \mathbf{6.0 \times 10^{-2}\ m}$$

$$\text{Wave number } (\bar{\nu}) = \frac{1}{\lambda} = \frac{1}{(6.0 \times 10^{-2} \text{ m})} = \mathbf{16.66\ m^{-1}.}$$

Q8. *What is the number of photons of light with wavelength 4000 pm which provide 1 J of energy ?*

Ans. Energy of photon (E) = $\frac{hc}{\lambda}$

$h = 6.626 \times 10^{-34}$ Js, $c = 3 \times 10^{8}$ ms^{-1}, λ = 4000 pm = $4000 \times 10^{-12} = 4 \times 10^{-9}$ m

$$\therefore \text{ Energy of photon } (E) = \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^{8} \text{ ms}^{-1})}{(4 \times 10^{-9} \text{ m})} = 4.969 \times 10^{-17} \text{ J}$$

Now, 4.965×10^{-17} J is the energy of photon = 1

$$\therefore \text{ 1 J is the energy of photons} = \frac{1}{4.969 \times 10^{-17}} = \mathbf{2.012 \times 10^{16}\ photons.}$$

Q9. *A photon of wavelength 4×10^{-7} m strikes on metal surface ; the work function of metal being 2.13 eV. Calculate (i) the energy of the photon, (ii) the kinetic energy of emission, (iii) the velocity of the photoelectron. (Given 1 eV = 1.6020×10^{-19} J).*

Ans. (*i*) *The energy of the photon*

$$\text{Energy } (E) = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^{8} \text{ ms}^{-1})}{(4 \times 10^{-7} \text{ m})} = 4.97 \times 10^{-19}$$

$$= \frac{(1\ eV)}{(1.602 \times 10^{-19} \text{ J})} \times (4.97 \times 10^{-19} \text{ J}) = \mathbf{3.1\ eV}$$

(*ii*) *Kinetic energy of emission*

Kinetic energy of emission = E – work function (*i.e,* kinetic energy of emitted electron)

= (3.1 – 2.13) = 0.97 eV

(*iii*) *Velocity of photoelectron*

$$KE \text{ of emission} = \frac{1}{2}mv^2 = 0.97 \text{ eV}$$

$$= 0.97 \times 1.602 \times 10^{-19} \text{ J} = 0.97 \times 1.602 \times 10^{-19} \text{ kg m}^2 \text{ s}^{-2}$$

or $$\upsilon^2 = \frac{2 \times 0.97 \times 1.602 \times 10^{-19} \left(\text{kg m}^2 \text{ s}^{-2}\right)}{\left(9.1 \times 10^{-31} \text{ kg}\right)} = 0.34 \times 10^{12} \text{ m}^2 \text{ s}^{-2}$$

or $$\upsilon = (0.34 \times 10^{12} \text{ m}^2 \text{ s}^{-2})^{1/2} = 0.583 \times 10^6 \text{ ms}^{-1} = \mathbf{5.83 \times 10^5 \text{ ms}^{-1}}$$

Q10. *Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1}.*

Ans. $E = hc/\lambda$

λ = 242 nm = 242×10^{-9} m, $c = 3 \times 10^8$ ms^{-1}, $h = 6.626 \times 10^{-34}$ Js

$$\therefore \quad E = \frac{\left(6.626 \times 10^{-34} \text{Js}\right) \times \left(3 \times 10^8 \text{ms}^{-1}\right)}{\left(242 \times 10^{-9} \text{m}\right)} = 0.0821 \times 10^{-17} \text{ J}$$

$$\text{Ionisation energy per mol } (E) = \frac{\left(0.0821 \times 10^{-17} \text{J}\right) \times \left(6.022 \times 10^{23} \text{mol}^{-1} \text{J}\right)}{1000}$$

$$= \mathbf{494 \text{ kJ mol}^{-1}}.$$

Q11. *A 25 watt bulb emits monochromatic yellow light of wavelength 0.57 μm. Calculate the rate of emission of quanta per second.*

Ans. Energy of one photon (E) = $h\nu = hc/\lambda$

$h = 6.626 \times 10^{-34}$ J s; $c = 3 \times 10^8$ ms^{-1}; $\lambda = 0.57 \times 10^{-6}$ m

$$E = \frac{\left(6.626 \times 10^{-34} \text{ J s}\right) \times \left(3 \times 10^8 \text{ms}^{-1}\right)}{\left(0.57 \times 10^{-6} \text{m}\right)} = \mathbf{3.48 \times 10^{-19} \text{ J}}$$

$$\text{Rate of emission of quanta per second} = \frac{\text{Power}}{\text{Energy}}$$

Power (P) = 25 watt = 25 Js^{-1}; $E = 3.48 \times 10^{-19}$ J

$$= \frac{(25 \text{ watt})}{\left(3.48 \times 10^{-19} \text{ J}\right)} = \frac{\left(25 \text{ Js}^{-1}\right)}{\left(3.48 \times 10^{-19} \text{ J}\right)} = \mathbf{7.18 \times 10^{19} \text{ s}^{-1}}.$$

Q12. *Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (ν_0) and work function (W_0) of the metal.*

Ans. $$\text{Threshold frequency } (\nu_0) = \frac{c}{\lambda} = \frac{\left(3 \times 10^8 \text{ ms}^{-1}\right)}{\left(68 \times 10^{-8} \text{ m}\right)} = \mathbf{4.41 \times 10^{14} \text{ s}^{-1}}$$

Work function (W_0) = $h\nu_0$ = (6.626×10^{-34} Js) × (4.41×10^{14} s^{-1}) = $\mathbf{2.92 \times 10^{-19}}$ **J.**

Q13. *What is the wavelength of the light emitted when the electron in a hydrogen atom undergoes transition from the energy level with n = 4 to energy level with n = 2?*

Ans. According to Balmer formula

$$\text{Wave number } (\bar{\nu}) = R_H\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]\text{cm}^{-1}; \quad n_1 = 2, \quad n_2 = 4, \quad R_H = 109678 \text{ cm}^{-1}$$

$$\therefore \quad \bar{\nu} = 109678\left[\frac{1}{2^2}-\frac{1}{4^2}\right]\text{cm}^{-1} = \frac{109678\times 3}{16}\text{cm}^{-1}$$

$$\lambda = \frac{1}{\bar{\nu}} = \frac{16}{109678\times 3}\text{cm} = \frac{16\times 10^7}{109678\times 3}\text{nm} = \mathbf{486\ nm.}$$

Q14. *How much energy is required to ionise a hydrogen atom if an electron occupies n = 5 orbit? Compare your answer with the ionisation energy of H atom (energy required to remove the electron from n = 1 orbit)*

Ans. Energy for a hydrogen electron present in a particular energy shell,

$$E_n = -\frac{13.12}{n^2}\times 10^5 \text{ J mol}^{-1} = -\frac{13.12\times 10^5}{n^2\times 6.022\times 10^{23}}\text{ J atom}^{-1}$$

$$= \frac{-2.18\times 10^{-18}}{n^2}\text{ J atom}^{-1}$$

Step I. *Ionisation energy for hydrogen electron present in orbit n = 5*

$$IE_5 = E_\infty - E_5 = 0-\left(\frac{-2.18\times 10^{-18}}{25}\right)\text{J atom}^{-1} = \mathbf{8.72\times 10^{-20}\ J}$$

Step II. *Ionisation energy for hydrogen electron present in orbit n = 1.*

$$IE_1 = E_\infty - E_1 = 0-\left(\frac{-2.18\times 10^{-18}}{1}\right) = 2.18\times 10^{-18}\text{ J atom}$$

On comparing: $$\frac{IE_5}{IE_1} = \frac{\left(8.72\times 10^{-20}\text{ J atom}^{-1}\right)}{\left(2.18\times 10^{-18}\text{ J atom}^{-1}\right)} = \mathbf{0.04.}$$

IE_1 is 25 times as compared to IE_5.

Q15. *What is the maximum number of emission lines when the excited electron of a atom in n = 6 drops to the ground state?*

Ans. The maximum no. of emission lines = $\frac{n(n-1)}{2} = \frac{6(6-1)}{2}$ = 3 × 5 = **15.**

The actual transitions which are taking place are as follows:

***n* = 6 to *n* = 1**	***n* = 5 to *n* = 1**	***n* = 4 to *n* = 1**	***n* = 3 to *n* = 1**	***n* = 2**
$6 \to 5$	$5 \to 4$	$4 \to 3$	$3 \to 2$	2
$6 \to 4$	$5 \to 3$	$4 \to 2$	$3 \to 1$	
$6 \to 3$	$5 \to 2$	$4 \to 1$		
$6 \to 2$	$5 \to 1$			
$6 \to 1$				
(5 lines)	(4 lines)	(3 lines)	(2 lines)	(1 line)

Q16. (*i*) *The energy associated with first orbit in hydrogen atom is -2.18×10^{-18} J atom^{-1}. What is the energy associated with the fifth orbit?*

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

Ans. (*i*) For an electron, the energies in two orbits may be compared as:

$$\frac{E_1}{E_2} = \left(\frac{n_2}{n_1}\right)^2$$

According to available data: $n_1 = 1$, $E_1 = -2.18 \times 10^{-18}$ J atom^{-1}, $n_2 = 5$

$$\therefore \quad \frac{\left(-2.18 \times 10^{-18} \text{ J atom}^{-1}\right)}{E_2} = \left(\frac{5}{1}\right)^2 = 25$$

or $$E_5 = \frac{\left(-2.18 \times 10^{-18} \text{ J atom}^{-1}\right)}{25} = \mathbf{-8.77 \times 10^{-20}\ J}$$

(*ii*) For hydrogen atom; $r_n = 0.529 \times n^2$ Å

$r_5 = 0.529 \times (5)^2 = 13.225$ Å = **1.3225 nm.**

Q17. *Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen.*

Ans. According to Balmer formula, $\bar{\nu} = \frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$.

