2 Structure of Atom

LESSON AT A GLANCE

- **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.
- **Properties of Cathode Rays**
  (*i*) Cathode rays travel in straight line.
  (*ii*) Cathode rays start from cathode and move towards the anode.
- **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.
- **Proton:** The smallest and lightest positive ion was obtained from hydrogen and was called proton.
  Mass of proton = 1.676×10^{-27} kg
  Charge on a proton = (+) 1.602×10^{-19} C
- **Neutron:** It is a neutral particle. It was discovered by Chadwick (1932).
- **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 (*z*):** 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.
- **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.
- **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.
- **Wave Length:** It is defined as the distance between any two consecutive crests or troughs. It is represented by g 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 *υ* (*nu*) and is expressed in Hertz (Hz).

  1 Hz = 1 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 trough. 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 cm length. Evidently it will be equal to the reciprocal of the wave length. 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 increassing wave lengths or decreasing frequencies, the complete spectrum obtained is called electromagnetic spectrum.
- **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.
- **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.
- **Magnetic Orbital Quantum Number (m or m_1):** 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.
- **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.

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 one electron $= 9.1094 \times 10^{-31}$ kg

$= 9.1094 \times 10^{-28}$ g

Number of electrons in 1 g $= \dfrac{1}{9.1094 \times 10^{-28}}$

$= 1.098 \times 10^{27}$ electrons.

(*ii*) Number of electrons in one mole $= 6.022 \times 10^{23}$

Mass of 1 mole of electrons

= Mass of 1 electron × No. of electrons in one mole

$= 9.1094 \times 10^{-31}$ kg $\times 6.022 \times 10^{23}$

$= 5.486 \times 10^{-7}$ kg

Charge on one electron = 1.6022×10^{-19} C

Total charge on 1 mole of electrons

$= 1.6022 \times 10^{-19}\ C \times 6.022 \times 10^{23}$

$= \mathbf{9.648 \times 10^{4}\ C.}$

Q2. (*i*) *Calculate the total number of electrons present in 1 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 molecule of methane (CH_4) contains total of ten electrons (6 electrons from carbon and 1 electron from each of the four hydrogen atoms).

One mole of CH_4 contains 6.022×10^{23} molecules.

So total number of electrons in one mole of CH_4

$= 10 \times 6.022 \times 10^{23} = \mathbf{6.022 \times 10^{24}}$

(*ii*) (*a*) Each atom of ^{14}C contains

($A - Z = 14 - 6 = 8$) 8 neutrons

14 g of ^{14}C isotope will contain 6.022×10^{23} atoms

7 mg of ^{14}C isotope will contain

$$\frac{6.022 \times 10^{23}}{14 \times 1000} \times 7 \text{ atoms.}$$

So total number of neutrons in 7 mg of ^{14}C

$$= 8 \times \frac{6.022 \times 10^{23} \times 7}{14 \times 1000} = \mathbf{2.4088 \times 10^{21}}.$$

(*b*) Total mass of neutrons in 7 mg of ^{14}C

= Total number of neutrons × mass of each neutron

= $2.4088 \times 10^{21} \times 1.675 \times 10^{-27}$ kg

= $\mathbf{4.035 \times 10^{-6}\ kg.}$

(*iii*) (*a*) Each molecule of NH_3 contains total of 10 protons (7 of N and 3 of H)

17 g of NH_3 contains 6.022×10^{23} molecules

So 34 mg of NH_3 will contain

$$\frac{6.022 \times 10^{23}}{17 \times 1000} \times 34 \text{ molecules}$$

$= 1.2044 \times 10^{21}$ molecules

Total number of protons in 34 mg of NH_3

$= 10 \times 1.2044 \times 10^{21}$

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

(*b*) Total mass of protons in 34 mg of NH_3

= Total number of protons × mass of each proton

$= 1.2044 \times 10^{22} \times 1.673 \times 10^{-27}$ kg

$= \mathbf{2.015 \times 10^{-5}}$ **kg.**

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

$${}^{13}_{6}C,\ {}^{16}_{8}O,\ {}^{24}_{12}Mg,\ {}^{56}_{26}Fe,\ {}^{88}_{38}Sr.$$

Ans.

Nucleus	*Z*	*A*	*No. of protons = Z*	*No. of neutrons = A – Z*
${}^{13}_{6}C$	6	13	6	13–6 = 7
${}^{16}_{8}C$	8	16	8	16 – 8 = 8
${}^{24}_{12}Mg$	12	24	12	24–12 = 12
${}^{56}_{26}Fe$	26	56	26	56 – 26 = 30
${}^{88}_{38}Sr$	38	88	38	88 – 38 = 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*) ${}^{35}_{17}Cl$ (*ii*) ${}^{233}_{92}U$ (*iii*) ${}^{9}_{4}Be$

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 the yellow light.*

Ans. Wavelength of yellow light (λ) = 580 nm = 580×10^{-9} m

Velocity of light (*c*) = 3×10^{8} ms^{-1}

Since $c = \nu\lambda$

So, $$\nu = \frac{c}{\lambda} = \frac{3 \times 10^{8}\ ms^{-1}}{580 \times 10^{-9}\ m} = 5.17 \times 10^{14}\ s^{-1}$$

Wave number, $$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{580 \times 10^{-9}\ m} = \mathbf{1.72 \times 10^{6}\ m^{-1}}.$$

