# XI CHEMISTRY ASSIGNMENT ON STRUCTURE OF ATOM

## DISCOVERY OF SUB-ATOMIC PARTICLES AND ATOMIC MODELS

- **John Dalton** (1767-1844), an english scientist, proposed his famous "Atomic theory" about the structure of matter. According to his concept,
  - ➢ Matter is composed of extremely small indivisible particles called atoms.
  - ➢ Atoms are structureless hard spherical particles.
  - ➢ Atoms of an element are similar in shape, size and properties but differ from atoms of other elements.
  - ➢ Atom is the smallest particle of matter that takes part in a chemical reaction.
  - ➢ Atoms combine in a definite proportion to form molecules.
- *J.J. Thomson, Rutherford, Goldstein, Chadwick* and *Bohr* have proved that atom is no longer a smallest indivisible particle but have a complex structure.
- The discovery of electron came form the experiments of **William crookes** (1879). He performed discharge tube experiments which led to discovery of cathode rays.
- Cathode rays have been found to consist of negatively charged particles, called electrons.
- **Properties of cathode rays:** The cathode rays have been Shown to possess the following properties by the experiments of **J.J. Thomson** and other scientists.
  - ➢ Cathode rays travel in a straight line.
  - ➢ Cathode rays possess particle nature.
  - ➢ Cathode rays contains negatively charged. Particles and are deflected towards the positive plate (cathode) on application of electric field.
  - ➢ Cathode rays are deflected by magnetic field.
  - ➢ Cathode rays produce heating effect.
  - ➢ Cathode rays produce X-rays when strikes on the surface of solids such as copper, etc.
  - ➢ Cathode rays produce fluorescence.
  - ➢ Cathode rays ionise the gas through which they pass.
  - ➢ Cathode rays affect the photographic plates.
  - ➢ Cathode rays penetrate through the thin foils of metals.
  - ➢ The nature of these rays does not depend upon the nature of gas and the cathode material used in discharge tube.
- J.J. Thomson found the ratio of charge to mass (e/m) constant by using different discharge tubes fitted with electrodes of different metals. He also used different gases in the discharge tube: The value of e/m was found to be **$1.76 \times 10^8$ coulombs/g.** Two important conclusions can be drawn from this experiment.

  (i) The electrons are universal constituents of all matter.

  (ii) For an electron e/m value is $1.76 \times 10^8$ coulomb/g or $1.76 \times 10^{11} Ckg^{-1}$
- The first accurate determination of the charge of an electron (e) was made by R.A. Millikan (1909) by his famous oil **drop experiment.**
- The charge on the electron is determined to be **$1.6 \times 10^{-19}$** coulombs. Since this is the smallest charge carried by any charged particle, it is referred to as **one unit charge.**
- By knowing the value of $\frac{e}{m}$ ratio and charge on the electron, it is possible to calculate the exact mass of electron.

$\therefore$ Mass of the electron, $m = \frac{e}{e/m} = \frac{1.60 \times 10^{-19} \text{ coulombs}}{1.76 \times 10^{8} \text{ coulombs/g}}$ $= 9.1 \times 10^{-28}$ g $= 9.1 \times 10^{-31}$ kg

- Mass of an electron $= \frac{1}{1837}$ × mass of one atom of hydrogen
- The first proof for the presence of ***proton*** in an atom was provided by the experiment of ***Goldstein.***
- **Properties of anode rays:**

➢ The anode rays travel in straight line.

➢ Anode rays are made up of material particles which are positively charged.

➢ Anode rays are deflected in the electrical as well as in the magnetic field.

➢ The ratio of charge to mass (i.e. e/m) and the value of the charge of the particles in the anode rays depends upon the nature of the gas taken in the discharge tube.

➢ Like cathode rays they also cause heating effect and affect the photographic plate.

- e/m ratio for a proton is determined to be $9.58 \times 10^4$ coulombs/g. The charge on proton is equal to electron in magnitude i.e. $1.6 \times 10^{-19}$ coulombs but opposite in sign. Thus, the mass of each proton is calculated as

Mass of proton $= \frac{e}{e/m} = \frac{1.6 \times 10^{-19}}{9.58 \times 10^{4}} = 1.67 \times 10^{-24}$ g $= 1.67 \times 10^{-27}$ kg

The value is nearly the same as the of hydrogen atom. Hence proton is 1837 times heavier than electron.

- A proton may be defined as a subatomic fundamental particle which carries one unit positive charge ($1.6 \times 10^{-19}$ coulombs) and has mass ($1.67 \times 10^{-27}$ kg) all most equal to that of an atom of hydrogen.
- The phenomenon of radioactivity discovered by **Henri Becqueral** (1896) further supported that atoms are divisible into subatomic particles.
- Radioactivity is a phenomenon of spontaneous emission of active radiations by a radioactive element.
- Rutherford (1902) resolved these radiations into three types by placing uranium mineral in a lead box and passing emitted radiations between two oppositely charged plates.

➢ The radiations bending towards negative plate (cathode) show that these radiations are carrying positive charge and thus named as alpha rays or alpha particles (alpha; $_2He^4$).

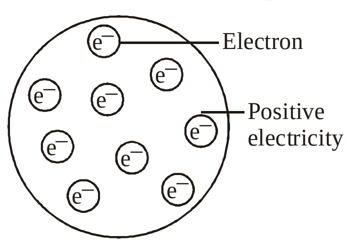
➢ The beam of rays which are deflected towards positive electrode (anode) show that these radiations are carrying **negative charge** and are known as **beta rays or beta particles (β).** These rays are deflected to much greater extent in comparison to deflection of α-rays showing that β-rays are lighter than α-rays.

➢ The third type of radiation, which are not deflected even in the strongest electric or magnetic field, are neutral in nature and are termed as **gamma rays.**

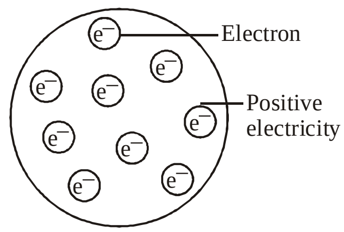
Order of penetrating power

$$\alpha\text{-particles (rays)} \quad < \quad \underset{\text{(100 times that of } \alpha\text{-particles)}}{\beta\text{ - rays}} \quad < \quad \underset{\text{(1000 times that of } \alpha\text{-particles)}}{\gamma\text{ - rays}}$$

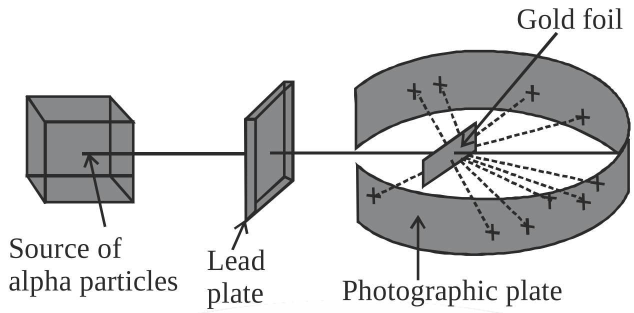
- According to **J.J. Thomson** the positive charge is spread over a sphere of the size of the atom (i.e. radius of sphere is of the order $10^{-8}$ cm). A sufficient number of electron so as to neutralise the positive charge are embedded over the sphere. This model of the atom is known as **"Plumpudding model".**




- **Rutherford** (1911) performed scattering experiment by bombarding fast moving α-particles, emitted from a radioactive substance, on thin foil ($4 \times 10^{-5}$ cm thick) of the metals like silver, gold, copper, platinum, etc).




According to Rutherford's model of atom, the atom consists of two parts.




Rutherford's scattering experiment

- **Nucleus:** It is a very small positively charged part of the atom. It is situated at the centre and carries almost the entire mass of the atom. The positive charge on the nucleus is due to the presence of protons in the nucleus.
- **Extranuclear part:** This is the space around the nucleus in which the electrons revolve at a high speed in a fixed path called orbits or shells. This extra nuclear part provides volume to an atom and thus called **atomic volume.**

**Rutherford's** model of atom is similar to our solar where the nucleus is like the sun and the electrons, are like the planets thus this model of atom is also called planetary model of atom.

- **Chadwick** (1932), while studying the bombardment of light elements, such as beryllium, boron and lithium, by fast moving $\alpha$-particles obtained highly penetrating radiations. These radiations were found to have high ionising power and were not deflected by electric or magnetic field. These neutral particles were found to have mass $1.675\times10^{-24}$ g and were named neutrons.
  The reactions responsible for the production of neutron were later found to be

  (i) $\underset{(\alpha-\text{particle})}{{}^{4}_{2}\text{He}} + {}^{9}_{4}\text{Be} \longrightarrow {}^{12}_{6}\text{C} + \underset{(\text{Neutron})}{{}^{1}_{0}\text{n}}$ (ii) ${}^{4}_{2}\text{He} + {}^{11}_{5}\text{Be} \longrightarrow {}^{14}_{7}\text{N} + {}^{1}_{0}\text{n}$
- A neutron is defined as a sub-atomic particle which has mass equal to that of a proton ($1.675 \times 10^{-27}$ kg) but has no charge.
- The nucleus consists of protons and neutrons and these are collectively termed as nucleons. The entire mass of the atom is due to the number of nucleons present in the atom.
- **Characteristics of fundamental particles**

| | | **Electron** | **Proton** | **Neutron** |
|---|---|---|---|---|
| 1 | Symbol | ${}_{-1}e^{0}$ | ${}_{+1}p^{1}$ | ${}_{0}n^{1}$ |
| 2 | Discoverer | J.J. Thomson (1897) | Goldstein (1919) | James Chadwick (1932) |
| 3 | Location in the atom | Extra nuclear part | Nucleus | Nucleus |
| 4 | Nature | Negatively charged | Positively charged | Neutral (uncharged) |
| 5 | Charge | | | |
| | (i) Relative | –1 | 1 | 0 |
| | (ii) Absolute | $-1.6022 \times 10^{-19}$C | $1.6022\times10^{-19}$C | 0 |
| 6 | Mass | | | |
| | (i) amu | 0.0005486 | 1.00727 | 1.00867 |
| | (ii) kg | $9.10939 \times 10^{-31}$ kg | $1.67262 \times 10^{-27}$ kg | $1.67493 \times 10^{-27}$ kg |
| | (iii) Relative to hydrogen atom | 1/1837 | 1 | 1 |

- The number of unit positive charges possessed by the nucleus of an atom is termed as the atomic number.