In order that the wavelength (λ) may be the maximum, wave number ($\bar{\nu}$) must be the least. This is possible in case $n_2 - n_1$ is minimum. Now, for Balmer series, $n_1 = 2$ and n_2 must be 3. Substituting these values in the Balmer formula,

$$\bar{\nu} = (1.097 \times 10^7 \text{ m}^{-1}) \left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 1.097 \times 10^7 \text{ m}^{-1} \left(\frac{5}{36}\right) = \mathbf{1.523 \times 10^6\ m^{-1}}.$$

Q18. *What is the energy in joules required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of light emitted when the electron returns to the ground state? The ground state electronic energy is -2.18×10^{-11} ergs.*

Ans. **Step I.** *Calculation of energy required*

The energy of electron $(E_n) = \frac{-2.18 \times 10^{-11}}{n^2}$ ergs $= \frac{-2.18 \times 10^{-18} \text{ J}}{n^2}$ ($\because$ 1 J = 10^7 ergs)

The energy in Bohr's first orbit $(E_1) = \frac{-2.18 \times 10^{-18} \text{ J}}{(1)^2} = \frac{-2.18 \times 10^{-18} \text{ J}}{1}$

The energy in Bohr's fifth orbit $(E_5) = \frac{-2.18 \times 10^{-18} \text{ J}}{(5)^2} = \frac{-2.18 \times 10^{-18} \text{ J}}{25}$

$\therefore$ Energy required $(\Delta E) = E_5 - E_1 = \left(\frac{-2.18}{25} \times 10^{-18} \text{ J}\right) - \left(-\frac{2.18}{1} \times 10^{-18} \text{ J}\right)$

$$= 2.18 \times 10^{-18} \left(1 - \frac{1}{25}\right) \text{ J}$$

$$= 2.18 \times 10^{-18} \times 24/25 = \mathbf{2.09 \times 10^{-18}\ J}$$

Step II. *Calculation of wavelength of light emitted*

$$\Delta E = h\nu = \frac{hc}{\lambda}$$

$$\therefore \qquad \lambda = \frac{hc}{\Delta E} = \frac{\left(6.626\times10^{-34}\text{ Js}\right)\times\left(3\times10^{8}\text{ ms}^{-1}\right)}{\left(2.09\times10^{-18}\text{ J}\right)}$$

$$= \mathbf{9.50 \times 10^{-8}\ m = 950\ \text{Å}.} \qquad (\because 1\ \text{Å} = 10^{-10}\text{ m})$$

Q19. *The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18})/n^2$ J. Calculate the energy required to remove an electron completely from the n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition?*

Ans. **Step I.** *Calculation of energy required*

The energy required is the difference in the energy when the electron jumps from orbit with $n = \infty$ to orbit with $n = 2$.

The energy required $(\Delta E) = E_\infty - E_2$

$$= 0 - \left(-\frac{2.18\times10^{-18}}{4}\text{ J}\right) = \mathbf{5.45 \times 10^{-19}\ J}$$

Step II. *Calculation of the longest wavelength of light in cm used to cause the transition*

$$\Delta E = h\nu = hc/\lambda$$

$$\lambda = \frac{hc}{\Delta E} = \frac{\left(6.626\times10^{-34}\text{ J s}\right)\times\left(3\times10^{8}\text{ ms}^{-1}\right)}{\left(5.45\times10^{-19}\text{ J}\right)}$$

$$= 3.644 \times 10^{-7}\text{ m} = 3.644 \times 10^{-7} \times 10^{2} = \mathbf{3.645 \times 10^{-5}\ cm.}$$

Q20. *Calculate the wavelength of an electron moving with a velocity of 2.05×10^7 ms^{-1}.*

Ans. According to de Broglie's equation, $\lambda = \dfrac{h}{m\upsilon}$

Mass of electron (m) = 9.1×10^{-31} kg

Velocity of electron (υ) = 2.05×10^{7} ms^{-1}

Planck's constant (h) = 6.626×10^{-34} kgm^2 s^{-1}

$$\lambda = \frac{\left(6.626\times10^{-34}\text{ kg m}^2\text{ s}^{-1}\right)}{\left(9.1\times10^{-31}\text{ kg}\right)\times\left(2.05\times10^{7}\text{ ms}^{-1}\right)} = \mathbf{3.55 \times 10^{-4}\ m.}$$

Q21. *The mass of an electron is 9.1×10^{-31} kg. If its kinetic energy is 3.0×10^{-25} J, calculate its wavelength.*

Ans. **Step I.** *Calculation of the velocity of electron*

Kinetic energy = $1/2\, m\upsilon^2 = 3.0 \times 10^{-25}\text{ J} = 3.0 \times 10^{-25}\text{ kg m}^2\text{ s}^{-2}$

$$\upsilon^2 = \frac{2\times K.E.}{m} = \frac{2\times\left(3.0\times10^{-25}\text{ kg m}^2\text{ s}^{-2}\right)}{\left(9.1\times10^{-31}\text{ kg}\right)} = 65.9 \times 10^{4}\text{ m}^2\text{ s}^{-2}$$

$$\upsilon = (65.9 \times 10^{4}\text{ m}^2\text{ s}^{-2})^{1/2} = 8.12 \times 10^{2}\text{ ms}^{-1}$$

Step II. *Calculation of wavelength of the electron*
According to de Broglie's equation,

$$\lambda = \frac{h}{m\upsilon} = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}\right)}{\left(9.1 \times 10^{-31} \text{ kg}\right) \times \left(8.12 \times 10^2 \text{ ms}^{-1}\right)}$$

$= 0.08967 \times 10^{-5}$ m $= 8967 \times 10^{-10}$ m = **8967 Å.** $(\because 1 \text{ Å} = 10^{-10}\text{m})$

Q22. *Which of the following are iso-electronic species?*
Na^+, K^+, Mg^{2+}, Ca^{2+}, S^{2-}, *Ar.*

Ans. Na^+ and Mg^{2+} are iso-electronic species (have 10 electrons) K^+, Ca^{2+}, S^{2-} are iso-electronic species (have 18 electrons).

Q23. (*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 the elements whose outermost electrons are represented by:*
(a) $3s^1$ (b) $2p^3$ (c) $3p^5$?
(*iii*) *Which atoms are indicated by the following configurations?*
(a) [He] $2s^1$ (b) [Ne] $3s^2$ $3p^3$ (c) [Ar] $4s^2$ $3d^1$.

Ans. (*i*) (*a*) $1s^2$ (*b*) $1s^2$ $2s^2$ $2p^6$ (*c*) $1s^2$ $2s^2$ $2p^6$ (*d*) $1s^2$ $2s^2$ $2p^6$.
(*ii*) (*a*) Na (Z = 11) has outermost electronic configuration = $3s^1$
(*b*) *N* (Z = 7) has outermost electronic configuration = $2p^3$
(*c*) Cl (*Z* = 17) has outermost electronic configuration = $3p^5$
(*iii*) (*a*) Li (*b*) *P* (*c*) Sc.

Q24. *What is the lowest value of n which allows 'g' orbital to exist?*

Ans. The lowest value of *l* where '*g*' orbital can be present = 4
The lowest value of *n* where '*g*' orbital can be present = 4 + 1 = **5**.

Q25. *An electron is in one of the 3d orbitals. Give the possible values of n, l and m_l for this electron.*

Ans. For electron in 3*d* orbital, $n = 3$, $l = 2$, $m_l = -2, -1, 0, +1, +2$.

Q26. *An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons and (ii) the electronic configuration of the element.*

Ans. No. of protons in a neutral atom = No. of electrons = 29
Electronic configuration = $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^{10}$ $4s^1$.

Q27. *Give the number of electrons in the species : H_2^+, H_2 and O_2^+.*

Ans. H_2^+ = one; H_2 = two; O_2^+ = 15.

Q28. (*i*) *An atomic orbital has n = 3. What are the possible values of l and m_l?*
(*ii*) *List the quantum numbers (m_l and l) of electron for 3d orbital.*
(*iii*) *Which of the following orbitals are possible?*
1p, 2s, 2p and 3f.

Ans. (*i*) For $n = 3$; $l = 0$, 1 and 2.
For $l = 0$; $m_l = 0$
For $l = 1$; $m_l = +1, 0, -1$
For $l = 2$; $m_l = -2, -1, 0, +1, +2$
(*ii*) For an electron in 3rd orbital; $n = 3$; $l = 2$; m_l can have any of the values -2, $-1, 0, +1, +2$.
(*iii*) 1*p* and 3*f* orbitals are not possible.