Q6. *Find 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*)

$$e = h\nu$$

$$h = 6.626 \times 10^{-34} \text{ J s}$$

$$\nu = 3 \times 10^{15} \text{ s}^{-1}$$

$$\therefore \quad e = 6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^{15} \text{ s}^{-1}$$

$$= \mathbf{1.9878 \times 10^{-18} \text{ J.}}$$

(*ii*)

$$e = h\nu$$

$$\text{and } \nu = \frac{c}{\lambda}$$

$$\therefore \quad e = \frac{hc}{\lambda}$$

$$h = 6.626 \times 10^{-34} \text{ J s}$$

$$c = 3 \times 10^{8} \text{ m s}^{-1}$$

$$\lambda = 50 \text{ Å} = 50 \times 10^{-10} \text{ m}$$

$$\therefore \quad e = \frac{6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^{8} \text{ ms}^{-1}}{50 \times 10^{-10} \text{ m}}$$

$$= \mathbf{3.976 \times 10^{-15} \text{ J.}}$$

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

Ans. Given that the period of the light wave

$$= 2.0 \times 10^{-10} \text{ s}$$

(*i*) Frequency (ν) of this light wave $= \dfrac{1}{\text{Period}}$

$$= \frac{1}{2 \times 10^{-10} \text{ s}} = \mathbf{5 \times 10^{9} \text{ s}^{-1}}$$

(*ii*) We know that $c = \nu\lambda$

$$\text{So, wavelength } (\lambda) = \frac{c}{\nu} = \frac{3 \times 10^{8} \text{ ms}^{-1}}{5 \times 10^{9} \text{ s}^{-1}} = \mathbf{6 \times 10^{-2} \text{ m}}$$

$$(iii) \text{ Wave number } (\bar{\nu}) = \frac{1}{\lambda} = \frac{1}{6 \times 10^{-2} \text{ m}} = \mathbf{16.67 \text{ m}^{-1}}.$$

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

Ans. Given:

$$\lambda = 4000 \text{ pm} = 4000 \times 10^{-12} \text{ m} = 4 \times 10^{-9} \text{ m}$$

Energy of 1 photon = $h\nu$

$$= \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J s})(3 \times 10^8 \text{ m s}^{-1})}{(4 \times 10^{-9} \text{ m})}$$

Number of photons that provide 1 J energy

$$= \frac{1 \text{ J}}{\text{Energy of one photon in joule}}$$

$$= \frac{(1 \text{ J})(4 \times 10^{-9} \text{ m})}{(6.626 \times 10^{-34} \text{ Js})(3 \times 10^8 \text{ m s}^{-1})}$$

$= \mathbf{2.012 \times 10^{16}}$ **photons.**

Q9. *A photon of wavelength 4×10^{-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 = 1.6020×10^{-19} J).*

Ans. Given: Work function w_0 ($h\nu_0$)= 2.13 eV

$\lambda = 4 \times 10^{-7}$ m and

(i) Energy of photon $= e = h\nu = \frac{hc}{\lambda}$

$$= \frac{(6.626 \times 10^{-34} \text{ J s})(3 \times 10^8 \text{ m s}^{-1})}{(4 \times 10^{-7} \text{ m})}$$

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

Since 1 eV = 1.602×10^{-19} J, energy of photon in electron volt,

$$\text{eV} = \frac{4.97 \times 10^{-19} \text{ J}}{1.602 \times 10^{-19} \text{ J(eV)}^{-1}} = \mathbf{3.10 \text{ eV}}.$$

(ii) Kinetic energy of emission $= \frac{1}{2}mv^2$

$= h\nu - h\nu_0 = 3.10 - 2.13 = \mathbf{0.97 \text{ eV}}$.

(iii) Velocity of photoelectron:

$\frac{1}{2}mv^2 = 0.97$ eV

$= 0.97 \text{ eV} \times 1.602 \times 10^{-19} \text{ J (eV)}^{-1}$

m = mass of electron = 9.11×10^{-31} kg

$\therefore \frac{1}{2} \times (9.11 \times 10^{-31} \text{ kg}) \times v^2$

$= 0.97 \times 1.602 \times 10^{-19}$ J

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

$$\text{or} \quad v^2 = \frac{2 \times 0.97 \times 1.602 \times 10^{-19} \text{ kg m}^2 \text{ s}^{-2}}{9.11 \times 10^{-31} \text{ kg}}$$

$= 3.41 \times 10^{11}$ m^2 s^{-1}

$$\text{or} \quad v = \sqrt{3.41 \times 10^{11} \text{ m}^2\text{s}^{-2}}$$

$$= \sqrt{34.1 \times 10^{10} \text{ m}^2\text{s}^{-2}} = \mathbf{5.84 \times 10^5 \text{ m s}^{-1}}.$$

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. Given: λ= 242 nm = 242 × 10^{-9} m

Energy per mole of photons

$$= E = N_A \times h\nu = \frac{N_A hc}{\lambda}$$

$$= \frac{(6.022 \times 10^{23} \text{ mol}^{-1}) \times (6.626 \times 10^{-34} \text{ J s}) \times (3 \times 10^8 \text{ m s}^{-1})}{(242 \times 10^{-9} \text{ m})}$$

$= \mathbf{494.65 \text{ kJ mol}^{-1}}$.

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

Ans. Given:

Energy emitted by bulb = 25 watt = 25 J s^{-1}

$\lambda = 0.57$ μm $= 0.57 \times 10^{-6}$ m

$$\text{Energy of one photon} = e = h\nu = \frac{hc}{\lambda}$$

$$= \frac{(6.626 \times 10^{-34} \text{ J s}) \times (3 \times 10^8 \text{m s}^{-1})}{(0.57 \times 10^{-6} \text{m})}$$

$= 3.48 \times 10^{-19}$ J

Energy emitted by bulb = 25 J s^{-1}

$$\therefore \quad \text{Number of photons emitted} = \frac{25 \text{ J s}^{-1})}{3.48 \times 10^{-19} \text{J}}$$

$= \mathbf{7.18 \times 10^{19} \text{ s}^{-1}}$

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. Since electrons are emitted with zero velocity the wavelength of radiation is **threshold wave length, λ_0**.