  **Atom number (Z) = Number of protons = Number of electrons**
- The total mass of the atom is due to the sum of the protons and neutrons i.e. nucleons, thus,

  **Mass number (A) = Number of proton + Number of neutrons**
- ***Yukawa*** (1935) suggested that **meson,** a sub atomic particle, acts as a nuclear glue to bind the nucleons (neutrons and protons) together in the nucleus. He suggested that mesons originate due to collisions between nuclear particles and play an important role in providing stability to the nucleus. The neutrons and protons are constantly undergoing exchange inside the nucleus through the involvement of meson.
- **Isotopes:** Atoms of the same element having same atomic number but different mass number are called isotopes. For example, ${}_1^1H, {}_1^2, {}_1^3H$ have same atomic number i.e. 1, but different mass number 1,2 and 3. These are named as protium (H), deuterium (D) and tritium (T) respectively. Other examples of isotopes are

  (i) ${}_6^{12}C, {}_6^{13}, {}_6^{14}C$ (ii) ${}_{17}Cl^{35}, {}_{17}Cl^{37}$ (iii) ${}_7^{14}N, {}_7^{15}N$ (iv) ${}_8^{16}O, {}_8^{17}O, {}_8^{18}O$

  As a result, all atoms of an element must have the same number of protons. But they don't have to contain the same number of neutrons.
- **Isobars:** Isobars are atoms of different elements which have same mass number but different atomic number. For example, $\frac{40}{18}Ar, {}_{19}^{40}K, {}_{20}^{40}Ca$

  Other examples of isobars are (i) ${}_{54}^{130}Xe, {}_{56}^{130}Ba$ (ii) ${}_7^{14}N, {}_6^{14}C$
- **Isotones:** Isotones are atoms of different elements which have different atomic number and mass number but have the same number of neutrons. For example, ${}_7^{15}N, {}_8^{16}O$.

  Other examples of isotones are (i) ${}_6^{14}C, {}_7^{15}N$ (ii) ${}_{15}^{31}P, {}_{16}^{32}S$

## ILLUSTRATIONS

**1.** The atomic number and mass number of calcium are 20 and 40 respectively. Find out the number of electrons, protons and neutrons in a $Ca^{2+}$ ion.

**Sol.** $Ca^{2+}$ ion is formed when the Ca-atom loses two electrons.

Atomic number of Ca = number of protons = number of electrons = 20

Mass number of Ca = number of protons + number of neutrons; So, Number of neutrons = 40 – 20 = 20

$Ca^{2+}$ possesses the same number of protons and neutrons as Ca-atom.

Hence, Number of protons in $Ca^{2+}$ = 20; Number of neutrons in $Ca^{2+}$ = 20

Number of electrons in $Ca^{2+}$ = 20 – 2 = 18

**2.** Which of the following are isotopes of sulphur? (At. no. of sulphur = 16)

[A] 16p + 16n; [B] 16n + 17p; [C] 19p + 16n; [D] 16p + 17n.

**Sol.** Isotopes are those atoms which possess the same atomic number, i.e., same number of protons. Thus, [A] = 16p + 16n and [D] = 16p + 17n are isotopes of sulphur.

**3.** How many protons, electrons and neutrons are present in 0.18 g ${}_{15}^{30}P$?

**Sol.** No. of products in one atom = No. of electrons in one atom = 15. No. of neutrons in one atom = (30 – 15) = 15

$0.18\text{ g } {}_{15}^{30}P = \frac{0.18}{30} = 0.006$ mole; No. of ${}_{15}^{30}P$ atoms in 0.006 mole = $0.006 \times 6.02 \times 10^{23}$

No. of protons in 0.006 mole ${}_{15}^{30}P = 15 \times 0.006 \times 6.02 \times 10^{23} = 5.418 \times 10^{22}$

So, no. of electrons = $5.418 \times 10^{22}$ and no. of neutrons = $5.418 \times 10^{22}$

**4.** Which of the following elements are isotopes and which are isobars?

$^{40}_{18}Ar$, $^{35}_{16}Cl$, $^{40}_{20}Ca$, $^{37}_{17}Cl$, $^{40}_{19}K$

**Sol.** $^{35}_{17}Cl$ and $^{37}_{17}Cl$ are isotopes because both have same atomic number but different atomic masses. $^{48}_{18}Ar$, $^{40}_{20}Ca$ and $^{40}_{19}K$ are isobars as these have different atomic numbers but same atomic mass.

**5.** How many nucleons are present in an atom of the element having the symbol $^{235}_{101}A$ ? Find also the number of neutrons in the atom.

**Sol.** From the symbol, it is evident that; Mass number = 235 = number of protons + number of neutrons = number of nucleons; Atomic number = number of protons = 101

So, Number of neutrons = mass number – atomic number = (235 – 101) = 134

## PRACTICE PROBLEMS

**1.** Give the symbol of element which is isoelectronic to $Cl^-$ and $S^{2-}$.

**2.** What is the number of electrons in 1 mol $N^{3-}$ ions?

**3.** Which of the following elements are isotopes and which are isobars?

(i) $^{48}_{18}Ar$, (ii) $^{35}_{17}Cl$, (iii) $^{40}_{20}Ca$, (iv) $^{37}_{17}Cl$, (v) $^{40}_{19}K$

**4.** Calculate the number of protons, electrons and neutrons in a. $^{15}_{8}O^{2-}$ b. $^{27}_{13}Al^{3+}$.

**5.** Calculate the number of electrons, protons and neutrons in Cl, $Cl^-$ and $Cl_2$. Atomic number of chlorine = 17; Mass number = 35.

## WAVE NATURE, BLACK BODY RADIATION, PHOTOELECTRIC EFFECT & BOHR MODEL

- **Drawbacks of Rutherford model:** According to Clark Maxwell, a charged particle moving under the influence of attractive force continuously loses energy in the form of electromagnetic radiations. The electron (a charged body) moving around the nucleus in an orbit must emit radiations and gradually lose energy. As a result of this the motion of electron should slow down and hence it should not be able to withstand the attraction of the nucleus. Consequently, the orbit would become smaller and smaller and the electron move closer and closer to the nucleus following a spiral path.

  Finally, the electron would fall into the nucleus and atom would collapse. Since atom is quite stable and such a collapse of the atom does not take place, there must be some drawbacks in the Rutherford's model of an atom.

  Another Drawback of Rutherford model is it provides no idea about electron distribution around the nucleus and what are the energies of these electrons.

- **Clark Maxwell** (1864) observed that an alternating current of high frequency is capable of radiating continuous energy in the form of waves. He termed these waves as electromagnetic waves or electromagnetic radiations.
- The electromagnetic radiations have following important characteristics.
  - The electromagnetic waves consist of electric and magnetic fields oscillating perpendicular to each other and both are perpendicular to the direction of propagation of radiation.
  - These waves do not require any medium for propagation.
  - All electromagnetic waves travel with the same speed. (i.e. $3.0 \times 10^8$ m $sec^{-1}$ or speed of light)
- All waves have five characteristics by which they can be recognised, viz. wave length, frequency, velocity, wave number and amplitude.
  - **Wave length (λ):** It is the distance between two neighbouring crests or troughs of the wave. It is denoted by Greek letter Lambda (λ) and is measured in Angstrom (Å) or nanometer (nm).

➢ **Frequency (ν):** It is defined as the number of waves which pass through a given point in one second. It is denoted by Greek letter **nu** (ν) and is expressed in units of cycles per second (cps) or Hertz (Hz).

➢ **Velocity:** he distance travelled by a wave in one second is called velocity of the wave. It is denoted by letter c. It is related to frequency and wavelength of the wave by the expression. $c = \nu \times \lambda$ or $\nu = c/\lambda$

➢ **Wave number $(\bar{\nu})$:** It is defined as the number of wave lengths per centimeter. It is equal to the inverse of wavelength expressed in centimeters. It is denoted by $\bar{\nu}$ **(nu bar)** and its unit is $cm^{-1}$.

Thus, $\text{Wave number } (\bar{\nu}) = \dfrac{1}{\text{Wave length } (\lambda)}$ ; $(\nu) = \dfrac{1}{\lambda}$ ; Frequency, $\nu = \dfrac{c}{\lambda} = c\,\bar{\nu}$

➢ **Amplitude (a):** It represent the height of the crest or depth of the trough of a wave. It is denoted by the letter 'a' and determines the intensity or brightness of radiation.

$$\text{Velocity} = \text{Wavelength} \times \text{frequency}$$
$$c = \lambda \times \nu$$

● **Electromagnetic spectrum:** It is defined as the arrangement of various types of electromagnetic radiation in terms of increasing (or decreasing) wave lengths (or frequency). The complete range of electromagnetic waves is called electromagnetic spectrum. The wavelength of various waves increases in the following order

Cosmic rays < γ-rays < X-rays < UV rays < Visible < IR rays < Micro waves < Radio waves

● The instrument used to separate and analyse the seven colour of spectrum is known as **spectrograph** or **spectroscope.**

**Limitations of electromagnetic wave theory:** Electromagnetic wave theory successfully explained the properties of light such as interference and diffraction but it could not explain the phenomenon of photo electric effect and black body radiations.

● **Planck's quantum theory:** To explain the phenomenon of photoelectric effect and black body radiations. Max Planck postulated theory of radiations, which was further extended by Einstein (1905).

Important postulates of this theory are

➢ The radiant energy is not emitted or absorbed continuously but discontinuously in the form of small packets of energy called **quantum.** For light the quantum of energy is termed as **photon.**

➢ The energy associated with each quantum is directly proportional to the frequency of the radiation i.e.

$$E \propto \nu \text{ or } E = h\nu \text{ ........(i)}$$

Where h is called Planck's constant.