Q29. *Using s, p and d notations, describe the orbitals with following quantum numbers :*
(a) $n = 1, l = 0$, (b) $n = 4, l = 3$, (c) $n = 3, l = 1$, (d) $n = 4, l = 2$

Ans. (*a*) 1*s* orbital (*b*) 4*f* orbital (*c*) 3*p* orbital (*d*) 4*d* orbital.

Q30. *From the following sets of quantum numbers, state which are possible. Explain why the others are not possible.*

(i) $n = 0, l = 0, m_l = 0, m_s = + 1/2$ *(ii)* $n = 1, l = 0, m_l = 0, m_s = - 1/2$

(iii) $n = 1, l = 1, m_l = 0, m_s = + 1/2$ *(iv)* $n = 1, l = 0, m_l = + 1, m_s = + 1/2$

(v) $n = 3, l = 3, m_l = - 3, m_s = + 1/2$ *(vi)* $n = 3, l = 1, m_l = 0, m_s = + 1/2$

Ans. (*i*) The set of quantum numbers is **not possible** because the minimum value of n can be 1 and not zero.

(*ii*) The set of quantum numbers is **possible.**

(*iii*) The set of quantum numbers is **not possible** because, for $n = 1$, l cannot be equal to 1. It can have 0 value.

(*iv*) The set of quantum numbers is **not possible** because for $l = 0$, m_l cannot be + 1. It must be zero.

(*v*) The set of quantum numbers is **not possible** because, for $n = 3$, $l \neq 3$.

(*vi*) The set of quantum numbers is **possible.**

Q31. *How many electrons in an atom may have the following quantum numbers?*

(a) $n = 4; m_s = - 1/2$ *(b)* $n = 3, l = 0.$

Ans. (*a*) **For $n = 4$**

Total number of electrons = $2n^2 = 2 \times 16 = 32$

Half out of these will have $m_s = - 1/2$

$\therefore$ Total electrons with m_s $(- 1/2) = 16$

(*b*) **For $n = 3$**

$l = 0$; $m_l = 0$, $m_s = + 1/2, - 1/2$ (two e^-)

Q32. *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.*

Ans. According to Bohr's theory,

$$m\upsilon r = \frac{nh}{2\pi} \qquad (n = 1, 2, 3, \text{... so on})$$

or

$$2\pi r = \frac{nh}{m\upsilon} \quad \text{or} \quad m\upsilon = \frac{nh}{2\pi r} \qquad ...(i)$$

According to de Broglie equation,

$$\lambda = \frac{h}{m\upsilon} \quad \text{or} \quad m\upsilon = \frac{h}{\lambda} \qquad ...(ii)$$

Comparing (*i*) and (*ii*),

$$\frac{nh}{2\pi r} = \frac{h}{\lambda} \quad \text{or} \quad 2\pi r = n\lambda$$

Thus, the circumference ($2\pi r$) of the Bohr orbit for hydrogen atom is an into the de Broglie wavelength.

Q33. *What transition in a hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum?*

Ans. For an atom, $\bar{\nu} = \dfrac{1}{\lambda} = R_H Z^2 \left(\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right)$

For He^+ spectrum : $Z = 2, n_2 = 4, n_1 = 2.$

$$\therefore \qquad \bar{\nu} = \frac{1}{\lambda} = R_H \times 4\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = \frac{3R_H}{4}$$

For hydrogen spectrum : $\bar{\nu} = \dfrac{3R_H}{4}$ and $Z = 1$

$$\therefore \qquad \bar{\nu} = \frac{1}{\lambda} = R_H \times 1\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\text{or} \qquad R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = \frac{3R_H}{4} \quad \text{or} \quad \frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{3}{4}$$

This corresponds to $n_1 = 1$, $n_2 = 2$ and means that the transition has taken Lyman series from $n = 2$ to $n = 1$.

Thus, the transition is from n_2 to n_1 in case of hydrogen spectrum.

Q34. *Calculate the energy required for the process :*

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

The ionisation energy for the H atom in the ground state is 2.18×10^{-18} J atom.

Ans. The expression for the ionisation energy of the atom:

$$E_n = \frac{2.18 \times 10^{-18} \times Z^2}{n^2} \text{ J atom}^{-1}$$

For H atom ($Z = 1$) $E_n = 2.18 \times 10^{-18} \times (1)^2$ J atom^{-1} (given)

For He^+ ion ($Z = 2$) $E_n = 2.18 \times 10^{-18} \times (2)^2$ = **8.72×10^{-18} J atom^{-1}**

(one electron species).

Q35. *If the diameter of carbon atom is 0.15 nm, calculate the number of carbon atoms which can be placed side by side in a straight line across a length of a scale of length 20 cm long.*

Ans. Length of scale = 20 cm = 20×10^7 nm = 2×10^8 nm

Diameter of carbon atom = 0.15 nm

$\therefore$ number of carbon atoms which can be placed side by side in the scale = $\dfrac{2 \times 10^8}{0.15}$

= **1.33×10^9**.

Q36. *2×10^8 atoms of carbon are arranged side by side. Calculate the radius of carbon if the length of this arrangement is 2.4 cm.*

Ans. The length of the arrangement = 2.4 cm

Total number of carbon atoms present = 2×10^8

Diameter of each carbon atom = $\dfrac{(2.4 \text{ cm})}{(2 \times 10^8)}$ = 1.2×10^{-8} cm

Radius of each carbon atom = $\dfrac{1}{2}$ (1.2×10^{-8}) = 6.0×10^{-9} cm = **0.06 nm.**

Q37. *The diameter of zinc atom is 2.6 Å. Calculate :*

(a) the radius of zinc atom in pm

(b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side length wise.

Ans. (*a*) Radius of zinc atom = $\dfrac{(2.6 \text{ Å})}{2}$ = 1.3 Å = 1.3×10^{-10} m = 130×10^{-12} m = **130 pm**

(*b*) Length of the scale = 1.6 cm = 1.6×10^{10} pm
Diameter of zinc atom = 260 pm

$$\therefore \text{ no. of zinc atoms placed side by side} = \frac{(1.6 \times 10^{10} \text{ pm})}{(260 \text{ pm})} = \mathbf{6.154 \times 10^{7}}.$$

Q38. *A certain particle carries 2.5×10^{-16} C of static electric charge. Calculate the number of electrons present in it.*

Ans. Magnitude of charge (q) = 2.5×10^{-16} C
Charge on one electron (e) = 1.602×10^{-19} C

$$\therefore \text{ No. of electrons present} = \frac{(2.5 \times 10^{-16} \text{ C})}{(1.602 \times 10^{-19} \text{ C})} = \mathbf{1560}$$

Q39. *In Millikan's experiment, the charge on the oil droplets was found to be -1.282×10^{-18} C. Calculate the number of electrons present in it.*

Ans. Charge on oil droplet = -1.282×10^{-18} C
Charge on an electron = -1.602×10^{-19} C

$$\text{Number of electrons} = \frac{q}{e} = \frac{(-1.282 \times 10^{-18} \text{ C})}{(-1.602 \times 10^{-19} \text{ C})} = \mathbf{8}$$

Q40. *In Rutherford experiment, generally the thin foil of heavy atoms like gold, platinum etc., have been used to be bombarded by the α-particles. If a thin foil of light atoms like aluminium etc., is used, what difference would be observed from the above results?*

Ans. Heavy atoms have a heavy nucleus and so a large amount of positive charge. Hence some α-particle deflected back on hitting the nucleus. If light atoms are used, their nuclei will be light and moreover, they will have small positive charge on the nucleus. As a result, the number of α-particles deflected will be quite less and the particles which are deflected back will be negligible.

Q41. *Symbols $^{79}_{35}Br$ and ^{79}Br can be written, whereas symbols $^{35}_{79}Br$ and ^{35}Br are not acceptable. Answer briefly.*

Ans. $^{35}_{79}Br$ is not acceptable because atomic number should be written as subscript while mass number should be written as superscript.
$_{35}Br$ is not acceptable because atomic number of an element is fixed. However, mass number is not fixed as it depends upon the isotope taken. Hence it is essential to indicate mass number.

Q42. *An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the symbol to the element.*

Ans. An element can be identified by its atomic number only. Let us find the atomic number.
Let the number of protons = x

$$\therefore \quad \text{Number of neutrons} = x + \frac{x \times 31.7}{100} = (x + 0.317\,x)$$

Now, Mass no. of element = No. of protons + No. of neutrons

$$81 = x + x + 0.317\,x = 2.317\,x \quad \text{or} \quad x = \frac{81}{2.317} = 35$$

$\therefore$ No. of protons = 35, No. of neutrons = 81 – 35 = 46

Atomic number of element (Z) = No. of protons = 35

The element with atomic number (Z) 35 is bromine $^{81}_{35}\mathbf{Br}$.