Given: $\lambda_0 = 6800 \text{ Å} = 6800 \times 10^{-10}$ m

$$\nu_0 = \frac{c}{\lambda_0} = \frac{3 \times 10^8 \text{ m s}^{-1}}{6800 \times 10^{-10} \text{ m}} = \mathbf{4.41 \times 10^{14} \text{ s}^{-1}}$$

Work function $= w_0 = h\nu_0$

$$= (6.626 \times 10^{-34} \text{J s})(4.41 \times 10^{14} \text{s}^{-1})$$

$$= \mathbf{2.92 \times 10^{-19} \text{ J}}.$$

Q13. *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?*

Ans. According to Rydberg formula,

$$\frac{1}{\lambda} = \bar{\nu} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \text{m}^{-1}$$

where R_H = Rydberg constant = $1.097 \times 10^7 \text{ m}^{-1}$

$n_1 = 2$

$n_2 = 4$

$$\bar{\nu} = 1.097 \times 10^7 \left[\frac{1}{2^2} - \frac{1}{4^2}\right] \text{m}^{-1}$$

$$= 1.097 \times 10^7 \times \frac{3}{16} \text{m}^{-1}$$

$$\lambda = \frac{1}{\bar{\nu}} = \frac{16}{1.097 \times 10^7 \times 3} \text{m} = 4.86 \times 10^{-7} \text{ m} = \mathbf{486 \text{ nm}}.$$

Q14. *How much energy is required to ionise a H atom if the 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 required to ionise a H atom if the electron is in 5 energy level

$$\Delta E = E_\infty - E_5$$

Energy for n^{th} energy level of hydrogen is given by the expression

$$E_n = -\frac{21.79 \times 10^{-19}}{n^2} \text{ J atom}^{-1}$$

and $E_\infty = 0$

So $\Delta E = E_\infty - E_5$

$$= 0 - \left(-\frac{21.79 \times 10^{-19}}{(5)^2}\right) \text{ J atom}^{-1}$$

$$= 8.716 \times 10^{-20} \text{ J atom}^{-1}$$

Ionisation energy is the energy of excitation from the orbit $n = 1$

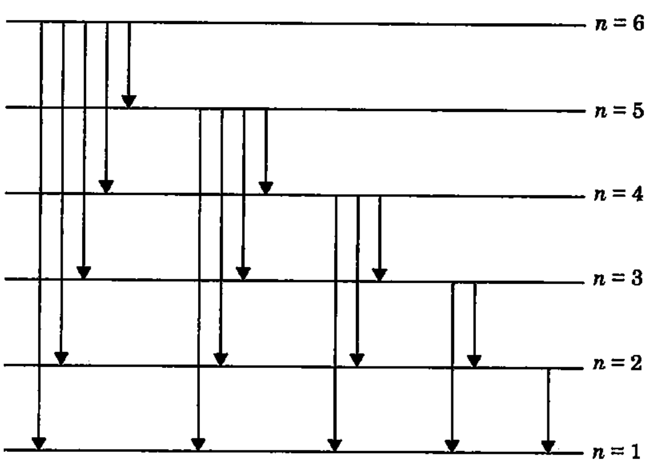
$$\Delta E = E_\infty - E_1$$

$$= 0 - \left(-\frac{21.79 \times 10^{-19}}{(1)^2}\right) \text{ J atom}^{-1}$$

$$= \mathbf{21.79 \times 10^{-19} \text{ J atom}^{-1}}.$$

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

Ans. The possible emission lines can be shown as below when the electron drops from $n = 6$.



∴ **The maximum number of emission lines is 15.**

Q16. (*i*) *The energy associated with the first orbit in the hydrogen atom is – 2.17 × 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*) Given that the energy of the first (Bohr) orbit in hydrogen atom

$$E_1 = -2.17 \times 10^{-18} \text{ J atom}^{-1}$$

Energy of the fifth Bohr orbit will be

$$E_5 = \frac{-2.17\times10^{-18}}{(5)^2} \text{ J atom}^{-1}$$

$$= \mathbf{8.68 \times 10^{-20} \text{ J atom}^{-1}}$$

(*ii*) Radius of n^{th} Bohr orbit in hydrogen atom

$$r_n = \frac{0.529 \text{ Å}(n^2)}{Z}$$

Here $n = 5$ and $z = 1$

So $r_1 = \frac{0.529 \text{ Å}\times(5)^2}{1} = 13.225 \text{ Å} = \mathbf{1.3225 \text{ nm.}}$

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

Ans. Longest wavelength transition means the transition of minimum energy and in Balmer series it will be from $n = 2$ to $n = 3$. According to Rydberg formula, wave number $\bar{\nu}$ is given as below:

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

When R_H = Rydberg constant = $1.097 \times 10^7 \text{ m}^{-1}$

$n_1 = 2$

$n_2 = 3$

Substituting these values, we obtain

$$\bar{\nu} = 1.097 \times 10^7 \text{ m}^{-1}\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= \frac{1.097\times10^7\times5}{36}\text{m}^{-1} = \mathbf{1.523 \times 10^6 \text{ m}^{-1}.}$$

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 the light emitted when the electron returns to the ground state? The ground state electron energy is -2.18×10^{-11} ergs.*

Ans. Given that the energy for the first Bohr orbit

$$E_1 = -2.18 \times 10^{-11} \text{ ergs}$$

Then energy for the fifth Bohr orbit will be

$$E_5 = \frac{E_1}{5^2} = -\frac{2.18\times10^{-11}}{25}\text{ergs}$$

$$= -0.0872 \times 10^{-11} \text{ ergs.}$$

Energy required to shift an electron from fifth Bohr orbit to first Bohr orbit is given as

$$\Delta E = E_5 - E_1$$
$$= -0.0872 \times 10^{-11} \text{ ergs} - (-2.18 \times 10^{-11} \text{ ergs})$$
$$= 2.1028 \times 10^{-11} \text{ ergs.}$$

We know that 1 erg = 10^{-7} J

So, the energy in Joules

$$= 2.1028 \times 10^{-11} \times 10^{-7} \text{ J} = \mathbf{2.1028 \times 10^{-18} \text{ J}}$$

When this electron comes back to first orbit then this much of energy will be liberated.