Numerically $h = 6.62 \times 10^{-34}$ Joules sec or $6.62 \times 10^{-27}$ erg sec.

➢ The total amount of energy emitted or absorbed by a body will be whole number multiple of quantum by an integer n i.e. $E = nh\nu$

where n is an integer e.g., n = 1, 2, 3 ........... etc.

**Photoelectric effect:**

Phenomenon of ejection of electrons from the surface of a metal when light of suitable frequency strikes on it is called photoelectric effect.

(I) **Threshold frequency ($\nu_0$):** The minimum frequency of incident radiation to cause the photoelectric effect is called threshold frequency.

(II) **Work function:** A part of the photons energy that is absorbed by the metal surface to release the electron is known as work function of the surface denoted by φ. The remaining part of the energy of photons goes into the Kinetic energy of the electron emitted.

If $\nu_0$ is the threshold frequency and ν the frequency of incident light then $\phi = h\nu_0$ and $E = h\nu$.

$K.E. = E - \phi = h\nu - h\nu_0 = h(\nu - \nu_0)$

Note: · K.E. is independent of the intensity of light.

● Number of photoelectrons ∝ intensity of light.

● K.E. is directly proportional to frequency of incident light.

- A **continuous spectrum** is one in which all wave lengths of radiations are so inter mixed that there is no line separation between two colours. As violet merges into blue, blue into green and so on.
- Spectrum of radiation emitted by a substance after absorption of energy is called an emission spectrum.
- When a gas, solid or transparent liquid is placed in the path of bright light emitted from any source and then analysed by the spectroscope, dark lines are observed in the spectrum. This type of spectrum is called **absorption spectrum.**

  The study of emission or absorption spectra is referred to as spectroscopy.
- A **line spectrum** is that spectrum in which only certain wave lengths or bright coloured lines appear on a dark background. Atomic spectra of most elements is line spectrum. Each element has a unique line emission spectra which is used in chemical analysis to identify unknown atoms.

- When radiations emitted by hydrogen gas at low pressure taken in a discharge tube and examined by a spectroscope the line spectrum so obtained is a called the **atomic spectrum of hydrogen.**
- The spectrum consists of a large number of sharp lines each corresponding to a particular wavelength (or frequency) of light emitted by hydrogen atoms. These lines, which are present in ultraviolet, visible and infrared regions are grouped together into different series named after the discoverers.

  Hydrogen spectrum has five series

| Spectrum Line | Region | $n_1$ | $n_2$ |
|---|---|---|---|
| Lyman Series | U.V. | 1 | 2, 3, 4.... |
| Balmer Series | Visible | 2 | 3, 4, 5... |
| Paschen Series | I.R. | 3 | 4, 5, 6... |
| Brackett Series | I.R. | 4 | 5, 6, 7... |
| Pfund Series | I.R. | 5 | 6, 7, 8... |

- **Rydberg** gave a general expression which can be applied to all the series of the hydrogen spectrum. This expression is called **Rydberg formula.**

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

where $n_2 > n_1$. For a particular series $n_1$ is constant i.e.

$R_H$ = Rydberg Constant, Z = charge on nucleus, $n_1$, $n_2$ = electronic levels involved in transition,

$\bar{v}$ = Wave number $\quad R_H = \frac{2\pi^2 me^4}{ch^3}$ = 109677.76 $cm^{-1}$ also for hydrogen

- **Bohr's model of atom**

  Bohr's model for hydrogen atom is based on the following postulates :
- An atom consist of a dense positively charged nucleus situated at the centre surrounded by electrons. The electrons revolves around the nucleus in certain selected circular paths, called orbits without emitting any energy.
- Only those orbits are permitted in which the angular momentum of the electron is an integral multiple of $\frac{h}{2\pi}$ (where h is plank's constant).
- The electrons revolve only in those orbits which have fixed value of energy.
- As long as the electron remains in a particular orbit. It neither lose or gain energy.
- When energy is supplied to an electron, it may jump instantaneously from a lower energy level (say K) to a higher energy level (say L, M, N, ......etc) by absorbing one or more quanta of energy.

  Frequency of radiation absorbed or emitted during transition $\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$
- Frequency (v) associated with the absorption and emission of photon

$$\nu = \frac{\Delta E}{h} = \frac{R_H}{h}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \quad ; \quad \vec{\nu} = \frac{\nu}{C} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = 1.09677 \times 10^7 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \text{ m}^{-1}$$

- **The main features of Bohr's model** are
  - It explains the stability of an atom.
  - It helped in calculating the energy of an electron in a particular orbit of hydrogen.
  - It explains the line spectrum of hydrogen.
  - Simultaneous appearance of a large number of lines in the spectrum of hydrogen: These spectral lines have been grouped into five series namely (i) Lyman series (ii) Balmer series (iii) Paschen series (iv) Brackett series (v) Pfund series.
  - It helped in calculated of radius of Bohr's orbit.
  - Calculation of velocity of electron in Bohr's orbit.
  - Bohr's theory can also be applied to the ions containing only one electron. For example, $He^+$, $Li^{2+}$, $Be^{3+}$ and so on.
- **Limitations of Bohr's theory:** Bohr's theory was unable to explain
  - line spectra of multielectron atom.
  - the presence of multiple spectral lines.
  - the splitting of spectral lines in magnetic field. (Zeeman effect).
  - the splitting of lines in electrical field (Stark effect).
  - the ability of atoms to form molecules by chemical bonds
- **The important formulas related to Bohr's model of atom**
  - The expression for energy of an electron in the nth orbit of hydrogen is $E_n = \frac{2k^2\pi^2 me^4 Z^2}{n^2h^2}$

    where, k = coulomb's law constant m = mass of the electron
    e = charge on the electron h = planck's constant Z = atomic number of element
    Substituting the values of constants.
    $\pi = 3.14$ $m = 9.1 \times 10^{-31}$ kg $k = 9 \times 10^9$ $Nm^2/c^2$
    $h = 6.62 \times 10^{-34}$ J $sec^{-1}$ $e = 1.6 \times 10^{-19}C$ Z = 1 (for hydrogen)

    $$E_n = -\frac{21.79 \times 10^{-19}}{n^2} \text{J/atom} = -\frac{1312}{n^2} \text{kJ/mole}$$

  - For a particular atom the relationship of energy of nth orbit ($E_n$) with that of first orbit ($E_1$) is given by the expression.

    $$E_n = \frac{E_1}{n^2} \qquad E_n = -R_H\left(\frac{1}{n^2}\right) \qquad R_H = 2.18 \times 10^{-18} \text{ J Rydberg constant}$$

    Energy of the electron of an atom (say $He^+$ of $Li^{2+}$) in $n^{th}$ orbit.

    $$E_n = -2.18 \times 10^{-18}\left(\frac{Z^2}{n^2}\right) \text{J} \qquad E_n(Li^{2+}) = \frac{E_n(H) \times Z^2}{n^2}$$ and radii by the expression $$r_n = \frac{52.9(n^2)}{Z} \text{ pm}$$

  - The expression for radius of $n^{th}$ orbit is: $r_n = \frac{n^2h^2}{4\pi^2 mZe^2\, k}$

    where $r_n$ = Radius of the orbit m = Mass of the electron ($9.1 \times 10^{-31}$ kg)

    Now substituting the values of π, h, m, e and k we get $r_0 = \frac{(6.62 \times 10^{-34})^2}{4 \times (3.14)^2 \times 9 \times 10^{-31} \times (1.6 \times 10^{-19}) \times 9 \times 10^9}$

    $= 5.290 \times 10^{-11}$ m = 0.529 Å

    Therefore, radius of the $n^{th}$ orbit for the hydrogen atom may be written as $r_n = 0.529 \times n^2$Å

  - Velocity of the revolving electron in $n^{th}$ orbit is given by the expression $v_n = \frac{2\pi\, k\, Ze^2}{nh}$

    where k = Columb's law constant Z = Atomic number of element
    e = Charge on the electron h = Plank's constant

## ILLUSTRATIONS

**1.** Calculate the energy of one mole quanta of radiation whose frequency is $5 \times 10^{10}$ sec$^{-1}$.

**Sol.** Energy of 1 mol quanta = $N h\nu = 6.023 \times 10^{23} \times 6.626 \times 10^{-34} \times 5 \times 10^{10} = 19.95$ J mol$^{-1}$

**2.** Calculate the energy associated with a photon of light having a wavelength of 6000 Å [$h = 6.624 \times 10^{-27}$ erg-sec].

**Sol.** We know that $E = h \cdot v = h \cdot \dfrac{c}{\lambda}$;

$h = 6.624 \times 10^{-27}$ erg-sec, $c = 3 \times 10^{10}$ cm/sec and $\lambda = 6000 \times 10^{-8}$ cm

So, $E = \dfrac{(6.624 \times 10^{-27}) \times (3 \times 10^{10})}{6 \times 10^{-5}} = 3.312 \times 10^{-12}$ erg

**3.** Find the number of quanta of radiations of frequency $4.67 \times 10^{13}$ s$^{-1}$ that must be absorbed in order to melt 5 g of ice. The energy required to melt 1 g of ice is 333 J.