Q43. *An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons, find the symbol of the ion.*

Ans. Let the no. of electron in the ion = x

$\therefore$ the no. of protons = $x - 1$ (as the ion has one unit negative charge)

and the no. of neutrons = $x + \dfrac{x \times 11.1}{100} = 1.111x$

Mass no. or mass of the ion = No. of protons + No. of neutrons

$(x - 1 + 1.111\,x)$

Given mass of the ion = 37

$\therefore$ $x - 1 + 1.111\,x = 37$ or $2.111\,x = 37 + 1 = 38$

$$x = \frac{38}{2.111} = 18$$

No. of electrons = 18 ; No. of protons = 18 – 1 = 17

Atomic no. of the ion = 17 ; Atom corresponding to ion = Cl

Symbol of the ion = $^{37}_{17}\mathbf{Cl}^{-1}$

Q44. *An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign symbol to the ion.*

Ans. Let the no. of electrons in the ion = x

$\therefore$ the no. of the protons = $x + 3$ (as the ion has three units positive charge)

and the no. of neutrons = $x + \dfrac{30.4x}{100} = x + 0.304\,x$

Now, mass no. of ion = No. of protons + No. of neutrons

$= (x + 3) + (x + 0.304\,x)$

$\therefore$ $56 = (x + 3) + (x + 0.304\,x)$ or $2.304\,x = 56 - 3 = 53$

$$x = \frac{53}{2.304} = 23$$

Atomic no. of the ion (or element) = 23 + 3 = 26

The element with atomic number 26 is iron (Fe) and the corresponding ion is $\mathbf{Fe^{3+}}$.

Q45. *Arrange the following type of radiations in increasing order of wavelength : (a) radiation from microwave oven (b) amber light from traffic signal (c) radiation from FM radio (d) cosmic rays from outer space and (e) X-rays.*

Ans. Cosmic rays < X-rays < amber colour < microwave < FM.

Q46. *Nitrogen laser produces a radiation of wavelength of 337.1 nm. If the number of photons emitted is 5.6×10^{24}, calculate the power of this laser.*

Ans. Power of the laser (E) = $Nh\upsilon = Nh\,c/\lambda$

$$= \frac{\left(5.6 \times 10^{24}\right) \times \left(6.626 \times 10^{-34}\ \text{Js}\right) \times \left(3 \times 10^{8}\ \text{ms}^{-1}\right)}{\left(337.1 \times 10^{-9}\ \text{m}\right)} = \mathbf{3.3 \times 10^{6}\ J.}$$

Q47. *Neon gas is generally used in sign boards. If it emits strongly at 616 nm, calculate : (a) frequency of emission (b) the distance travelled by this radiation in 30s (c) energy of quantum (d) number of quanta present if it produces 2 J of energy.*

Ans. (*a*) Frequency of emission$(\nu) = \dfrac{c}{\lambda} = \dfrac{(3.0\times10^{8}\ \text{ms}^{-1})}{(616\times10^{-9}\ \text{m})} = \mathbf{4.87\times10^{14}\ s^{-1}}$.

(*b*) Velocity of radiation $(c) = 3.0\times10^{8}\ \text{ms}^{-1}$

Distance travelled in 30s $= (3.0\times10^{8}\ \text{ms}^{-1})\times(30\text{s}) = \mathbf{9.0\times10^{9}\ m}$.

(*c*) Energy of quanta $(E) = h\nu = \dfrac{hc}{\lambda} = \dfrac{(6.626\times10^{-34}\ \text{Js})\times(3\times10^{8}\ \text{ms}^{-1})}{(616\times10^{-9}\ \text{m})}$

$= \mathbf{32.27\times10^{-20}\ J}$.

(*d*) Number of quanta present in 2 J of energy

$$= \frac{\text{Total energy}}{\text{Energy per quanta}} = \frac{(2\ \text{J})}{(32.27\times10^{-20}\ \text{J})} = \mathbf{6.2\times10^{18}}.$$

Q48. *In astronomical observations, signals observed from the distant starts are generally weak. If the photon detector receives a total of 3.15×10^{-18} J from the radiations of 600 nm, calculate the number of photons received by the detector.*

Ans. Total energy received $= 3.15\times10^{-18}$ J

$\lambda = 600\ \text{nm} = 600\times10^{-9}\ \text{m} = 6\times10^{-7}\ \text{m}$

Energy of one photon, $E = \dfrac{hc}{\lambda} = \dfrac{(6.626\times10^{-34}\ \text{Js})\times(3\times10^{8}\ \text{ms}^{-1})}{(6\times10^{-7}\ \text{m})}$

$= 3.3125\times10^{-19}$ J

$\therefore$ No. of photons $= \dfrac{3.15\times10^{-18}\ \text{J}}{3.3125\times10^{-19}\ \text{J}} = \mathbf{10}$.

Q49. *Lifetimes of the molecules in the excited state are often measured by using pulsed radiation source of duration nearly in the nano second range. If the radiation source has the duration of 2ns and the number of photons emitted during the pulse source is 2.5×10^{15}, calculate the energy of the source.*

Ans. Frequency $= \dfrac{1}{2\times10^{-9}\ \text{s}}$

$= 0.5\times10^{9}\ \text{s}^{-1}$

Energy $= Nh\nu$

$= (2.5\times10^{5})\ (6.26\times10^{-39}\ \text{Js})\ (0.5\times10^{9}\ \text{s}^{-1})$

$= 8.28\times10^{-10}$ J

Q50. *The longest wavelength doublet absorption transition is observed at 589 nm and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.*

Ans. $\lambda_1 = 589\ \text{nm} = 589\times10^{-9}\ \text{m}$

$$\nu_1 = \frac{c}{\lambda_1} = \frac{3\times10^{8}\ \text{ms}^{-1}}{589\times10^{-9}\ \text{m}} = \frac{3000}{589}\times10^{14}\ \text{s}^{-1} = \mathbf{5.0934\times10^{14}\ s^{-1}}$$

$$\lambda_2 = 589.6 \text{ nm} = 589.6 \times 10^{-9} \text{ m}$$

$$\nu_2 = \frac{c}{\lambda_2} = \frac{3 \times 10^8 \text{ ms}^{-1}}{589.6 \times 10^{-9} \text{ m}} \; \frac{3000}{589.6} \times 10^{14} \text{ s}^{-1} = \mathbf{5.0882 \times 10^{14}\ s^{-1}}$$

$$\Delta E = E_1 - E_2 = \frac{hc}{\lambda_1} = \frac{hc}{\lambda_2} = hc\left[\frac{1}{\lambda_1} - \frac{1}{\lambda_2}\right]$$

$$= (6.626 \times 10^{-34} \text{ Js} \times 3 \times 10^8 \text{ ms}^{-1}) \left[\frac{1}{589 \times 10^{-9} \text{ m}} - \frac{1}{589.6 \times 10^{-9}}\right]$$

$$= \frac{19.878 \times 10^{-34} \times 10^8}{10^{-9}} \left[\frac{589.6 - 589}{589.6 \times 589}\right] \text{J}$$

$$= \frac{19.878 \times 10^{-17} \times 0.6}{589.6 \times 589} \text{J} = \mathbf{3.43 \times 10^{-22}\ J.}$$

Q51. *The work function for caesium atom is 1.9 eV. Calculate (a) the threshold wavelength and (b) the threshold frequency of the radiation. If the caesium element is irradiated a wavelength 500 nm, calculate the kinetic energy and the velocity of the photoelectron.*

Ans. $E_0 = 1.9 \text{ eV} = 1.9 \times 1.602 \times 10^{-19} \text{ J}$

$$\text{Threshold frequency } (\nu_0) = \frac{E_0}{h} = \frac{1.9 \times 1.602 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ Js}} = 459 \times 10^{15} \text{ s}^{-1}$$

$$= \mathbf{4.59 \times 10^{-19}\ J}$$

$$\text{Threshold wavelength } (\lambda_0) = \frac{c}{\nu_0} = \frac{3 \times 10^8 \text{ ms}^{-1}}{4.59 \times 10^{14} \text{ s}^{-1}} = 0.6536 \times 10^{-6} \text{ m}$$

$$= \mathbf{653.6\ nm \approx 654\ nm}$$

$$E = E_0 + \frac{1}{2}m\upsilon^2$$

$$\text{Kinetic energy } \left(\frac{1}{2}m\upsilon^2\right) = E - E_0 = hc\left[\frac{1}{\lambda} - \frac{1}{\lambda_0}\right]$$

$$= \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^8 \text{ ms}^{-1})}{10^{-9} \text{ m}} \times \left[\frac{1}{500} - \frac{1}{654}\right]$$

$$= \frac{6.626 \times 3 \times 154}{500 \times 654} \times 10^{-34+8+9} = \mathbf{9.36 \times 10^{-20}\ J}$$

$$\text{Velocity } (\upsilon) = \sqrt{\frac{2 \times 9.36 \times 10^{-20}}{m}} \text{J} = \sqrt{\frac{2 \times 9.36 \times 10^{-20} \text{ kg m}^2 \text{ s}^{-2}}{9.1 \times 10^{-31} \text{ kg}}}$$

$$= \sqrt{2.057 \times 10^{11} \text{ m}^2 \text{ s}^{-2}} = \sqrt{20.57 \times 10^{10} \text{ m}^2 \text{ s}^{-2}}$$

$$= \mathbf{4.5356 \times 10^5\ m/s.}$$

Q52. *Following results are observed when sodium metal is irradiated with different wavelengths. Calculate threshold wavelength.*

λ (nm) :	*500*	*450*	*400*
$\upsilon \times 10^{-5}$ (cm s^{-1}) :	*2.55*	*4.35*	*5.35*
λ (m) :	*500×10^{-9}*	*450×10^{9}*	*400×10^{-9}*
υ (ms^{-1}) :	*2.55×10^5*	*4.35×10^5*	*5.35×10^5*