Wavelength of the emitted radiation is given by the relation

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

Where h = Planck's constant = 6.626×10^{-34} J s

c = velocity of light = 3×10^8 m s^{-1}

and ΔE = 2.1028×10^{-18} J

Substituting these values we obtain

$$\lambda = \frac{6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^8 \text{ m s}^{-1}}{2.1028 \times 10^{-18} \text{ J}}$$
$$= 9.453 \times 10^{-8} \text{ m} = \mathbf{945.3 \text{ Å}}.$$

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. Energy of the electron in second orbit is

$$E_2 = -\frac{2.18 \times 10^{-18}}{(2)^2} \text{ J} = -5.45 \times 10^{-19} \text{ J}$$

So, the energy required to remove an electron from this orbit is equal to 5.45×10^{-19} J. Wavelength corresponding to this energy can be calculated as

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

where, h = Planck's constant = 6.626×10^{-34} J s

c = Velocity of light = 3×10^8 m s^{-1}

E = Energy of transition = 5.45×10^{-19} J

Substituting these values, we obtain

$$\lambda = \frac{6.626\times10^{-34}\ \text{J s}\times3\times10^{8}\ \text{m s}^{-1}}{5.45\times10^{-19}\ \text{J}}$$

$= 3.647 \times 10^{-7}$ m = **3.647×10^{-5} cm.**

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

Ans. Wavelength of the moving electron is given according to de Broglie equation

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

where, h = Planck's constant

$= 6.626 \times 10^{-34}$ kg m^2 s^{-1}

m = mass of the electron = 9.1×10^{-31} kg

v = 2.05×10^7 m s^{-1}

Substituting these values, we obtain

$$\lambda = \frac{6.626\times10^{-34}\ \text{kg m}^2\ \text{s}^{-1}}{9.1\times10^{-31}\ \text{kg}\times2.05\times10^{7}\ \text{m s}^{-1}}$$

= **3.55×10^{-11} m.**

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

Ans. Kinetic energy = $\frac{1}{2}mv^2$

where m = mass of the electron = 9.1×10^{-31} kg

v = velocity of electron =?

K.E. = 3.0×10^{-25} J = 3.0×10^{-25} kg m^2 s^{-2}

From the formula, we get $v = \left(\frac{2\times\text{K.E.}}{m}\right)^{1/2}$

Substituting the values, we get

$$v = \left(\frac{2\times3.0\times10^{-25}\ \text{kg m}^2\ \text{s}^{-2}}{9.1\times10^{-31}\ \text{kg}}\right)^{1/2} = 812\ \text{m s}^{-1}$$

According to de Broglie equation $\lambda = \frac{h}{mv}$

where, h = Planck's constant

$= 6.626 \times 10^{-34}$ kg m^2 s^{-1}

m = mass of electron = 9.1×10^{-31} kg

v = velocity of electron = 812 m s^{-1}

So, wavelength (λ)

$$= \frac{6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{9.1 \times 10^{-31} \text{ kg} \times 812 \text{ m s}^{-1}}$$

$$= 8.967 \times 10^{-7} \text{ m} = \mathbf{8967\ Å}.$$

Q22. *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.

Ans. The number of electrons present in each species are shown against their symbol

Na^+	–	10
K^+	–	18
Mg^{2+}	–	10
Ca^{2+}	–	18
S^{2-}	–	18
Ar	–	18

So, Na^+ and Mg^{2+} are isoelectronic. Similarly K^+, Ca^{2+}, S^{2-} and Ar are all isoelectronic.

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 elements whose outermost electrons are represented by (a) $3s^1$; (b) $2p^3$ and (c) $3p^5$?*

(*iii*) *Which atoms are indicated by the following configuration?*

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

Ans. (*i*)

	Ion	Electronic configuration
(*a*)	H^-	$1s^2$
(*b*)	Na^+	$1s^2\ 2s^2\ 2p^6$
(*c*)	O^{2-}	$1s^2\ 2s^2\ 2p^6$
(*d*)	F^-	$1s^2\ 2s^2\ 2p^6$

(*ii*)

	Outer electronic configuration	Atomic number of the element
(*a*)	$3s^1$	11
(*b*)	$2p^3$	7
(*c*)	$3p^5$	17

(*iii*)

Electronic configuration	Name of the element
(*a*) [He] $2s^1$	Lithium (Li)
(*b*) [Ne] $3s^2\ 3p^3$	Phosphorus (P)
(*c*) [Ar] $4s^2\ 3d^1$	Scandium (Sc)

Q24. *What is the lowest value of n that allows g-orbitals to exist?*

Ans. *g*-orbitals can exist for the principal quantum number, n = 5 and onwards.

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

Ans. Possible values for different quantum numbers for an electron in 3*d*– orbitals are

$$n = 3;\ l = 2$$

$$m_l = -2, -1, 0, +1, +2.$$

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. (*i*) Number of proton = 29 (because there are 29 electrons)

(*ii*) 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.

Species	Number of electrons
H_2^+	1
H_2	2
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 electrons for 3d– orbital.*

(*iii*) *Which of the following orbitals are possible?*

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

Ans. (*i*) For n = 3 the possible values of l and m_l are:

l	m_l
0	0
1	– 1, 0, + 1
2	– 2, – 1, 0, + 1, + 2

(*ii*) Possible values of l and m_l for 3*d* electrons are:

$l = 2$

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

(*iii*) The possible orbitals are 2*s* and 2*p*.

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

Ans.