**Sol.** Energy required to melt 5 g of ice = 5 × 333 = 1665 J

Energy associated with one quantum = $h\nu = (6.62 \times 10^{-34}) \times (4.67 \times 10^{13}) = 30.91 \times 10^{-21}$ J

Number of quanta required to melt 5 g of ice $= \dfrac{1665}{30.91 \times 10^{-21}} = 53.8 \times 10^{21} = 5.38 \times 10^{22}$

**4.** Calculate the radius of Bohr's 3rd orbit in $Li^{2+}$ ion.

**Sol.** We know that, $r_n = \dfrac{n^2}{Z} \times 0.529$ Å; When n = 3 and Z = 3, $r_3 = \dfrac{3^2}{3} \times 0.529$Å $= 3 \times 0.529 = 1.587$ Å

**5.** Calculate the shortest and longest wavelength in hydrogen spectrums of Lyman series.

Or

Calculate the wavelengths of the first line & the series limit for the Lyman series for hydrogen. [$R_H$=109678 cm$^{-1}$]

**Sol.** For Lyman series $n_1 = 1$.

For shortest wavelength in Lyman series (i.e., series limit), the energy difference in two state showing transition should be maximum, i.e., $n_2 = \infty$.

$$\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{(\infty)^2}\right] = R_H\,; \qquad \lambda = \frac{1}{109678} = 9.117 \times 10^{-6} \text{ cm} = 911.7 \text{ Å}$$

For longest wavelength in Lyman series (i.e., first line) the energy difference in two states showing transition should be minimum, i.e., $n_2 = 2$.

$$\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{2^2}\right] = \frac{3}{4}R_H \qquad \text{or} \qquad \lambda = \frac{4}{3} \times \frac{1}{R_H} = \frac{4}{3 \times 109678} = 1215.7 \times 10^{-8} \text{ cm} = 1215.7 \text{ Å}$$

**6.** The minimum energy required to overcome the attractive forces between an electron and the surface of Ag metal is $5.52 \times 10^{-19}$ J. What will be the maximum kinetic energy of electrons ejected out from Ag which is being exposed to UV-light of $\lambda$ = 360 Å?

**Sol.** Energy of the photon absorbed $= \dfrac{hc}{\lambda} = \dfrac{6.626 \times 10^{-27} \times 3 \times 10^{10}}{360 \times 10^{-8}} = 5.52 \times 10^{-11}$ erg $= 5.52 \times 10^{-18}$ J

E(photon) = Work function + K.E.; K.E. = $5.52 \times 10^{-18} - 7.52 \times 10^{-19} = 47.68 \times 10^{-19}$ J

**7.** When certain metal was irradiated with light of frequency $3.2 \times 10^{16}$ Hz, the photoelectrons emitted had twice the kinetic energy as did photoelectrons emitted when the same metal was irradiated with light of frequency $2.0 \times 10^{16}$ Hz. Calculate $v_0$ for the metal.

**Sol.** Applying photoelectric equation, K.E. = hν – hν$_0$ or $(\nu - \nu_0) = \frac{K.E.}{h}$

Given: $K.E._2 = 2K.E._1$; $\nu_2 - \nu_0 = \frac{K.E._2}{h}$ ...(i) and $\nu_1 - \nu_0 = \frac{K.E._1}{h}$ ... (ii)

Dividing equation (i) by equation (ii), $\frac{\nu_2 - \nu_0}{\nu_1 - \nu_0} = \frac{K.E._2}{K.E._1} = \frac{K.E._2}{K.E._1} = 2$ or $\nu_2 - \nu_0 = 2\nu_1 - 2\nu0$

or $\nu_0 = 2\nu_1 - \nu_2 = 2(2.0 \times 10^{16}) - (3.2 \times 10^{16}) = 8.0 \times 10^{15}$ Hz

**8.** Threshold wavelength of a metal is 230 nm. What will be the kinetic energy of photoelectrons ejected when the metal is irradiated with wavelength 180 nm? ($h = 6.626 \times 10^{-34}$ J sec).

**Sol.** Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

$\frac{hc}{\lambda} = \frac{hc}{\lambda_0}$ + K.E. or K.E. $= hc\left[\frac{1}{\lambda} - \frac{1}{\lambda_0}\right] = 6.626 \times 10^{-34} \times 3 \times 10^8 \left[\frac{1}{180 \times 10^{-9}} - \frac{1}{230 \times 10^{-9}}\right]$

$= 6.626 \times 10^{-34} \times 3 \times 10^8 \times 10^9 \left[\frac{1}{180} - \frac{1}{230}\right] = 1.9878 \times 10^{-16} [5.55 \times 10^{-3} - 4.347 \times 10^{-3}] = 2.4 \times 10^{-19}$ J

## PRACTICE PROBLEMS

**1.** At the closest approach, the distance between Mars and Earth is found to be 58 million km. When the two planets are at this nearest distance, how long would it take to send a radio message from a space probe from Mars to Earth.

**2.** Calculate the frequency and wave number of radiation with wavelength 480 nm.

**3.** Calculate the velocity of an electron in Bohr's first orbit of hydrogen atom. (Given, $r = 0.53 \times 10^{-10}$ m)

**4.** Calculate the wavelength of the spectral line obtained in the spectrum of $Li^{2+}$ ion, when the transition takes place between two levels whose sum is 4 and difference is 2.

**5.** If the energy difference between two electronic states is 214.68 kJ mol$^{-1}$, calculate the frequency of light emitted when an electron is dropped from the higher shell to lower shell. (Planck's constant, $h = 39.79 \times 10^{-14}$ kJ sec mol$^{-1}$.)

**6.** The energy required to remove an electron from the surface of sodium metal is 2.3 eV. What is the longest wavelength of radiation with which it can show photoelectric effect?

**7.** Calculate the kinetic energy of the electron ejected when yellow light of frequency $5.2 \times 10^{14}$ sec$^{-1}$ falls on the surface of potassium metal. Threshold frequency of potassium is $5 \times 10^{14}$ sec$^{-1}$.

## DUAL NATURE, HEISENBERG'S PRINCIPLE & QUANTUM MECHANICAL MODEL

- **Einstein** (1905) suggested that light has dual nature i.e. it behaves like a wave as well as like a particle.
- In 1924 **De Broglie** suggested that just like light all material objects including electron have dual nature.

  **DeBroglie's** concept was experimentally verified by **Davisson and Germer** by observing diffraction effect (a property shown by waves) with an electron beam.
- **De Broglie** derived the following relationship between wavelength (λ) of the wave associated with a particle of mass m moving with velocity v i.e.,

  $\lambda = \frac{h}{mv}$ or $\lambda = \frac{h}{p}$ where h in Planck's constant and p is the momentum of the particle.
- The **Heisenberg's Uncertainty Principle** states that it is impossible to determine simultaneously the exact velocity (or momentum) and position of a small moving particle (i.e., electron) with absolute accuracy. Mathematically, it

may be given as $\Delta x \times \Delta p \geq \frac{h}{4\pi}$ or $\Delta x \times m\Delta v \geq \frac{h}{4\pi}$ $\therefore$ $\Delta x \times \Delta v \geq \frac{h}{4\pi m}$

where $\Delta$ x = Uncertainty in determination of position

$\Delta$ p = Uncertainty in measurement of momentum ($\Delta p = m\,\Delta v$)

h = Planck's constant

This principle is applicable for motion of microscopic objects not applicable in case of macroscopic objects.

- **Schrodinger** (1927) gave a mathematical expression known as a **Schrodinger wave equation.** His theory is based on quantum mechanical model of atom in which the concept of probability of finding the electron at any position around the nucleus at any instant of time is considered.

## ILLUSTRATIONS

**1.** An electron beam emerges from an accelerator with kinetic energy 100 eV. What is its de Broglie wavelength? [$m = 9.1 \times 10^{-31}$ kg, $h = 6.6 \times 10^{-34}$ Js, 1 eV = $1.6 \times 10^{-19}$ J]

**Sol.** Kinetic energy of electron = 100 eV = $100 \times 1.6 \times 10^{-19}$ J = $1.6 \times 10^{-17}$ J

We know, $\lambda = \frac{h}{\sqrt{2Em}} = \frac{6.626 \times 10^{-34}}{\sqrt{2 \times 1.6 \times 10^{-17} \times 9.1 \times 10^{-31}}}$ = $1.228 \times 10^{-10}$ m

**2.** Calculate the wavelength of 1000 kg rocket moving with a velocity of 3000 kilometre per hour.

**Sol.** Mass of rocket (m) = 1000 kg; Velocity of rocket = 3000 km/hr $= \frac{3000 \times 1000}{3600}$ m/sec = 833.33 m/sec

Wavelength associated with the rocket can be calculated as : $\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{1000 \times 833.33}$ = $7.9512 \times 10^{-40}$ m

**3.** Find the number of waves made by a Bohr electron in one complete revolution in the 3rd orbit.

**Sol.** Velocity of the electron in 3rd orbit = $\frac{3h}{2\pi mr}$ where, m = mass of electron and r = radius of 3rd orbit.

Applying de Broglie equation, $\lambda = \frac{h}{mv} = \frac{h}{m} \times \frac{2\pi mr}{3h} = \frac{2\pi r}{3}$; No. of waves $= \frac{2\pi r}{\lambda} = \frac{2\pi r}{2\pi r} \times 3 = 3$

**4.** As electron has a speed of 40 m s$^{-1}$ accurate upto 99.99%. What is the uncertainty in its location?

**Sol.** Given, $\Delta v = \frac{0.01}{100} \times 40$ = 0.004 m s$^{-1}$; We know, $\Delta x \cdot \Delta v \geq \frac{h}{4\pi m}$; $\Delta x \geq \frac{h}{4\pi m \Delta v}$

$\Delta x \geq \frac{6.626 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times 0.004}$ = 0.0145 m

**5.** Calculate the uncertainty in the position of a particle when the uncertainty in momentum is a. $1 \times 10^{-3}$ g cm sec$^{-1}$, b. zero.

**Sol.** Given $\Delta p = 1 \times 10^{-3}$ g cm sec$^{-1}$; $h = 6.62 \times 10^{-27}$ erg-sec; $\pi = 3.142$

According to uncertainty principle, $\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$; So, $\Delta x \geq \frac{h}{4\pi} \cdot \frac{1}{\Delta p}$; $\geq \frac{6.62 \times 10^{-27}}{4 \times 3.142} \times \frac{1}{10^{-3}}$ = $0.527 \times 10^{-24}$ cm

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

**6.** Calculate the uncertainty in velocity of a cricket ball of mass 150 g if the uncertainty in its position is of the order of 1 Å ($h = 6.6 \times 10^{-34}$ kg m$^2$ s$^{-1}$).