Ans. Let threshold wavelength = λ_0 nm = $\lambda_0 \times 10^{-9}$ m

According to photoelectric effect :

$$h(\upsilon - \upsilon_0) = \frac{1}{2} m\upsilon^2$$

$$hc\left(\frac{1}{\lambda} - \frac{1}{\lambda_0}\right) = \frac{1}{2} m\upsilon^2$$

Substituting the results of three experiments in the above equation :

$$\frac{hc}{10^{-9}}\left(\frac{1}{500} - \frac{1}{\lambda_0}\right) = \frac{1}{2} m \ (2.55 \times 10^5)^2 \qquad ...(i)$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{450} - \frac{1}{\lambda_0}\right) = \frac{1}{2} m \ (4.35 \times 10^5)^2 \qquad ...(ii)$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{400} - \frac{1}{\lambda_0}\right) = \frac{1}{2} m \ (5.35 \times 10^5)^2 \qquad ...(iii)$$

Divide eqn. (*ii*) by eqn. (*i*)

$$\frac{(\lambda_0 - 450)}{450 \times \lambda_0} \times \frac{500 \times \lambda_0}{(\lambda_0 - 500)} = \frac{\left(4.35 \times 10^5\right)^2}{\left(2.99 \times 10^5\right)^2}$$

or $$\frac{(\lambda_0 - 450)}{(\lambda_0 - 500)} = \frac{(4.35)^2}{(2.99)^2} \times \frac{450}{500} = 2.619$$

or $$(\lambda_0 - 450) = 2.619\ (\lambda_0 - 500) = 2.619\ \lambda_0 - 1309.5$$

or $$1.619\ \lambda_0 = 859.5 \quad \therefore \quad \lambda_0 = \frac{859.5}{1.619} = \textbf{531 nm.}$$

Q53. *The ejection of the photoelectrons from the silver metal in the photoelectric effect experiment can be stopped by applying the voltage of 0.35 V when the radiation 256.7 nm is used. Calculate the work function for silver metal.*

Ans. $$\lambda = 256.7 \text{ nm} = 256.7 \times 10^{-9} \text{ m ; K.E.} = 0.35 \text{ eV}$$

$$E = \frac{hc}{\lambda} = \frac{\left(6.626 \times 10^{-34} \text{ Js}\right)\left(3 \times 10^8 \text{ ms}^{-1}\right)}{\left(256.7 \times 10^{-9} \text{ m}\right)}$$

$$= \frac{6.626 \times 3}{256.7} \times 10^{-17} \text{ J} = \frac{6.626 \times 3 \times 10^{-17}}{256.7 \times 1.602 \times 10^{-19}} \text{ eV}$$

$$= \frac{662.6 \times 3}{256.7 \times 1.602} \text{ eV} = 4.83 \text{ eV}$$

$$E = E_0 + \text{K.E.}$$

$$4.83 \text{ eV} = E_0 + 0.35 \text{ eV}$$

$$E_0 = 4.83 - 0.35 = \mathbf{4.48 \text{ eV.}}$$

Q54. *If the photon of the wavelength 150 pm strikes an atom, and one of its inner bound electrons is ejected out with a velocity of 1.5 × 10^7 ms^{-1}, calculate the energy with which it is bound to the nucleus.*

Ans. $\lambda = 150 \text{ pm} = 150 \times 10^{-12} \text{ m} = 1.5 \times 10^{-10} \text{ m}$; $\upsilon = 1.5 \times 10^7 \text{ ms}^{-1}$

$$K.E. = \frac{1}{2}mv^2 = \frac{1}{2} \times 9.1 \times 10^{-31} \text{ kg} \times (1.5 \times 10^7 \text{ ms}^{-1})^2$$

$$= \frac{9.1 \times 1.5 \times 1.5}{2} \times 10^{-31+14} \text{ J}$$

$$= 10.2375 \times 10^{-17} \text{ J} = 1.02375 \times 10^{-16} \text{ J}$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js})(3 \times 10^8 \text{ ms}^{-1})}{(1.5 \times 10^{-10} \text{ m})}$$

$$= \frac{6.626 \times 3}{1.5} \times 10^{-34+8+10}$$

$$= \frac{6.626 \times 3}{1.5} \times 10^{-16} \text{ J} = 13.386 \times 10^{-16} \text{ J}$$

$$E = E_0 + \text{K.E.}$$

$$E_0 = E - K.E. = (13.386 - 1.024) \times 10^{-16} \text{ J}$$

$$= 12.362 \times 10^{-16} \text{ J} = \frac{12 \cdot 362 \times 10^{-16}}{1 \cdot 602 \times 10^{-19}} = 7.7167 \times 10^3 \text{ eV.}$$

Q55. *Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n can be represented as v = 3.29 × 10^{15} (Hz) [1/3^2 – 1/n^2]*
Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

Ans.

$$v = (3.29 \times 10^{15} \text{ Hz}) \left(\frac{1}{3^2} - \frac{1}{n^2}\right)$$

$$\lambda = 1285 \text{ nm} = 1285 \text{ x } 10^{-19} \text{ m} = 1.285 \text{ x } 10^{-16} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{(3 \times 10^8 \text{ ms}^{-1})}{(1.285 \times 10^{-6} \text{ m})} = 2.3346 \times 10^{14} \text{ s}^{-1}$$

$$2.3346 \times 10^{14} = 3.29 \text{ x } 10^{15} \left[\frac{1}{3^2} - \frac{1}{n^2}\right]$$

$$\frac{2 \cdot 3346}{32 \cdot 9} = \frac{1}{3^2} - \frac{1}{n^2} \text{ or } 0.71 = \frac{1}{9} - \frac{1}{n^2}$$

$$\frac{1}{n^2} = \frac{1}{9} - 0.071 = 0.111 - 0.071 = 0.04$$

$$n^2 = \frac{1}{0.04} = 25 \text{ or } \boldsymbol{n = 5}$$

Paschen series lies in infrared region of the spectrum.

Q56. *Calculate the wavelength for the emission transition if it starts from the orbit have radius 1.3225 nm and ends at 211.6 pm. Name the series to which this transition being and the region of the spectrum.*

Ans. Radius of the orbit of H like species = $\frac{0.529}{Z} n^2$ Å = $\frac{52.9}{Z} n^2$ pm

$$r_1 = 1.3225 \text{ nm} = 1322.5 \text{ pm} = \frac{52 \cdot 9\, n_1^2}{Z} \quad ...(i)$$

$$r_2 = 211.6 \text{ pm} = \frac{52 \cdot 9\, n_2^2}{Z} \quad ...(ii)$$

$$\frac{r_1}{r_2} = \frac{1322.5}{211.6} = \frac{n_1^2}{n_2^2}$$

or $$\frac{n_1^2}{n_2^2} = 6.25 \text{ or } \frac{n_1}{n_2} = (6.25)^{1/2} = 2.5$$

Thus, if $n_1 = 2$, $n_2 = 5$. This transition corresponds to transition from 5th orbit to 2nd orbit. This means that the transition belongs to **Balmer series**.

Now, Wave number $(\overline{v})$ = $(1.097 \times 10^7 \text{ m}^{-1}) \times \left(\frac{1}{2^2} - \frac{1}{5^2}\right)$; $= 1.097 \times 10^7 \times \frac{21}{100} \text{ m}^{-1}$

$$= 23.037 \times 10^5 \text{ m}^{-1}$$

$$\text{Wave length } (\lambda) = \frac{1}{\overline{v}} = \frac{1}{(23 \cdot 037 \times 10^5 \text{ m}^{-1})} = 434 \times 10^{-9} \text{ m or } 434 \text{ nm}$$

It lies in the visible region of light.

Q57. *Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is 1.6×10^6 ms^{-1}, calculate de Brogile wavelength associated with this electron.*

Ans.

$$\lambda = \frac{h}{m.v} = \frac{(6.626 \times 10^{-34} \text{ Js})}{(9 \cdot 1 \times 10^{-31} \text{kg}) \times (1 \cdot 6 \times 10^6 \text{ms}^{-1})}$$

$= 0.455 \times 10^{-34 + 25}$ m = 0.455 nm = **455 pm.**

Q58. *Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.*

Ans.

$$\lambda = 800 \text{ pm} = 800 \times 10^{-12} \text{ m} = 8 \times 10^{-10} \text{ m}$$

$$m = 1.675 \times 10^{-27} \text{ kg}$$

$$v = \frac{h}{m\lambda} = \frac{(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1})}{(1.675 \times 10^{-27} \text{kg}) \times (8 \times 10^{-10} \text{m})}$$

$$= \frac{6.626}{1.675 \times 8} \times 10^{-34+37} \text{ ms}^{-1} = \frac{6.626 \times 10^3}{1.675 \times 8} \text{ ms}^{-1}$$

$$= \mathbf{0.494 \times 10^3 \text{ ms}^{-1} = 494 \text{ ms}^{-1}.}$$

Q59. *If the velocity of the electron in Bohr's first orbit is 2.19 × 10⁶ ms⁻¹, calculate the de Brogile wavelength associated with it.*

Ans.