Quantum numbers	Notation for the orbital
(*a*) $n = 1, l = 0$	1*s*
(*b*) $n = 3, l = 1$	3*p*
(*c*) $n = 4, l = 2$	4*d*
(*d*) $n = 4, l = 3$	4*f*

Q30. *Explain, giving reasons, which of the following sets of quantum numbers are not possible.*

(*a*)	$n = 0,$	$l = 0,$	$m_l = 0,$	$m_s = +1/2$
(*b*)	$n = 1,$	$l = 0,$	$m_l = 0,$	$m_s = -1/2$
(*c*)	$n = 1,$	$l = 1,$	$m_l = 0,$	$m_s = +1/2$
(*d*)	$n = 2,$	$l = 1,$	$m_l = 0,$	$m_s = -1/2$
(*e*)	$n = 3,$	$l = 3,$	$m_l = -3,$	$m_s = +1/2$
(*f*)	$n = 3,$	$l = 1,$	$m_l = 0,$	$m_s = +1/2$

Ans. (*a*) This set of quantum number is not possible because $n = 0$ is not permitted. Values of n are 1, 2, 3 ... etc.

(*b*) Possible.

(*c*) This set of quantum number is not possible because for $n = 1$, the only permitted value of l is 0.

(*d*) Possible.

(*e*) This set of quantum numbers is not possible because for $n = 3$, the permitted values of l are 0, 1 and 2; $l = 3$ is not possible for $n = 3$.

(*f*) 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*) 16 electrons

There can be maximum of 32 electrons in different orbitals with $n = 4$, *i.e.*, 4*s*, 4*p*, 4*d* and 4*f*. Out of these, one-half, *i.e.*, 16 can have $+\frac{1}{2}$ value for m_s and the other one-half, *i.e.*, 16 can have $-\frac{1}{2}$ value.

(*b*) It represents 3*s* orbital so there can be only 2 electrons with $n = 3, l = 0$.

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 de Broglie equation, $\lambda = \frac{h}{mv}$

According to Bohr's theory,

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

$\therefore$ Circumference of Bohr orbit $= 2\pi r = \frac{nh}{mv}$

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

Circumference $= 2\pi r = n\lambda$

= Integral multiple of de Broglie wavelength of electron.

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

Ans. When wavelength of a particular line of hydrogen spectrum and that of the given Balmer transition of He^+ are the same, their energies would also be the same.

The energy of transition of electron in a one electron hydrogen like particle

$$= \Delta E = R_H Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where Z = atomic number of the element

For $n = 4$ to $n = 2$, transition of He^{2+} $(Z = 2)$

$$\Delta E = R_H \times 2^2\left(\frac{1}{2^2} - \frac{1}{4^2}\right)$$

$$= R_H \times 4 \times \frac{16-4}{4 \times 16} = \frac{3}{4}R_H$$

For hydrogen spectrum $(Z = 1)$

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

This would be true if $n_1 = 1$ and $n_2 = 2$, hence for the transition $n = 2$ to $n = 1$.

Q34. *Calculate the energy required for the process*

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

The ionisation energy for the H atom in the ground state is 2.18×10^{-18} *J atom*$^{-1}$.

Ans. Energy of an electron in hydrogen like particle,

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

For the process $He^+ (g) \rightarrow He^{2+} (g) + e^-$ the transition is from $n = 1$ to $n = \infty$ and $Z = 2$ for He.

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

$$= -2.18 \times 10^{-18} J \times 4(0 - 1)$$

$$= +2.18 \times 10^{-18} \times 4 J = \mathbf{8.72 \times 10^{-18} J.}$$

Q35. *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 length of scale of length 20 cm long.*

Ans. Diameter of one carbon atom

$$= 0.15 \text{ nm} = 0.15 \times 10^{-9} \text{ m}$$

Total length of all the carbon atoms

$$= 20 \text{ cm} = 20 \times 10^{-2} \text{ m}$$

$$\therefore \quad \text{No. of carbon atoms} = \frac{\text{Total length}}{\text{Diameter of one atom}}$$

$$= \frac{20 \times 10^{-2} \text{ m}}{0.15 \times 10^{-9} \text{ m}} = \mathbf{1.33 \times 10^9.}$$

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

Ans. Total length = sum of diameters of all the carbon atoms

$$= 2.4 \text{ cm} = 2.4 \times 10^{-2} \text{ m}$$

No. of carbon atoms = 2×10^8

$$\therefore \text{Diameter of one carbon atom} = \frac{\text{Total length}}{\text{No. of atoms}}$$

$$= \frac{2.4 \times 10^{-2} \text{ m}}{2 \times 10^8} = 1.2 \times 10^{-10} \text{ m}$$

Radius of one carbon atom $= \frac{1}{2} \times 1.2 \times 10^{-10}$

$= 0.60 \times 10^{-10}$ m

$= 0.060 \times 10^{-9}$ m = **0.060 nm.**

Q37. *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.*

Ans. Given:

Diameter of zinc atom = 2.6 Å = 2.6×10^{-10} m

(*a*) Radius of zinc atom $= \frac{2.6 \times 10^{-10} \text{ m}}{2}$

$= 1.3 \times 10^{-10}$ m

$= 130 \times 10^{-12}$ m = 130 pm

(*b*) Total length = 1.6 cm = 1.6×10^{-2} m

Diameter of one atom = 2.6×10^{-10} m

$\therefore$ No. of zinc atoms $= \frac{1.6 \times 10^{-2} \text{ m}}{2.6 \times 10^{-10} \text{ m}}$ = **6.154×10^7.**

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

Ans. Given: Total (negative) charge on the particle

$= 2.5 \times 10^{-16}$ C

Charge (negative) on one electron

$= 1.6022 \times 10^{-19}$ C

$\therefore$ No. of electrons present

$= \frac{\text{Total charge}}{\text{Charge on one electron}}$

$= \frac{2.5 \times 10^{-16} \text{ C}}{1.6022 \times 10^{-19} \text{ C}}$ = **1560.**

Q39. *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 -1.282×10^{-18} C, calculate the number of electrons present on it.*

Ans. Given:

Total charge on the oil drop = -1.282×10^{-18} C

Charge on one electron = -1.6022×10^{-19} C

∴ No. of electrons present

$$= \frac{\text{Total charge}}{\text{Charge on one electron}}$$

$$= \frac{-1.282 \times 10^{-18}\ \text{C}}{-1.6022 \times 10^{-19}\ \text{C}} = 8.$$

Q40. *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?*

Ans. Charge present on the nuclei of heavy metals is large enough to deflect very few α-particles back or through large angles and some of these through small angles. The charge on the nuclei of light metals like aluminium is small and would not be able to deflect α-particles back or by large angles. Only some of them would be deflected by small angles.