**Sol.** $\Delta x \cdot m\,\Delta v = \frac{h}{4\pi}$; $\Delta v = \frac{h}{4\pi\,\Delta x \cdot m} = \frac{6.6 \times 10^{-34}}{4 \times 3.143 \times 10^{-10} \times 0.150}$ = $3.499 \times 10^{-24}$ m s$^{-1}$

## PRACTICE PROBLEMS

**1.** A moving particle is associated with wavelength $5 \times 10^{-8}$ m. If its momentum is reduced to half of its value, calculate the new wavelength.

**2.** Find out the wavelength of a track star running a 100 metre dash in 10.1 sec, if his weight is 75 kg.

**3.** At what velocity ratio are the wavelengths of an electron and a proton be equal? ($m_e = 9.1 \times 10^{-28}$ g and $m_p = 1.6725 \times 10^{-24}$ g)

**4.** Calculate the uncertainty product of velocity and displacement for an electron of mass $9.1 \times 10^{-31}$ kg according to Heisenberg's uncertainty principle. ($h = 6.6 \times 10^{-34}$ Js)

**5.** The uncertainty in measuring the speed of an accelerated electron is $1.2 \times 10^5$ m $s^{-1}$. Calculate the uncertainty in finding its location while it is in motion. ($h = 6.6 \times 10^{-34}$ Js, mass of electron = $9.1 \times 10^{-31}$ kg)

**6.** A proton is accelerated to a velocity of $3 \times 10^7$ and m $s^{-1}$. If the velocity can be measured with a precision of ± 0.5% calculate the uncertainty in the position of the proton.

## ATOMIC ORBITAL, QUANTUM NUMBERS & ELECTRONIC CONFIGURATION

- The region of space around the nucleus within which the probability of finding the electron of a given energy is maximum (nearly 90–95%) is called the atomic orbital.
- An atomic orbital is the wave function ψ for an electron in an atom. The probability of finding an electron at a point within an atom is proportional to the square of the orbital wave function. i.e. $|\psi|^2$.
- **Quantum numbers** may be defined as set of numbers which display complete information about size, shape and orientation of the orbital. These are designated as **principal quantum number (*n*), azimuthal quantum number (*l*), magnetic quantum number (m)** and **spin quantum number(s).**
- The fourth quantum number designated as **spin quantum number** (s), represent the spin of electron i.e. rotation of electron about its own axis.
- **Principal quantum number** refer to the average distance of the electron from the nucleus. It denotes the energy level to which the electron belongs. It denotes the energy level to which electron belongs. It gives information about the maximum number of electrons that can be accommodated in any shell. Number of electrons in any shell is given by expression $2n^2$.
- **Azimuthal quantum number or angular momentum quantum number (*l*):** It tells about the number of subshells within a given principal energy shell to which the electrons belong. For a given values of principal quantum number, 'n', the azimuthal quantum number, '*l*', may have all integral values from 0 to (n–1), each of which represent a different subenergy level or subshell.

  The value of *l* refer to the shapes of the subshells. For example.

| *l* | Subshell | Shape of subshell |
|---|---|---|
| 0 | s | Spherical |
| 1 | p | Dumbbell |
| 2 | d | Double dumbbell |
| 3 | f | complex shape. |

It gives information for the orbital angular momentum, whose value is equal to $\frac{h}{2\pi}\sqrt{l(l+1)}$

**Magnetic Quantum Number (m):** This quantum number describes the behaviour of electron in a magnetic field. This quantum number provides following information about the electron.

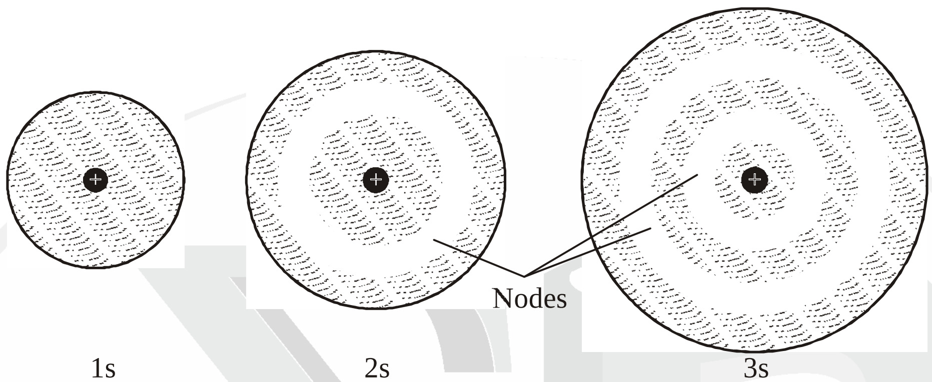
(i) Magnetic quantum number gives the number of permitted orientations of subshells. For example, for a given value of '*l*', the possible values of 'm' range from –*l* through 0 to +*l*. Each value of m corresponds to one atomic orbital. For example,

For s-subshell, $l = 0$ $m = 0$ i.e., s-subshell has **one** orbital

| | | | |
|---|---|---|---|
| For p-subshell, | $l = 1$ | $m = -1, 0, +1$ | i.e., p-subshell has **three** orbitals |
| For d-subshell, | $l = 2$ | $m = -2, -1, 0, +1, +2$ | i.e., d-subshell has **five** orbitals |
| For f-subshell, | $l = 3$ | $m = -3, -2, -1, 0, +1, +2, +3$ | i.e., f-subshell has **seven** orbitals |

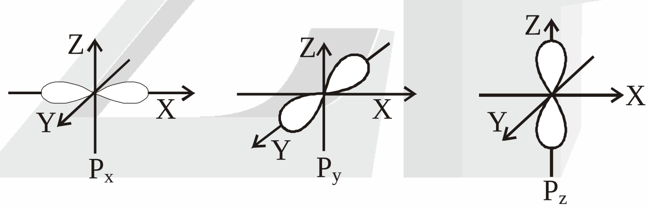
(ii) It explained successfully the splitting of spectral lines in the magnetic field i.e., Zeeman effect.

- **Spin quantum number(s)** accounts for the spinning orientation of the electron. The electron in an orbital can have only two types of spins i.e. in clockwise and anticlockwise direction. Therefore, the spin quantum number can have only two values i.e. $+\frac{1}{2}$ or $-\frac{1}{2}$.
- **Shapes of orbitals —** For s-orbitals $l = 0$, so there is only one value of m i.e. m = 0. Thus, s-orbital can have only one orientation i.e. the probability of finding electron is same in all directions at a given distance from the nucleus. Hence s-orbital is symmetrical around the nucleus and thus has spherical shape. Between each s-orbital there is some region where the probability of finding electrons is minimum or may be considered as zero. This region in called as nodel plane or node.

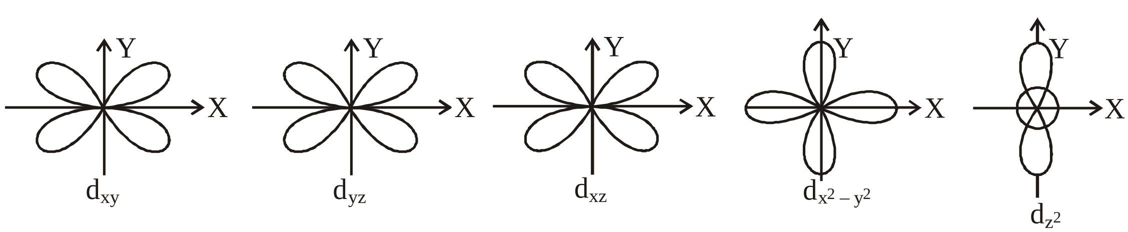



Shapes of 1s, 2s and 3s orbitals

For p-orbitals $l = 1$, the permissible values of m are +1, 0 and –1. Thus, there are three p-orbitals, designated as $p_x$, $p_y$ and $p_z$, in each p-subshell. Each p-orbital has two lobes on the opposite side of the nucleus and separated by the nodal plane as depicted in fig. These three orbitals are identical in shape, size and energy but have orientation along x-axis, y-axis and z-axis.




For d-orbitals l = 2 i.e. the permissible values of m are –2, –1, 0, +1, +2. This indicates for the five d-orbitals designated as $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and $d_{z^2}$. Each d-orbital is identical in shape, size and energy and looks like double dumbbell in shape shown in fig.




The orbitals having the same energy are called degenerate.

- A diagram showing the relative energies of various orbitals in an atom is called energy level diagram.

  The relative order of energies of various sub-shells in a multielectron atom can be predicted with the help of **Bohr-Bury's** rule also known as (n + $\ell$) rule. The orbital with lower value (n + $\ell$) will be filled first

(i) Lower the value of $(n + \ell)$ for an orbital, the lower is its energy. For example 3s orbital has lower energy than 3p orbital.

(ii) If two orbitals have the same value of $(n + \ell)$, the orbital with lower value of n has lower energy. For example, between 2p$(n + \ell = 2 + 1 = 3)$ and 3s $(n + \ell = 3 + 0 = 3)$, 2p orbital has lower energy than 3s.

- The energies of the different subshells present within the same principle shell are in the following order.

$$\xrightarrow[\text{Increasing energy}]{s<p<d<f}$$

- The size and energy of s-orbital increases with the increase of principal quantum number n i.e.

$$\xrightarrow[\text{size and energy increases}]{1s<2s<3s<4s}$$

- **Aufbau's rule:** According to this rule the orbitals are filled progressively in order of their increasing energy beginning with the orbital of lowest energy.
- **Pauli's exclusion principle:** According to this principle "no two electrons in an atom can have the same set of all the four quantum number".
- **Hund's rule of maximum multiplicity:** This rule governs the arrangement of electrons in the orbitals of identical energies e.g., within $p_x$, $p_y$ and $P_z$ or amongst $d_{xy}$, $d_{yz}$, $d_{zx}$, $d_{x^2-y^2}$ and $d_{z^2}$ : According to this rule, electron pairing in orbitals of same energy (i.e. orbitals of p, d or f subshell) shall take place only when all the orbitals of the subshell contain one electron each.
- Elements with atomic number 24 and 29 have shown little deviation form aufbau's rule. Here one electron form 4s orbital shifts to higher energy level 3d orbital. This migration of electron is due to higher stability of fully-filled and half-filled orbitals. The electronic configuration of these elements is given below.

| Element | Atomic number | Expected electronic configuration | observed electronic configuration |
|---|---|---|---|
| Cr | 24 | $[Ar]^{18}3d^4 4s^2$ | $[Ar]^{18}3d^5 4s^1$ |
| Cu | 29 | $[Ar]^{18}3d^9 4s^2$ | $[Ar]^{18}3d^{10} 4s^1$ |

- Orbitals which are either half filled or fully filled are more symmetrical and therefore possess lower energy i.e. extra stability.
- When orbitals are half filled or fully filled, the exchange of electrons between the orbitals is maximum. Such exchanges leads to greater stability of electrons in the orbitals, because low exchange energy results in to higher stabilization.