$$v = 2.19 \times 10^6 \text{ ms}^{-1}$$

$$\lambda = \frac{h}{mv} = \frac{\left(6.626 \times 10^{-34} \text{kg m}^2 \text{ s}^{-1}\right)}{\left(9.1 \times 10^{-31} \text{kg}\right) \times \left(2.19 \times 10^6 \text{ms}^{-1}\right)}$$

$$= \frac{6.626}{9.1 \times 2.19} \times 10^{-34+25} \text{ m} = 0.33243 \times 10^{-9} \text{ m} = \mathbf{332.43 \text{ pm.}}$$

Q60. *The velocity associated with a proton moving in a potential difference of 1000 V is 4.37×10^5 ms^{-1}. If the hockey ball of mass 0.1 kg is moving with this velolcity, calculate the wavelength associated with this velocity.*

Ans.

$$v = 4.37 \times 10^5 \text{ ms}^{-1} \text{ ; m} = 0.1 \text{ kg}$$

$$\lambda = \frac{h}{mv} = \frac{\left(6.626 \times 10^{-34} \text{kg m}^2 \text{ s}^{-1}\right)}{\left(0.1\text{kg}\right) \times \left(4.37 \times 10^5 \text{ms}^{-1}\right)} = \frac{6.626}{0.437} \times 10^{-34-5} \text{m}$$

$$= 15.16 \times 10^{-39} \text{ m} = \mathbf{1.516 \times 10^{-38} \text{ m}}$$

Q61. *If the position of the electron is measured within an accuracy of ± 0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is h/4π × 0.05 nm. Is there any problem in defining this value?*

Ans.

$$\Delta x = 0.002 \text{ nm} = 0.002 \times 10^{-9} \text{ m} = 2.0 \times 10^{-12} \text{ m}$$

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

$$\Delta p = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}\right)}{4 \times 3.142 \times \left(2 \times 10^{-12} \text{ m}\right)} = 2.638 \times 10^{-23} \text{ kg ms}^{-1}$$

$$\text{Actual momentum } (p) = \frac{h}{4\pi \times 0.05 \text{ nm}} = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}\right)}{4 \times 3.142 \times \left(5 \times 10^{-11} \text{ m}\right)}$$

$$= 1.055 \times 10^{-24} \text{ kg mp}$$

Since actual momentum is smaller than the uncertainty in measuring momentum, therefore the momentum of electron cannot be defined.

Q62. *The quantum numbers of six electrons are given below. Arrange them in order of increasing energies. If any of these combination (s) has/have the same energy?*

(i) $n = 4, l = 2, m_l = -2, m_s = -1/2$ *(ii)* $n = 3, l = 2, m_l = 1, m_s = +1/2$

(iii) $n = 4, l = 1, m_l = 0, m_s = +1/2$ *(iv)* $n = 3, l = 2, m_l = -2, m_1 = -1/2$

(v) $n = 3, l = 1, m_l = -1, m_s = +1/2$ *(vi)* $n = 4, l = 1, m_l = 0, m_s = +1/2$

Ans. The electrons may be assigned to the following orbitals :

(*i*) 4*s* (*ii*) 3*d* (*iii*) 4*p* (*iv*) 3*d* (*v*) 3*p* (*vi*) 4*p*.

The increasing order of energy is :

(*v*) < (*ii*) = (*iv*) < (*vi*) = (*iii*) < (*i*)

Q63. *The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electrons in 4p orbital. Which of these electron experiences the lowest effective nuclear charge?*

Ans. 4*p* electron experiences the lowest effective nuclear charge because of the maximum magnitude of screening or shielding effect. It is farthest from the nucleus.

Q64. *Among the following pairs of orbitals, which orbital will experience more effective nuclear charge? (i) 2s and 3s (ii) 4d and 4f (iii) 3d and 3p?*

Ans. Please note that greater the penetration of the electron present in a particular orbital toward the nucleus, more will be the magnitude of the effective nuclear charge. Based upon this,

(*i*) 2*s* electron will experience more effective nuclear charge.

(*ii*) 4*d* electron will experience more effective nuclear charge.

(*iii*) 3*p* electron will experience more effective nuclear charge.

Q65. *The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus?*

Ans. Configuration of the two elements are:

$$\text{Al (Z = 13) : [Ne]}^{10}\ 3s^2 3p^1\ ;\ \text{Si (Z = 14) : [Ne]}^{10} 3s^2 3p^2$$

The unpaird electrons in silicon (Si) will experience more effective nuclear charge because the atomic number of the element Si is more than that of Al.

Q66. *Indicate the number of unpaired electrons in :*

(a) P (b) Si (c) Cr (d) Fe and (e) Kr.

Ans. (*a*) P (Z = 15) : $[\text{Ne}]^{10}3s^2 3p^3$; No. of unpaired electrons = 3

(*b*) Si (Z = 14) : $[\text{Ne}]^{10}3s^2 3p^2$; No. of unpaired electrons = 2

(*c*) Cr (Z = 24) : $[\text{Ar}]^{18}4s^1 3p^5$; No. of unpaired electrons = 6

(*d*) Fe (Z = 26) : $[\text{Ar}]^{18}4s^2 3d^6$; No. of unpaired electrons = 4

(*e*) Kr (Z = 36) : $[\text{Ar}]^{18}4s^2 3d^{10} 4p^6$; No. of unpaired electrons = Nil.

Q67. *(a) How many sub-shells are associated with n = 4?*

(b) How many electrons will be present in the sub-shells having m_s value of –1/2 for n = 4?

Ans. (*a*) For $n = 4$; No. of sub-shells = ($l = 0$, $l = 1$, $l = 2$, $l = 3$) = **4**.

(*b*) Total number of orbitals which can be present = $n^2 = 4^2 =$ **16**.

Each orbital can have an electron with $m_s = -1/2$

$\therefore$ Total no. of electrons with $m_s = \frac{1}{2}$ is **16.**

MORE QUESTIONS SOLVED

I. VERY SHORT ANSWER TYPE QUESTIONS

Q1. *Give the relation between wavelength and momentum of moving microscopic particle. What is the relation known as?*

Ans. Relation: $\lambda = \dfrac{h}{mV}$

The relation is known as de Broglie's relationship.

Q2. *Write the electronic configuration and number of unpaired electrons in Fe^{2+} ion.*

Ans. Fe (Z = 26) : $[Ar]^{18}\ 3d^6 4s^2$

Fe^{2+}ion : $[Ar]^{18}\ 3d^6$

No. of unpaired electrons = 4

Q3. *What are degenerate orbitals?*

Ans. Orbitals having same energy belonging to the same subshell.

Q4. *What is the most important application of de Broglie concept?*

Ans. In the construction of electron microscope used for the measurement of objects of very small size.

Q5. *Which one Fe^{3+}, Fe^{2+} is more paramagnetic and why?*

Ans. As Fe^{3+} contains 5 unpaired electrons while Fe^{2+} contains only 4 unpaired electrons. Fe^{3+} is more paramagnetic.

Q6. *Which element does not have any neutron?*

Ans. Hydrogen.

Q7. *What is value of Planck's constant in S.I. units?*

Ans. 6.62×10^{-34} Js.

Q8. *Arrange X-rays, cosmic rays and radio waves according to frequency.*

Ans. Cosmic rays > X-rays > radio waves.

Q9. *Which series of lines of the hydrogen spectrum lie in the visible region?*

Ans. Balmer series.

Q10. *What is the difference between ground state and excited state?*

Ans. Ground state means the lowest energy state. When the electrons absorb energy and jump to outer orbits, this state is called excited state.

Q11. *What is common between d_{xy} and $d_{x^2-y^2}$ orbitals?*

Ans. Both have identical shape, consisting of four lobes.

Q12. *If n is equal to 3, what are the values of quantum numbers l and m?*

Ans. $l = 0, 1, 2$

$m = -2, -1, 0, +1, +2$

and $S = +1/2$ and $-1/2$

for each value of m.

Q13. *Calculate the number of protons, neutrons and electrons in ${}^{80}_{35}Br$?*

Ans. Here, $Z = 35$

$A = 80$

$\therefore$ No. of protons = Atomic no. = 35

No. of electrons = No. of protons = 35

No. of neutrons = A – Z

= 80 – 35 = 45

Q14. *An electron beam after hitting a neutral crystal produces a diffraction pattern? What do you conclude?*

Ans. Electron has wave nature.

Q15. *An electron beam on hitting a ZnS screen produces a scientillation on it. What do you conclude?*

Ans. Electron has particle nature.

Q16. *Discuss the similarities and differences between a 1s and a 2s orbital.*

Ans. Similarities:

(*i*) Both have spherical shape.

(*ii*) Both have same angular momentum.

Differences:

(*i*) 1s has no node while 2s has one node.

(*ii*) Energy of 2*s* is greater than that of 1*s*.

Q17. *What will be the order of energy levels 3s, 3p and 3d in case of H-atom?*

Ans. All have equal energy.

Q18. *How many unpaired electrons are present in Pd (Z = 46)?*

Ans. The electronic configuration of the element palladium (Z = 46) is

$$[Kr]^{36}\ 4d^{10}\ 5s^{0}.$$

This means that it has no unpaired electron.

Q19. *Distinguish between a photon and quantum.*

Ans. A quantum is a bundle of energy of a definite magnitude (E = hv) and it may be from any source. However, a photon is quantum of energy associated with light only.