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. Atomic number of each element is fixed and even if it is not written it has the same value. Thus, $^{79}_{35}\text{Br}$ and ^{79}Br mean the same nucleus.

However, different isotopes of a given element have different mass numbers, it must be written otherwise the information is incomplete. Also, written this way, the position of atomic number and mass number have been interchanged which is in violation of the convention.

Q42. *An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol. Also, written this way, the positions of atomic number and mass number have been interchanged which is in violation of the convention.*

Ans. Given: Mass number, $A = p + n = 81$

Let the number of protons, $p = x$

∴ The number of neutrons, n = 31.7% more than protons

$$= x + \frac{31.7}{100}x = 1.317x$$

$$\therefore \quad A = p + n = x + 1.317x = 81$$

or $2.317x = 81$

$$x = \frac{81}{2.317} = 35$$

$\therefore$ No. of protons = Z = 35

The element is Br and the symbol is $^{81}_{35}Br$.

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 number of electrons in the ion = x

$\therefore$ Number of neutrons

= 11.1% more than the no. of electrons

$$n = x + \frac{11.1}{100}x = 1.111x$$

Since the ion carries one unit of negative charge, the number of electrons in the neutral atom = $x - 1$

$\therefore$ No. of protons $= x - 1$

Given $A = p + n = 37$

or $(x - 1) + 1.111x = 37$

$$2.111x - 1 = 37$$

$$2.111x = 38$$

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

$\therefore$ No. of protons, $p = Z = 18 - 1 = 17$

The element is chlorine and the symbol is $^{37}_{17}Cl^{-1}$.

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

Ans. Let the number of electrons = x

No. of neutrons, n = 30.4% more than the no. of electrons

$$= x + \frac{30.4}{100}x = 1.304x$$

Since the ion carries 3 units of positive charge, the no. of electrons in neutral atom = $x + 3$

$\therefore$ No. of protons, $p = x + 3$

Given: $A = p + n = 56$

$$x + 3 + 1.304x = 56$$

$$2.304x = 56 - 3 = 53$$

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

No. of protons, p = Atomic No. $Z = x + 3$
$= 23 + 3 = 26$

The element is Fe and the symbol is $^{56}_{26}Fe^{3+}$.

Q45. *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.*

Ans. FM < microwave < amber light < X-rays < cosmic rays.

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

Ans. Given:

$$\lambda = 337.1 \text{ nm} = 337.1 \times 10^{-9} \text{ m}$$

No. of photons emitted (per second)

$$N = 5.6 \times 10^{24} \text{ s}^{-1}$$

Power = Energy emitted (per second)
= No. of photons emitted (per second) × Energy of one photon

$$= \frac{N \times h \times c}{\lambda}$$

$$= \frac{(5.6 \times 10^{24} \text{s}^{-1}) \times (6.626 \times 10^{-34} \text{J s}) \times (3 \times 10^{8} \text{m s}^{-1})}{(337.1 \times 10^{-9} \text{m})}$$

$= 3.3 \times 10^6$ J (per second) = **3.3×10^6 watt.**

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

Ans. Given: $\lambda = 616 \text{ nm} = 616 \times 10^{-9}$ m

(*a*) Frequency, $\nu = \frac{c}{\lambda} = \frac{3 \times 10^8 \text{ ms}^{-1}}{616 \times 10^{-9} \text{ m}} = \mathbf{4.87 \times 10^{14}\ s^{-1}}$

(*b*) Velocity of radiation, $c = \mathbf{3 \times 10^8\ m\ s^{-1}}$

Distance travelled in 30 s = Time × Velocity
$= (30 \text{ s}) \times (3 \times 10^8 \text{ m s}^{-1}) = \mathbf{9 \times 10^9\ m}$

(*c*) Energy of quantum, $e = h\nu$

$= (6.626 \times 10^{-34}\ \text{J s}) \times (4.87 \times 10^{14}\ \text{s}^{-1})$

$= 32.27 \times 10^{-20}\ \text{J}$

(*d*) Total energy produced = 2 J

No. of quanta emitted

$$= \frac{\text{Total energy}}{\text{Energy of one quantum}}$$

$$= \frac{2\ \text{J}}{32.27 \times 10^{-20}\ \text{J}} = \mathbf{6.2 \times 10^{18}}.$$

Q48. *In astronomical observations, signals observed from the distant stars 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. Given: $\lambda = 600\ \text{nm} = 600 \times 10^{-9}\ \text{m}$

Total energy received $= 3.15 \times 10^{-8}\ \text{J}$

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

$$= \frac{(6.626 \times 10^{-34}\ \text{J s}) \times (3 \times 10^{8}\ \text{m s}^{-1})}{(600 \times 10^{-9}\ \text{m})}$$

$= 3.313 \times 10^{-19}\ \text{J}$

No. of photons received $= \frac{3.15 \times 10^{-8}\ \text{J}}{3.313 \times 10^{-19}\ \text{J}} = 9.51 = \mathbf{10}$

(since no. of photons can't be fraction).