**THE ELECTRONIC CONFIGURATION OF THE ELEMENTS**

| Atomic number | Symbol | Electronic Configuration configuration |
|---|---|---|
| 1 | H | $1s^1$ |
| 2 | He | $1s^2$ |
| 3 | Li | [He] $2s^1$ |
| 4 | Be | [He] $2s^2$ |
| 5 | B | [He] $2s^2\ 2p^1$ |
| 6 | C | [He] $2s^2\ 2p^2$ |
| 7 | N | [He] $2s^2\ 2p^3$ |
| 8 | O | [He] $2s^2\ 2p^4$ |
| 9 | F | [He] $2s^2\ 2p^5$ |

| | | |
|---|---|---|
| 10 | Ne | [He] $2s^2\ 2p^6$ |
| 11 | Na | [Ne] $3s^1$ |
| 12 | Mg | [Ne] $3s^2$ |
| 13 | Al | [Ne] $3s^2\ 3p^1$ |
| 14 | Si | [Ne] $3s^2\ 3p^2$ |
| 15 | P | [Ne] $3s^2\ 3p^3$ |
| 16 | S | [Ne] $3s^2\ 3p^4$ |
| 17 | Cl | [Ne] $3s^2\ 3p^5$ |
| 18 | Ar | [Ne] $3s^2\ 3p^6$ |
| 19 | K | [Ar] $4s^1$ |
| 20 | Ca | [Ar] $4s^2$ |
| 21 | Sc | [Ar] $4s^2\ 3d^1$ |
| 22 | Ti | [Ar] $4s^2\ 3d^2$ |
| 23 | V | [Ar] $4s^2\ 3d^3$ |
| 24 | Cr | [Ar] $4s^1\ 3d^5$ |
| 25 | Mn | [Ar] $4s^2\ 3d^5$ |
| 26 | Fe | [Ar] $4s^2\ 3d^6$ |
| 27 | Co | [Ar] $4s^2\ 3d^7$ |
| 28 | Ni | [Ar] $4s^2\ 3d^8$ |
| 29 | Cu | [Ar] $4s^1\ 3d^{10}$ |
| 30 | Zn | [Ar] $4s^2\ 3d^{10}$ |
| 31 | Ga | [Ar] $4s^2\ 3d^{10}\ 4p^1$ |
| 32 | Ge | [Ar] $4s^2\ 3d^{10}\ 4p^2$ |
| 33 | As | [Ar] $4s^2\ 3d^{10}\ 4p^3$ |
| 34 | Se | [Ar] $4s^2\ 3d^{10}\ 4p^4$ |
| 35 | Br | [Ar] $4s^2\ 3d^{10}\ 4p^5$ |
| 36 | Kr | [Ar] $4s^2\ 3d^{10}\ 4p^6$ |

## ILLUSTRATIONS

**1.** How many unpaired electrons are present in Fe, $Fe^{2+}$ and $Fe^{3+}$ ? (At. No. of Fe is 26)

**Sol.** The electronic configurations of Fe, $Fe^{2+}$ and $Fe^{3+}$ are:

| | | 3d$^6$ | 4s$^2$ | |
|---|---|---|---|---|
| Fe | $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ | [Ar] $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ | $\boxed{\uparrow\downarrow}$ | 4 unpaired electrons |
| $Fe^{2+}$ | $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$ | [Ar] $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ | | 4 unpaired electrons |

$Fe^{3+}$ $1s^2 2s^2 2p^6$ $3s^2 3p^6 3d^5$ $\quad$ [Ar] $\overset{3d^5}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}$ $\quad$ 5 unpaired electrons

**2.** Write the electronic configuration of the element with Z = 16 and predict the (i) number of p-electrons (ii) number of filled orbitals and (iii) number of half-filled orbitals.

**Sol.** The electronic configuration of the element with Z = 16 is: $1s^2\ 2s^2\ 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$

(i) No. of p-electrons = 10 $\quad$ (ii) No. of filled orbitals = 7 $\quad$ (iii) No. of half-filled orbitals = 2

**3.** One unpaired electron in an atom contributes a magnetic moment of 1.1 B.M. Calculate the magnetic moment for manganese (At. No. of Mn = 25).

**Sol.** Electronic configuration of manganese atom is : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

Number of unpaired electrons = 5; $\quad$ Magnetic moment for Mn = 1.1 × 5 = 5.5 B. M.

**4.** Give the atomic number of the elements whose outermost electrons are represented by :

(i) $3d^5$ $\quad$ (ii) $3p^3$ $\quad$ (iii) $3s^2$

**Sol.** (i) Configuration of element : $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^5, 4s^2$

Total number of electrons = 25; $\quad$ Atomic number of element = 25

(ii) Configuration of element: $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^3$; Total number of electrons = 15; Atomic number of element =15

(iii) Configuration of element : $1s^2, 2s^2\ 2p^6, 3s^2$; Total number of electrons = 12; Atomic number of element= 12

**5.** Arrange the electrons represented by the following sets of quantum number in decreasing order of energy :

(1) n = 4 $\quad l = 0 \quad m_e = 0 \quad m_s = +\frac{1}{2}$ $\qquad$ (2) n = 3 $\quad l = 1 \quad m_e = 1 \quad m_s = -\frac{1}{2}$

(3) n = 3 $\quad l = 2 \quad m_e = 0 \quad m_s = +\frac{1}{2}$ $\qquad$ (4) n = 3 $\quad l = 0 \quad m_e = 0 \quad m_s = -\frac{1}{2}$

**Sol:** Find (n + $l$) for each set

For (1): (n + $l$) = 4, i.e., 4s $\qquad$ For (2) : (n + $l$) = 5, i.e., 3p

For (3) : (n + $l$) = 5, i.e., 3d $\qquad$ For (4) : (n + $l$) = 3, i.e., 3s

(1) Lower is the value of (n + $l$), lower is energy level.

(2) If (n + $l$) are same then orbital with lower values of n possess lower energy.

Decreasing order of energy 3 > 1 > 2 > 4.

**6.** Give orbital angular momentum of the following : (i) 3s $\quad$ (ii) 4p $\quad$ (iii) 3d

**Sol.** Orbital angular momentum = $\sqrt{l(l+1)}\dfrac{h}{2\pi}$

(i) For 3s : $l = 1$ $\quad \therefore$ Orbital angular momentum = 0

(ii) For 4p : $l = 1$ $\quad \therefore$ Orbital angular momentum = $\sqrt{l(l+1)}\dfrac{h}{2\pi} = \sqrt{2}\dfrac{h}{2\pi}$

(iii) For 3d : $l = 2$ $\quad \therefore$ Orbital angular momentum = $\sqrt{l(l+1)}\dfrac{h}{2\pi} = \sqrt{2\times 3}\dfrac{h}{2\pi} = \sqrt{6}\dfrac{h}{2\pi}$

**7.** Arrange following orbitals in decreasing order of effective nuclear charge : 3s, 3p, 3d

**Sol.** Closer is the orbital to the nucleus, greater is its effective nuclear charge. $\xrightarrow[\text{Decreasing effective nuclear charge}]{3s > 3p > 3d}$

**8.** Electronic configuration of both $Na^+$ and $Mg^{2+}$ are similar. In which of these ions, the effective nuclear charge of 2p electrons will be greater?

**Sol.** $Na^+ = 1s^2, 2s^2\ 2p^6$; $Mg^{2+} = 1s^2, 2s^2\ 2p^6$; Effective nuclear charge is calculated as, $Z^* = Z - \sigma$

Due to greater nuclear charge (Z), the 2p electrons as $Mg^{2+}$ will experience more effective nuclear charge.

## PRACTICE PROBLEMS

**1.** Which of the following orbitals are not possible?

(i) 1p (ii) 3f (iii) 5g (iv) 4f (v) 2d

**2.** Designate the orbitals having:

(i) n = 2, $l$ = 0 (ii) n = 5, $l$ = 4 (iii) n = 4, $l$ = 3 (iv) n = 6, $l$ = 2

**3.** How many electrons can be accommodated at most in :

(i) n = 4, $s = +\frac{1}{2}$ (ii) n = 5, $l$ = 3 (iii) all the orbitals with n = 3 (iv) n = 3, $l$ = 2, m = –1 and $s = -\frac{1}{2}$

**4.** What is the atomic number of element with valence shell configuration $6s^2\ 6p^4$ ?

**5.** Which atoms have as may 's' electrons as 'p' electrons?

**6.** Write down all four quantum numbers for : a. 19th electron of $_{24}Cr$ b. 21st electron of $_{21}Sc$.

## TEXT BOOK QUESTIONS

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

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

**2.** (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 the mass of a neutron = $1.675 \times 10^{-27}$ 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 temperature and pressure are changed?

**3.** How many neutrons and protons are there in the following nuclei? $^{13}_{6}C$, $^{16}_{8}O$, $^{24}_{12}Mg$, $^{56}_{26}Fe$, $^{88}_{38}Sr$

**4.** 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

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

**6.** Find energy of each of the photons which

(i) correspond to light of frequency $3 \times 10^{15}$ Hz (ii) have wavelength of 0.50 Å

**7.** Calculate the wavelength, frequency and wave number of a light wave whose period is $2.0 \times 10^{-10}$s.

**8.** What is the number of photons of light with a wavelength of 4000 pm that provide 1J of energy?

**9.** A photon of wavelength $4 \times 10^{-7}$ m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate

(i) the energy of the photon (eV) (ii) the kinetic energy of the emission and

(iii) the velocity of the photoelectron (1 eV = $1.6020 \times 10^{-19}$ J).