Q20. *What type of metals are used in photoelectric cells? Give one example.*

Ans. The metals with low ionisation enthalpies are used in photoelectric cells. Caesium (Cs), an alkali metal belonging to group 1 is the most commonly used metal.

Q21. *When is the energy of electron regarded as zero?*

Ans. The energy of the electron is regarded as zero when it is at infinite distance from the nucleus.

At that point force of attraction between the electron and the nucleus is almost nil. Therefore, its energy is regarded as zero.

Q22. *What is difference between the notations l and L?*

Ans. '*l*' signifies the secondary quantum number.

'*L*' signifies second energy level (n = 2).

II. SHORT ANSWER TYPE QUESTIONS

Q1. *The uncertainty in the position of a moving bullet of mass 10 g is 10^{-5} m. Calculate the uncertainty in its velocity?*

Ans. According to uncertainty principle,

$$\Delta x.\ m\Delta v = \frac{h}{4\pi} \text{ or } \Delta v = \frac{h}{4\pi m\Delta x};\ h = 6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1};\ m = 10 \text{ g} = 10^{-2} \text{ kg}$$

$$\Delta x = 10^{-5} \text{ m};\ \Delta v = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}\right)}{4 \times 3.143 \times \left(10^{-2} \text{ kg}\right) \times \left(10^{-5} \text{ m}\right)} = \mathbf{5.27 \times 10^{-28}\ mv}$$

Q2. *The uncertainty in the position and velocity of a particle are 10^{-10} m and 5.27×10^{-24} ms^{-1} respectively. Calculate the mass of the particle.* (Haryana Board 2000)

Ans. According to uncertainty principle,

$$\Delta x.\ m\Delta v = \frac{h}{4\pi} \text{ or } m = \frac{h}{4\pi\ \Delta x\ \Delta v};\ h = 6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}$$

$$\Delta x = 10^{-10} \text{ m};\ \Delta v = 5.27 \times 10^{-24} \text{ ms}^{-1}$$

$$m = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}\right)}{4 \times 3.143 \times \left(10^{-10} \text{m}\right) \times \left(5.27 \times 10^{-24} \text{ ms}^{-1}\right)} = \mathbf{0.1\ kg}$$

Q3. *With what velocity must an electron travel so that its momentum is equal to that of a photon of wavelength = 5200 Å?*

Ans. According to de Broglie equation, $\lambda = \dfrac{h}{mv}$

Momentum of electron, $mv = \dfrac{h}{\lambda} = \dfrac{(6.626 \times 10^{-34}\,\text{kg m}^2\,\text{s}^{-1})}{(5200 \times 10^{-10}\,\text{m})}$

$= 1.274 \times 10^{-27}$ kg ms^{-1} ...(*i*)

The momentum of electron can also be calculated as = $mv = (9.1 \times 10^{-31}\text{ kg}) \times v$ ...(*ii*)

Comparing (*i*) and (*ii*)

$(9.1 \times 10^{-31}\text{ kg}) \times v = (1.274 \times 10^{-27}\text{ kg ms}^{-1})$

$$v = \frac{(1.274 \times 10^{-27}\,\text{kg ms}^{-1})}{(9.1 \times 10^{-31}\,\text{kg})} = \mathbf{1.4 \times 10^3\ ms^{-1}}$$

Q4. *Using Aufbau principle, write the ground state electronic configuration of following atoms. (i) Boron (Z = 5) (ii) Neon (Z = 10), (iii) Aluminium (Z = 13) (iv) Chlorine (Z = 17), (v) Calcium (Z = 20) (vi) Rubidium (Z = 37)*

Ans. (*i*) Boron (Z = 5) ; $1s^2\,2s^2\,2p^1$

(*ii*) Neon (Z = 10) ; $1s^2\,2s^2\,2p^6$

(*iii*) Aluminium (Z = 13) ; $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$

(*iv*) Chlorine(Z = 17) ; $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$

(*v*) Calcium (Z = 20) ; $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$

(*vi*) Rubidium (Z = 37) ; $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}4s^2\,4p^6\,5s^1$.

Q5. *Calculate the de Broglie wavelength of an electron moving with 1% of the speed of light?*

Ans. According to de Broglie equation, $\lambda = \dfrac{h}{mv}$

Mass of electron = 9.1×10^{-31} kg ; Planck's constant = 6.626×10^{-34} kgm^{2}s^{-1}

Velocity of electron = 1% of speed of light = $3.0 \times 10^8 \times 0.01 = 3 \times 10^6$ ms^{-1}

Wavelength of electron (λ) = $\dfrac{h}{mv} = \dfrac{(6.626 \times 10^{-34}\ \text{kg m}^2\text{s}^{-1})}{(9.1 \times 10^{-31}\,\text{kg}) \times (3 \times 10^6\,\text{ms}^{-1})}$

$= 2.43 \times 10^{-10}$ m.

Q6. *The kinetic energy of an electron is 4.55×10^{-25} J. The mass of electron 9.1×10^{-1} kg. Calculate velocity, momentum and the wavelength of the electron?*

(Haryana Board, 2004, AI CBSE 2000)

Ans. **Step I.** Calculation of the velocity of electron

Kinetic energy = $1/2\ mv^2 = 4.55 \times 10^{-25}$ J = 4.55×10^{-25} kg m^2 s^{-2}

or $$v^2 = \frac{2 \times \text{KE}}{m} = \frac{2 \times (4.55 \times 10^{-25}\ \text{kg m}^2\text{s}^{-2})}{(3 \times 10^6\, ms^{-1})} = 10^6\ \text{m}^2\ \text{s}^{-2}$$

or Velocity (v) = $(10^6\ \text{m}^2\ \text{s}^{-2})^{1/2}$ = $\mathbf{10^3\ ms^{-1}}$

Step II. Calculation of the momentum of the electron
Momentum of electron = mv = $(9.1 \times 10^{-31}$ kg$) \times (10^3$ m s$^{-1}) = 9.1 \times 10^{-28}$ kg m^{-1}
Step III. Calculation of the wavelength of the electron
According to de Broglie equation:

$$\lambda = \frac{h}{mv} = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}\right)}{\left(9.1 \times 10^{-31}\text{kg}\right) \times \left(10^3 \text{ m s}^{-1}\right)}$$

$$= 0.728 \times 10^{-6} \text{ m} = \mathbf{7.28 \times 10^{-7} \text{ m}}$$

Q7. *What is the wavelength for the electron accelerated by 1.0×10^4 volts?*

Ans. **Step I.** Calculation of the velocity of electron
Energy (kinetic energy) of electron = 1.0×10^4 volts.

$$= 1.0 \times 10^4 \times 1.6 \times 10^{-19} \text{ J} = 1.6 \times 10^{-15} \text{ J}$$

$$= 1.6 \times 10^{-15} \text{ kg m}^2 \text{ s}^{-2}$$

or $$1/2\, mv^2 = 1.6 \times 10^{-15} \text{ kg m}^2 \text{ s}^{-2}$$

or $$v = \left(\frac{2 \times 1.6 \times 10^{-15} \text{ kg m}^2\text{s}^{-2}}{9.1 \times 10^{-31}\text{kg}}\right)^{1/2} = 5.93 \times 10^7 \text{ ms}^{-1}$$

Step II. Calculation of the wavelength of electron
According to de Broglie equation,

$$\lambda = \frac{h}{mv};\ \lambda = \frac{\left(6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}\right)}{\left(9.1 \times 10^{-31} \text{ kg}\right) \times \left(5.93 \times 10^7 \text{ m } s^{-1}\right)} = \mathbf{1.22 \times 10^{-11} \text{ m.}}$$

Q8. *In a hydrogen atom, the energy of an electron in first Bohr's orbit is 13.12×10^5 J mol^{-1}. What is the energy required for its excitation to Bohr's second orbit?*

Ans. The expression for the energy of electron of hydrogen is:

$$E_n = -\frac{2\pi^2 m_e{}^4}{n^2 h^2}$$

When $$n = 1,\ E_1 = -\frac{2\pi^2 m_e{}^4}{(1)^2 h^2} = -13.12 \times 10^5 \text{ J mol}^{-1}$$

When $$n = 2,\ E_2 = -\frac{2\pi^2 m_e{}^4}{(2)^2 h^2} = -\frac{13 \cdot 12 \times 10^5}{4} \text{ J mol}^{-1}$$

$$= -3.28 \times 10^5 \text{ J mol}^{-1}.$$

The energy required for the excitation is :

$$\Delta E = E_2 - E_1 = (-3.28 \times 10^5) - (-13.12 \times 10^5) = \mathbf{9.84 \times 10^5 \text{ J mol}^{-1}.}$$

Q9. *What are the two longest wavelength lines (in manometers) in the Lyman series of hydrogen spectrum?*

Ans. According to Rydberg-Balmer equation.

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

The wavelength (λ) will be the longest when n_2 is the smallest *i.e.,* $n_2 = 2$ and 3 for two longest wavelength lines.