Q49. *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 2 ns and the number of photons emitted during the pulse source is 2.5×10^{15}, calculate the energy of the source.*

Ans. No. of photons emitted $= 2.5 \times 10^{15}$

Duration of pulse radiation $= 2\ ns = 2 \times 10^{-9}\ \text{s}$

$\therefore$ Frequency of radiation

= No. of pulses per second

$$= \frac{1}{\text{Duration of one pulse}}$$

$$= \frac{1}{2 \times 10^{-9}\ \text{s}} = 0.5 \times 10^{9}\ \text{s}^{-1}$$

Energy of the source

= No. of photons emitted × Energy of one photon

= $Nh\nu$

= $(2.5 \times 10^{15}) \times (6.626 \times 10^{-34}$ J s$)$
$\times (0.5 \times 10^{9}\ s^{-1})$

= **8.28×10^{-10} J.**

Q50. *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.*

Ans. Frequency, $\nu = \dfrac{c}{\lambda}$

Frequency of radiations:

(*i*) λ_1 589 nm = 589×10^{-9} m

$$\therefore \quad \nu_1 = \frac{3 \times 10^{8}\ m\,s^{-1}}{589 \times 10^{-9}\ m} = 5.0934 \times 10^{14}\ s^{-1}$$

(*ii*) λ_2 = 589.6 nm = 589.6×10^{-9} m

$$\therefore \quad \nu_2 = \frac{3 \times 10^{8}\ m\,s^{-1}}{589.6 \times 10^{-9}\ m} = 5.0882 \times 10^{14}\ s^{-1}$$

Energy Difference

Energy = $\lambda\nu$

$e_1 = \lambda\nu_1 = 6.626 \times 10^{-34}$ J s $\times 5.0934 \times 10^{14}\ s^{-1}$

$e_1 = \lambda\nu_2 = 6.626 \times 10^{-34}$ J s $\times 5.0882 \times 10^{14}\ s^{-1}$

$\Delta e = (e_1 - e_2)$

$= 6.626 \times 10^{-34}$ J $\times (5.0934 - 5.0882)\ 10^{14}\ s^{-1}$

$= 6.626 \times 10^{-34}$ J $\times 0.0052 \times 10^{14}\ s^{-1}$

= **3.45×10^{-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 with a wavelength 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.*

Ans. Given:

Work function, $w_0 = h\nu_0$ = 1.9 eV = $1.9 \times 1.602 \times 10^{-19}$ J

(*a*) & (*b*)

Threshold frequency

$$\nu_0 = \frac{h\nu_0}{h} = \frac{1.9 \times 1.602 \times 10^{-19}\ J}{6.626 \times 10^{-34}\ J\,s} = \mathbf{4.59 \times 10^{14}\ s^{-1}}$$

Threshold wavelength, $\lambda_0 = \dfrac{c}{\nu_0}$

$$= \frac{3 \times 10^8 \text{ m s}^{-1}}{4.59 \times 10^{14} \text{ s}^{-1}} = 6.54 \times 10^{-7} \text{ m}$$

$= 654 \times 10^{-9}$ m = **654 nm**

Wavelength of radiation used $= \lambda$

$= 500 \text{ nm} = 500 \times 10^{-9}$ m

Frequency of radiation used

$$= \nu = \frac{c}{\lambda} = \frac{3 \times 10^8 \text{ m s}^{-1}}{500 \times 10^{-9} \text{ m}} = \mathbf{6 \times 10^{14} \text{ s}^{-1}}$$

Kinetic energy of effected electron $= \dfrac{1}{2}mv^2$

$= h\nu - h\nu_0 = h(\nu - \nu_0)$

$= (6.626 \times 10^{-34} \text{ J s}) \times (6 \times 10^{14} - 4.59 \times 10^{14}) \text{ s}^{-1}$

$= (6.626 \times 10^{-34} \text{ J s}) \times (1.41 \times 10^{14} \text{ s}^{-1})$

$= 9.34 \times 10^{-20}$ J

Kinetic energy $= \dfrac{1}{2}mv^2 = 9.34 \times 10^{-20} \text{ kg m}^2 \text{ s}^{-2}$

$\because \quad m = 9.11 \times 10^{-31}$ kg

$$\therefore \quad v^2 = \frac{2 \times 9.34 \times 10^{-20} \text{ kg m}^2\text{s}^{-2}}{9.11 \times 10^{-31}\text{kg}}$$

$= 2.055 \times 10^{11} \text{ m}^2 \text{ s}^{-2} = 20.55 \times 10^{10} \text{ m}^2 \text{ s}^{-2}$

$v = \sqrt{20.55 \times 10^{10}\text{m}^2\text{s}^{-2}} = \mathbf{4.53 \times 10^5 \text{ m s}^{-1}}$.

Q52. *Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength and, (b) Planck's constant.*

λ (nm)	*500*	*450*	*400*
$v \times 10^{-5}$ (cm s^{-1})	*2.55*	*4.35*	*5.35*

Ans. Let the threshold wavelength be λ_0 nm

$= \lambda_0 \times 10^{-9}$ m

Kinetic energy $= \dfrac{1}{2}mv^2 = h\nu - h\nu_0 = hc\left(\dfrac{1}{\lambda} - \dfrac{1}{\lambda_0}\right)$

Since all wavelengths are in nm,

$$\frac{hc}{10^{-9}}\left(\frac{1}{\lambda} - \frac{1}{\lambda_0}\right) = \frac{1}{2}mv^2$$

Substituting the given data sets,

$$\frac{hc}{10^{-9}}\left(\frac{1}{500}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m(2.55\times10^{6})^2 \quad ...(i)$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{450}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m(4.35\times10^{6})^2 \quad ...(ii)$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{400}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m(5.35\times10^{6})^2 \quad ...(iii)$$

λ_0 can be calculated from (*i*) and (*ii*) relations.