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

**11.** A 25 watt bulb emits monochromatic yellow light of wavelength of 0.57$\mu$m. Calculate the rate of emission of quanta per second.

12. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency ($\nu_0$) and work function ($W_0$) of the metal.

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

14. How much energy is required to ionise a H atom if the electron occupies n = 5 orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from n = 1 orbit).

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

16. (i) The energy associated with the first orbit in the hydrogen atom is $-2.18 \times 10^{-18}$ J atom$^{-1}$. What is the energy associated with the fifth orbit?

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

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

18. 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 the light emitted when the electron returns to the ground state? The ground state electron energy is $-2.18 \times 10^{-11}$ ergs.

19. 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?

20. Calculate the wavelength of an electron moving with a velocity of $2.05 \times 10^7$ ms$^{-1}$.

21. The mass of an electron is $9.1 \times 10^{-31}$ kg. If its K.E. is $3.0 \times 10^{-25}$ J, calculate its wavelength.

22. Which of the following are isoelectronic species i.e., those having the same number of electrons?

$Na^+$, $K^+$, $Mg^{2+}$, $Ca^{2+}$, $S^{2-}$, Ar.

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

a. $H^-$ b. $Na^+$ c. $O^{2-}$ d. $F^-$

(ii) What are the atomic numbers of elements whose outermost electrons are represented by

a. $3s^1$ b. $2p^3$ c. $3p^5$?

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

a. |He| $2s^1$ b. |Ne| $3s^2 3p^3$ c. |Ar| $4s^2 3d^1$.

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

25. An electron is in one of the 3d orbitals. Give the possible values of n, $\ell$ and $m_\ell$ for this electron.

26. 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.

27. Give the number of electrons in the species $H_2^+$, $H_2$ and $O_2^+$.

28. (i) An atomic orbital has n = 3. What are the possible values of $l$ and $m_\ell$?

(ii) List the quantum numbers ($m_\ell$ and $\ell$) of electrons for 3d– orbital.

(iii) Which of the following orbitals are possible? 1p, 2s, 2p and 3f.

29. Using s, p, d notations, describe the orbital with the following quantum numbers:

a. n = 1, $\ell = 0$ b. n = 3, $\ell = 1$ c. n = 4, $\ell = 2$ d. n = 4, $\ell = 3$.

30. Explain, giving reasons, which of the following sets of quantum numbers are not possible.

a. n = 0, $\ell = 0$, $m_1 = 0$, $m_s = +1/2$ b. n = 1, $\ell = 0$, $m_1 = 0$, $m_s = -1/2$

c. n = 1, $\ell = 1$, $m_1 = 0$, $m_s = +1/2$ d. n = 2, $\ell = 1$, $m_1 = 0$, $m_s = -1/2$

e. n = 3, $\ell = 3$, $m_1 = -3$, $m_s = +1/2$ f. n = 3, $\ell = 1$, $m_1 = 0$, $m_s = +1/2$

31. How many electrons in an atom may have the following quantum numbers? a. n = 4, $m_s = -1/2$ b. n = 3, $\ell = 0$

32. 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.

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

34. Calculate the energy required for the process $He^+(g) \longrightarrow He^{2+}(g) + e^-$

The ionization energy for the H atom in the ground state is $2.18 \times 10^{-18}$ J $atom^{-1}$.

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

36. $2 \times 10^8$ atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

37. The diameter of zinc atom is 2.6 Å. Calculate a. radius of zinc atom in pm and b. number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

38. A certain particle carries $2.5 \times 10^{-16}$C of static electric charge. Calculate the number of electrons present in it.

39. In Milikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the static electric charge on the oil drop is $-1.282 \times 10^{-18}$ C, calculate the number of electrons present on it.

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

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

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

43. 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.

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

45. Arrange the following type of radiations in increasing order of frequency:

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.

46. Nitrogen laser produces a radiation at a wavelength of 337.1 nm. If the number of photons emitted is $5.6 \times 10^{24}$, calculate the power of this laser.

47. Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate

a. the frequency of emission b. distance traveled by this radiation in 30 s

c. energy of quantum and d. number of quanta present if it produces 2 J of energy.

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

49. Lifetimes of the molecules in the excited states 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 phons emitted during the pulse source is $2.5 \times 10^{15}$, calculate the energy of the source.

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

**51.** 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 with a wavelength 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

**52.** Following results are observed when sodium metal is irradiated with different wavelengths. Calculate a. threshold wavelength and b. Planck's constant.

| | | | |
|---|---|---|---|
| $\lambda$ (nm) | 500 | 450 | 400 |
| $v \times 10^{-5}$ ($cms^{-1}$) | 2.55 | 4.35 | 5.35 |

**53.** The ejection of the photoelectron 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.

**54.** 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 $1.5 \times 10^7$ m $s^{-1}$, calculate the energy with which it is bound to the nucleus.

**55.** Emission transitions in the Paschen series end at orbit n = 3 and start form orbit n and can be represented as $\nu = 3.29 \times 10^{15}$ (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.

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

**57.** 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 $1.6 \times 10^6$ $ms^{-1}$, calculated de Broglie wavelength associated with this electron.

**58.** 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.

**59.** If the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6$ $ms^{-1}$, calculate the de Broglie wavelength associated with it.

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

**61.** 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 $h/4\pi$ m × 0.05 nm, is there any problem in defining this value.

**62.** 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) $n = 4, \ell = 2, m_\ell = -2, m_s = -1/2$ (ii) $n = 3, \ell = 2, m_\ell = -1, m_s = +1/2$

(iii) $n = 4, \ell = 1, m_\ell = 0, m_s = +1/2$ (iv) $n = 3, \ell = 2, m_\ell = -2, m_s = -1/2$

(v) $n = 3, \ell = 1, m_\ell = -1, m_s = +1/2$ (vi) $n = 4, \ell = 1, m_\ell = 0, m_s = +1/2$

**63.** 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?

**64.** Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge?

(i) 2s and 3s (ii) 4d and 4f (iii) 3d and 3p

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

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

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

**67.** 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?

# EXERCISE

## PART I VERY SHORT ANSWER QUESTIONS

**1.** From the following nuclei, select the isotopes and isobars: ${}^{238}_{92}U$, ${}^{234}_{90}Th$, ${}^{234}_{92}U$, ${}^{234}_{91}Pa$

**2.** Name a species which has no neutron.

**3.** Which experiment led to the discovery of neutrons?

**4.** The atomic number of Uranium is 92. How many electrons are there in the neutral atom? How many electrons and protons are there in the $U^{+2}$ ion?

**5.** Name the element whose isotope has mass number 14 and 8 neutrons.

**6.** An element of atomic weight Z consists of two isotopes of mass number Z – 1 and Z + 2. Calculate the % of higher isotope.

**7.** Give an example of a neutral molecule which is iso—electronic with $CIO^{-1}$?

**8.** What did Rutherford's alpha scattering experiment prove?

**9.** What is the ratio of mass of electron to the mass of proton?

**10.** How many protons and neutrons are in an element which is represented by the symbol ${}^{125}X_{56}$ have in its nucleus?

**11.** How does the intensity of spectral line vary with wave length?

**12.** What is Stark effect?

**13.** What is Zeeman effect?

**14.** What type of spectrum is obtained when light emitted form discharge tube containing hydrogen gas is analysed with a spectroscope?

**15.** Explain why are Bohr's orbitals also called energy levels?

**16.** What is value of angular momentum for an electron in 5th orbit according to Bohr's theory?

**17.** Which has the greater energy–a photon of violet light or a photon of green light?

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

**19.** Name the experiment evidence to support the wave nature of light.

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

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

**22.** How is photon different from proton?

**23.** Why splitting of spectral lines takes place when the source giving the spectrum is placed in a magnetic field?

**24.** One photon of ultraviolet light can eject a photo electron from the surface of a certain metal. When the same metal is irradiated with 2 photons of red light having total energy equal to that of ultraviolet light; will photoelectron be ejected?

**25.** Calculate the radius of first two orbits of $Li^{2+}$?

**26.** If the electron absorbs 12.1 eV of energy, it will jump to which orbit? (Given energy of electron is –13.6eV)

**27.** If the wavelength of green light is about 5000 Å, what is its frequency? ($c = 3 \times 10^8$ m/s)

**28.** Calculate the kinetic energy of the ejected electron when ultra-violet radiation of frequency $1.6 \times 10^{15}$ $s^{-1}$ strikes the surface of potassium metal. Threshold frequency of potassium is $5 \times 10^{14}$ $s^{-1}$. ($h = 6.63 \times 10^{-34}$ Js)

29. When moving with the same velocity which one of the following particles has the largest de Broglie wavelength and why? a. Electron b. Proton c. $\alpha$-particle.

30. How many electrons in the p-orbital of argon (Z = 18) have same spin?

31. Can an electron have the quantum number values as n = 2, $l$ = 2, $m_l$ = +2?

32. What in the sequence of energies of 3s, 3p and 3d orbitals in a. a hydrogen atom b. a multielectron atom

33. Which energy level do not have p-orbital?

34. Which orbital is non-directional in nature?

35. What is the lowest shell which has an f-subshell?

36. How are $d_{xy}$ and $d_{x^2-y^2}$ orbitals related?

37. How many nodal spheres does a 5s electron charge cloud have?

38. What is the probability of finding a 4d electron right at the nucleus?

39. What is the maximum number of electrons that can occupy g subshell ($l$ = 4)?

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

41. What is difference between the notations $l$ and L?

42. Why is 4s orbital filled before 3d orbital?

43. How many quantum numbers specify an a. electron, b. orbital? Name them

44. State physical significance of $\psi^2$.

45. The electronic configuration of N(7) is $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ and not $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1$. Why?