For $n_2 = 2$:
$$\frac{1}{\lambda} = (1.097 \times 10^{-2}\ \text{nm}^{-1}) \left[\frac{1}{1^2} - \frac{1}{2^2}\right]$$
$$= (1.097 \times 10^{-2}\ \text{nm}^{-1}) \times \frac{3}{4} = 8.228 \times 10^{-3}\ \text{nm}^{-1} \text{ or } \lambda = \mathbf{121.54\ nm}$$

For $n_2 = 3$:
$$\frac{1}{\lambda} = (1.097 \times 10^{-2}\ \text{nm}^{-1}) \left[\frac{1}{1^2} - \frac{1}{3^2}\right]$$
$$= (1.097 \times 10^{-2}\ \text{nm}^{-1}) \times (8/9) = 9.75 \times 10^{-3}\ \text{nm}^{-1}\ ;\ \lambda = \mathbf{102.56\ nm}$$

III. LONG ANSWER TYPE QUESTIONS

Q1. *(a) What is the limitations of Rutherford model of atoms?*

(b) How has Bohr's theory helped in calculating the energy of hydrogen electron in different energy levels?

Ans. (*a*) **Limitations of Rutherford Model:**

(*i*) When a body is moving in an orbit, it achieves acceleration (even if body is moving with constant speed in an orbit, it achieves acceleration due to change in direction). So an electron moving around nucleus in an orbit is under acceleration. However, according to radiation theory of Maxwell, the charged particles when accelerated must emit energy as electromagnetic radiations. This means that the revolving electron must also lose energy continuously in the form of electromagnetic radiation. The loss of energy in revolution of the electron around the nucleus must bring it closer to the nucleus and the electron must ultimately fall into the nucleus by the spiral path. This means that the atom must collapse. But we all know that atom is quite stable in nature.

(*ii*) Rutherford's model could not explain the existence of different spectral lines in the hydrogen spectrum.

(*b*) Based upon the postulates of Bohr's theory, it is possible to calculate the energy of the hydrogen electron and also one electron species. (He^+, Li^{2+} etc.) The mathematical expression for the energy in the n^{th} orbit is

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

By substituting the values of m_e (mass of electron), e (charge of electron) and h (Planck's constant), the value of energy comes out to be

$$E_n = -\frac{2\cdot 178 \times 10^{-18} \times Z^2}{n^2} \text{ J per atom.}$$
$$= -\frac{1312 \times Z^2}{n^2} \text{ KJ mol}^{-1}$$

For hydrogen electron,

$$Z = 1$$
$$E_n = -\frac{1312}{n^2} \text{ KJ mol}^{-1}$$

The value for $n = 1$, gives the energy of the hydrogen electron in the ground state.

By assigning values, energy in different excited states can be calculated.

Q2. *Define atomic number, mass number and neutron. How are the three related to each other?*

Ans. **Atomic Number (Z):** The atomic number of an element is equal to the number of protons present inside the nucleus of its atoms.

Since, an isolated atom has no net charge on it, in neutral atoms, the total number of electrons is equal to its atomic number.

$$\text{Atomic number } (Z) = \text{Number of protons in the nucleus of an atom}$$
$$= \text{Number of electrons in the neutral atoms}$$

Mass Number (A): The sum of the number of neutrons and protons in the nucleus of an atom is called its mass number. Mass number is denoted by A. Thus, for an atom,

$$\text{Mass number } (A) = \text{Number of protons } (p) + \text{Number of neutrons } (n)$$
$$A = p + n$$

Neutron: It is neutral particle. It is present in the nucleus of an atom. Expect hydrogen (which contains only one electron and one proton but no neutron), the atoms of all other elements including isotopes of hydrogen contain all the three fundamental particles called neutron, proton and electron.

The relation between mass number, Atomic no. and no. of neutrons is given by the equation:

$$\boxed{A = Z + n}$$

Where A = Mass number

Z = Atomic number

n = Number of neutrons in the nucleus.

Q3. *What were the weaknesses or limitations of Bohr's model of atoms? Briefly describe the quantum mechanical model of atom.*

Ans. **Limitations of Bohr's model of an atom:**

(*i*) It could not explain spectrum of multi-electron atoms.

(*ii*) It could not explain Zeeman and Stark effects.

(*iii*) It could not explain shape of molecules.

(*iv*) It was not in accordance with Heisenberg's uncertainty principle.

Quantum Mechanical Model: It was developed on the basis of Heisenberg's uncertainty principle and dual behaviour of matter.

Main features of this model are given below :

(*i*) The energy of electrons in an atom is quantized *i.e.* can only have certain values.

(*ii*) The existence of quantized electronic energy levels is a direct result of the wave like properties of electrons.

(*iii*) Both, the exact position and velocity of an electron in an atom cannot be determined simultaneously.

(*iv*) The orbitals are filled in increasing order of energy. All the information about the electron in an atom is stored in orbital wave function ψ.

(*v*) From the value of ψ^2 at different points within atom, it is possible to predict the region around the nucleus where electron most probably will be found.

Q4. *State and explain the following:*

(i) Aufbau principle

(ii) Pauli exclusion principle.

(iii) Hund's rule of maximum multiplicity.

Ans. (*i*) **Aufbau Principle:** In the ground state of the atoms, the orbitals are filled in the order of their increasing energies. In other words, electrons first occupy the lowest-energy orbital available to them and enter into higher energy orbitals only after the lower energy orbitals are filled.

The order in which the energies of the orbitals increase and hence the order in which the orbitals are filled is as follows:

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, 5*f*, 6*d*, 7*p*.........

(*ii*) **Pauli Exclusion Principle:** An orbital can have maximum of two electrons and these must have opposite signs.

For example: Two electrons in an orbital can be represented by

$$\underset{(i)}{\boxed{\uparrow\downarrow}} \quad \text{or} \quad \underset{(ii)}{\boxed{\downarrow\uparrow}}$$

The two electrons have opposite spin, if one is revolving clockwise, the other is revolving anticlockwise or vice versa.

(*iii*) **Hund's Rule of Maximum Multiplicity:** Electron pairing in *p*, *d* and *f* orbitals cannot occur until each orbital of a given subshell contains one electron each or is single occupied.

For example: For the element nitrogen which contains 7 electrons, the following configuration can be written.

1s	2s	$2p_x$	$2p_y$	$2p_z$
$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow$

Total spin of unpaired electrons

$$= \frac{1}{2}+\frac{1}{2}+\frac{1}{2} = 1\frac{1}{2}.$$

IV. MULTIPLE CHOICE QUESTIONS

1. Cathode rays are deflected by

(*a*) electric field only (*b*) electric and magnetic field

(*c*) magnetic field only (*d*) none of these

2. In a sodium atom (atomic number = 11 and mass number = 23) and the number of neutrons is

(*a*) equal to the number of protons

(*b*) less than the number of protons

(*c*) greater than the number of protons

(*d*) none of these

3. The Balmer series in the spectrum of hydrogen atom falls in

(*a*) ultraviolet region (*b*) visible region

(*c*) infrared region (*d*) none of these

4. The idea of stationary orbits was first given by

(*a*) Rutherford (*b*) J.J. Thomson (*c*) Niels Bohr (*d*) Max Planck

5. de Broglie equation is

(*a*) $\lambda = \dfrac{h}{mv}$ (*b*) $\lambda = \dfrac{hv}{m}$ (*c*) $\lambda = \dfrac{mv}{h}$ (*d*) $\lambda = hmv$

6. The orbital with n = 3 and l = 2 is

(*a*) 3*s* (*b*) 3*p* (*c*) 3*d* (*d*) 3*j*

7. The outermost electronic configuration of manganese (at. no. = 25) is

(*a*) $3d^5\ 4s^2$ (*b*) $3d^6\ 4s^1$ (*c*) $3d^7\ 4s^0$ (*d*) $3d^6\ 4s^2$

8. The energy needed to remove a single electron (most loosely bound) from an isolated gaseous atom is called

(*a*) ionisation energy (*b*) electronegativity

(*c*) kinetic energy (*d*) electron affinity

9. The maximum number of electrons in a sub-shell is given by the equation

(*a*) n^2 (*b*) $2n^2$ (*c*) $2l - 1$ (*d*) $2l + 1$

10. If the value of azimuthal quantum number is 2, what will be the values for magnetic quantum number?

(*a*) 2 (*b*) 3 (*c*) 4 (*d*) 5

Ans. 1. (*b*) 2. (*c*) 3. (*b*) 4. (*c*) 5. (*a*)

6. (*c*) 7. (*a*) 8. (*a*) 9. (*d*) 10. (*d*)

V. HOTS QUESTIONS

Q1. *Give the name and atomic number of the inert gas atom in which the total number of d-electrons is equal to the difference between the numbers of total p and total s-electrons.*

Ans. Electronic configuration of Kr (atomic no. = 36) = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

Total no. of s-electrons = 8

Total no. of p-electrons = 18

Difference = 10, no. of d-electrons = 10

Q2. *What is the minimum product of uncertainty in position and momentum of an electron?*

Ans. $h/4\pi$

Q3. *Which orbital is non-directional?*

Ans. s-orbital.

Q4. *What is the difference between the notations l and L?*

Ans. *l* represents the subshell and *L* represents shell.

Q5. *How many electrons in an atom can have $n + l = 6$?*

Ans. 18.

Q6. *An anion A^{3+} has 18 electrons. Write the atomic number of A.*

Ans. 15.

Q7. *Arrange the electron (e), protons (p) and alpha particle (α) in the increasing order for the values of e/m (charge/mass).*

Ans. $\alpha < p < e$.

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