46. What would you predict for the atomic number of the noble gas beyond Rn, if such an element had sufficient stability to be prepared or observed? Assume that 'g' orbitals are still not occupied in the ground states of the preceding elements.

47. Explain why chromium has only one electron in its 4s sub-shell?

48. Which properties of the elements depend on the electronic configuration of the atoms and which do not?

49. What is the difference between angular momentum of an electron present in 2p and in 3p orbital?

## PART II SHORT ANSWER QUESTIONS

1. Explain how can you say electrons and protons are fundamental particles of all the atoms.

2. What is the origin of anode rays in the discharge tube? Name the particles which form anode rays.

3. Describe the drawback of Rutherford's model of atom.

4. What is the evidence that cathode rays are a part of all matter?

5. Why does the charge to mass ratio of positive rays depends on the gas taken in the discharge tube whereas charge to mass ratio of cathode rays is same for all gases?

6. Calculate and compare the energies of two radiations one with a wavelength of 800 nm and other with wavelength of 400 nm.

7. Calculate the wavelength of an electron that has been accelerated in a particle accelerator through a potential difference of 100 million volts.

   (1 eV = $1.6 \times 10^{-19}$ J, me = $9.1 \times 10^{-31}$ kg, h = $6.6 \times 10^{-34}$ Js, c = $3.0 \times 10^{8}$ $ms^{-1}$)

8. In photoelectric effect experiment irradiation of a metal with light of frequency $5 \times 10^{20}$ $s^{-1}$ yields electrons with maximum K.E. = $6.63 \times 10^{-14}$ J. Calculate $v_0$ (threshold frequency) for the metal.

**9.** What is the wavelength for the electron accelerated by $1.0 \times 10^4$ volts?

**10.** The first member ($H_\alpha$ line) of the Balmer series of hydrogen has a wavelength of 6563 Å. Calculate the wavelength of the second member ($H_\beta$ line).

**11.** Iodine molecule dissociates into atoms after absorbing light of 4500 Å. If one quantum of radiation is absorbed by each molecule, calculate the kinetic energy of iodine atoms. (Bond energy of $I_2$ = 240 kJ $mol^{-1}$)

**12.** Energy in a Bohr orbit is given to be equal to $-\frac{B}{n^2}$, where B = $2.179 \times 10^{-18}$ J. Calculate the wavelength of the emitted radiation when electron jumps from the third orbit to the second.

**13.** A sheet of silver is illuminated by monochromatic ultraviolet radiations of wavelength = 1810 Å. What is the maximum energy of the emitted electron? Threshold wavelength of silver is 2640 Å.

**14.** An electron beam can undergo diffraction by crystals. Through what potential should a beam of electrons be accelerated so that its wavelength becomes equal to 1.54 Å.

**15.** Two particles A and Bare in motion. If the wavelength associated with particle A is $5 \times 10^{-8}$m, calculate the wavelength associated with particle B if its momentum is half of A.

**16.** A photon of wavelength $4 \times 10^{-7}$ m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate

a. the energy of the photon in eV. b. the kinetic energy of the emitted electron in joules.

c. the velocity of the photoelectron. (1 eV = $1.6020 \times 10^{-19}$ J)

**17.** One of the lines in the Balmer series of the hydrogen atom emission spectrum is at 397 nm. It results from a transition from an upper energy level to n = 2. What is the principal quantum number of the upper level?

**18.** Ionization energy of a H-atom is 13.6 eV/atom. It requires a photon of energy 1.5 times the minimum which is required to remove the electron. Calculate the wavelength of the emitted electron.

**19.** What is the frequency and wave-length of a photon emitted during a transition from the n = 5 state to the n = 2 state in the hydrogen atom.

**20.** Calculate (i) wave number (ii) frequency of yellow radiation having wavelength of 5800 Å.

**21.** What are the main points of Planck's quantum theory?

**22.** The ionization energy of $He^+$ is $19.6 \times 10^{-18}$ J $atom^{-1}$. Calculate the energy of the stationary state of $Li^{2+}$.

**23.** Which of the following relate to light as wave, particle, or both?

A) diffraction B) photoelectric effect C) $E = mc^2$ D) $E = h\nu$.

**24.** An electron in a hydrogen atom in its ground state absorbs 1.5 times as much energy as the minimum required for it to escape from the atom. What is the wavelength of the emitted electron?

**25.** An electron and a proton are possessing the same amount of kinetic energy. Which of the two have greater wavelength?

**26.** When light of frequency v is thrown on a metal surface with threshold frequency $v_0$, photo-electrons are emitted with maximum kinetic energy = $1.3 \times 10^{-18}$J. If the ratio, $v : v_0 = 3 : 1$, calculate the threshold frequency $v_0$.

**27.** Define (i) Photo-electric effect (ii) Black body radiations.

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

**29.** According to de Broglie, matter should exhibit dual behaviour, that is both particle and wave like properties. However, a cricket ball of mass 100 g does not move like a wave when it is thrown by a bowler at a speed of 100 km/h. Calculate the wavelength of the ball and explain why it does not show wave nature.

30. On the basis of Heisenberg Uncertainty principle, show that an electron cannot exist within the nucleus having a radius = $10^{-15}$ m. (h = $6.626 \times 10^{-34}$ Js)

31. Calculate the wavelength of de-Broglie waves associated with a proton of kinetic energy 500eV.

(given: $m_p = 1.67 \times 10^{-27}$kg, h = $6.626 \times 10^{-34}$Js and 1eV = $1.6 \times 10^{-9}$ J)

32. An electron has a speed of 500 $ms^{-1}$ with an uncertainty of 0.02%. What is the uncertainty in locating its position?

33. 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

34. How many degenerate atomic orbitals are there that can be designated (i) 6p (ii) 5d (iii) 6f'?

35. (i) An atomic orbital has n = 3. What are the possible values of *l* and m?

(ii) List the quantum numbers (m and *l*) of electron for 3d orbital.

(iii) Which of the following orbitals are not possible? 1s, 2p, 1p, 3f

36. (i) Write the electronic configurations of the following ions. a. $H^-$ b. $Na^+$ c. $O^{2-}$ d. $F^-$

(ii) What are the atomic numbers of elements whose outermost electrons are represented by

a. $3s^1$; b. $2p^3$; c. $3a^6$?

37. Differentiate between orbit and orbital.

38. Discuss the similarities and differences between 1s and 2s orbital.

39. Which out of $Cu^{2+}$, $Fe^{2+}$, $Cr^{3+}$ has highest para-magnetism and why? (Given At. No. Cu = 29, Fe = 26, Cr = 24)

## PART III LONG ANSWER QUESTIONS

1. (i) Calculate the wavelength of photon which will be emitted when the electron of hydrogen atom jumps from the fourth shell to the first shell. The ionization energy of hydrogen atom is $1.312 \times 10^3$ kJ $mol^{-1}$.

(ii) Which orbital is each of the following pairs is lower in energy in a multi-electron system?

a. 2s, 2p b. 3p, 3d c. 3s, 4s 4d, 5f

2. (i) Find number of photons emitted per second by a 25 watt source of monochromatic light of wavelength 6000 Å?

(ii) A certain photochemical reaction is found to require $7.61 \times 10^{-17}$ J energy per molecule. Calculate the number of photons per molecule for light of wavelength 300nm that is just sufficient to initiate the reaction.

3. (i) Calculate shortest and the longest wavelengths of the Lyman series. Given, Rydberg constant = 10967700 $m^{-1}$.

(ii) Calculate the frequency, energy and wavelength of the radiation corresponding to the spectral line of lowest frequency in Lyman series in the spectrum of hydrogen atom.

4. (i) What are the two longest wavelength lines (in nanometers) in the Lyman series of hydrogen spectrum?

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

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

6. (i) What is an emission spectrum? (ii) Explain the hydrogen spectrum.

7. Complete the following statements:

a. Two electrons in the same _____________ must have opposite spins.

b. The presence of unpaired electrons in an atom gives rise to ____________.

c. When 'l' = 3, $m_1$ may have value from ___________ to ____________.

d. The neutral fourth-period atom having a total of six 'd' electrons is ____________.

e. Orbitals with the same energy are said to be _________.

f. The electronic configuration of Sn is [Kr]– ____________.

g. The 2p orbitals of an atom have identical shapes but differ in their _____________.

h. A nodal surface is one at which the probability of finding an electron is __________.

i. Electronic configuration of Li is not $1s^3$, it is in accordance with ___________.

j. Balmer series in hydrogen spectrum is observed in the ___________region.

**8.** (i) Write outer electronic configuration of Cr atom. Why are half filled orbitals more stable?

(ii) State Heisenberg's uncertainty principle. An electron has a velocity of 50 $ms^{-1}$ accurate upto 99.99%. Calculate the uncertainty in locating its position. (Mass of electron = $9.1 \times 10^{-31}$ kg, h = $6.6 \times 10^{-34}$ Js)

## PART IV SKILL ANALYSER

**Time:** 30 min. **Max. Marks :** 15

**Directions:** (i) Attempt at questions (ii) Question 1 to 3 carry 1 mark each. (iii) Question 4 and 5 carry 2 marks each. (iv) Question 6 carry 3 marks (v) Question 7 carry 5 marks

**1.** Distinguish between a photon and quantum.

**2.** Calculate the mass and charge of one mole of electron.

**3.** How many orbitals are in subshells with 'l' equal to a. 0, b. 2?

**4.** Explain why is electronic energy negative?

**5.** Electro-magnetic radiation of wave-length 242 nm is just sufficient to ionize the sodium atom. Calculate the ionization energy of sodium in kJ/mol (h = $6.626 \times 10^{-34}$Js; c = $3 \times 10^8$ $ms^{-1}$).

**6.** Why can't we overcome the uncertainty predicted by Heisenberg's principle by building more precise devices to reduce the error in the measurement below the h/4 $\pi$ limit?

**7.** Which of the four quantum number (n, l, $m_1$, $m_s$) determine a. the energy of orbital in hydrogen atom and in multi-electron atom, b. the size of orbital, c. the shape of an orbital, d. the orientation of orbital in space, e. orientation of the spin of the electron?