Dividing *eqn.* (*ii*) by *eqn.* (*i*),

$$\left(\frac{\lambda_0-450}{450\lambda_0}\right)\times\left(\frac{500\lambda_0}{\lambda_0-500}\right)=\left(\frac{4.35}{2.55}\right)^2$$

$$\left(\frac{\lambda_0-450}{\lambda_0-500}\right)=\frac{450}{500}\left(\frac{4.35}{2.55}\right)^2=2.619$$

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

$$=2.619\lambda_0-1309.5$$

$$1309.5-450=2.619\lambda_0-\lambda_0$$

$$859.5=1.619\lambda_0$$

or $\lambda_0=531$ nm

Substituting this in equation (*iii*)

$$\frac{h\times3\times10^{8}\ \text{m s}^{-1}}{10^{-9}\ \text{m}}\left(\frac{1}{400}-\frac{1}{531}\right)$$

$$=\frac{1}{2}\times(9.11\times10^{-31}\ \text{kg})(5.35\times10^{6}\ \text{m s}^{-1})^2$$

$$\frac{h\times3\times10^{8}\text{m s}^{-1}}{10^{-9}\ \text{m}}\left(\frac{531-400}{400\times531}\right)$$

$$=\frac{1}{2}\times(9.11\times10^{-31}\ \text{kg})\times(5.35\times10^{6}\ \text{m s}^{-1})^2$$

$$h\times\frac{(3\times10^{8}\text{m s}^{-1})\times131}{10^{-9}\text{m}\times400\times531}$$

$$=\frac{1}{2}\times(9.11\times10^{-31}\ \text{kg})\times(28.62\times10^{12}\ \text{m}^2\text{s}^{-2})$$

$$h = \frac{(9.11\times10^{-31}kg)\times(28.62\times10^{12}m^2s^{-2})\times10^{-9}m\times400\times531}{2\times(3\times10^8 m\, s^{-1})\times131}$$

$= 7.046 \times 10^{-32}$

Thus, $\lambda_0 = 531$ nm and h = **7.046×10^{-32} J s.**

Q53. *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.*

Ans. Energy of incident radiation

= Work function + Kinetic energy of photoelectron

Work function = Energy of incident radiation

– kinetic energy of photo electron

Energy of incident radiation:

Given: $\lambda = 256.7$ nm $= 256.7 \times 10^{-9}$ m

$$e = h\nu = \frac{hc}{\lambda}$$

$$= \frac{(6.626\times10^{-34}\text{ J s})\times(3\times10^{8}\text{ m})}{256.7\times10^{-9}\text{m}}$$

$= 7.74 \times 10^{-19}$ J

$$= \frac{7.74\times10^{-19}\text{J}}{1.602\times10^{-19}\text{J(eV)}^{-1}} = \mathbf{4.83\ eV}$$

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

Kinetic energy of photoelectron: The applied potential of 0.35 V is sufficient to stop the photoelectrons. Thus, the energy imparted to each electron with charge e = 1e × 0.35 V =0.35 eV is equal to the kinetic energy.

∴ Work function = 4.83 – 0.35 = **4.48 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 m s^{-1}, calculate the energy with which it is bound to the nucleus.*

Ans. Energy of incident radiation

= Binding energy of electron

+ Kinetic energy of ejected electron

∴ Binding energy of electron

= Energy of incident radiation

– Kinetic energy of electron

Energy of incident radiation:

Given: $\lambda = 150 \text{ pm} = 150 \times 10^{-12} \text{ m}$

$$e = h\nu = \frac{hc}{\lambda}$$

$$= \frac{(6.626 \times 10^{-34} \text{ J s}) \times (3 \times 10^{8} \text{ m})}{150 \times 10^{-12} \text{m}}$$

$$= \mathbf{13.25 \times 10^{-16} \text{ J}}$$

Kinetic energy of ejected electron:

Given: Velocity $v = 1.5 \times 10^{7} \text{ m s}^{-1}$

$$\text{Kinetic energy} = \frac{1}{2}mv^2$$

$$= \frac{1}{2}(9.11 \times 10^{-31} \text{ kg})(1.5 \times 10^{7} \text{ m s}^{-1})^2$$

$$= 1.025 \times 10^{-16} \text{ J}$$

$\therefore$ Binding energy $= (13.25 \times 10^{-16} - 1.025 \times 10^{-16}) \text{ J}$

$$= 12.225 \times 10^{-16} \text{ J.}$$

$$= \frac{12.225 \times 10^{-16} \text{J}}{1.602 \times 10^{-19} \text{ J (eV)}^{-1}} = \mathbf{7.63 \times 10^{3} \text{ eV.}}$$

Q55. *Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as* $v = 3.29 \times 10^{15} \text{ (Hz) } [1/3^2 - 1/n^2]$

Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

Ans. Given:

λ of radiation causing transition = 1285 nm

$$= 1285 \times 10^{-9} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{3 \times 10^{8} \text{ m s}^{-1}}{1285 \times 10^{-9} \text{ m}}$$

From the given relation,

$$\nu = \frac{3 \times 10^{8}}{1285 \times 10^{-9}} \text{s}^{-1}$$

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

$$\frac{1}{9} - \frac{1}{n^2} = \frac{3 \times 10^{8} \text{s}^{-1}}{1285 \times 10^{-9} \times 3.29 \times 10^{15} \text{s}^{-1}}$$

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

$$\frac{1}{n^2} = 0.111 - 0.071 = 0.04 = \frac{1}{25}$$

$$\therefore \quad n^2 = 25 \text{ and } n = 5.$$

The wavelength of incident radiation 1285 nm lies in the infrared region.

Q56. *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.*

Ans. According to Bohr's theory, the radius of n^{th} orbit of a hydrogen like particle with atomic number Z is given by

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

Radius of initial higher energy orbit ($r = r_2$, $n = n_2$)

$$r_2 = 1.3225 \text{ nm} = 1322.5 \text{ pm} = \frac{52.9\, n_2^2}{Z}$$

Radius of final lower energy orbit ($r = r_1$; $n = n_1$)

$$r_1 = 211.6 \text{ pm} = \frac{52.9\, n_1^2}{Z}$$

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

$$\text{or} \quad 6.25 = \left(\frac{n_2}{n_1}\right)^2 \quad \text{or} \quad \frac{n_2}{n_1} = \sqrt{6.25} = 2.5.$$

This ratio is possible if $n_2 = 5$ and $n_1 = 2$. Since the n (n_1) for lower energy orbit is 2 and the transition is from 5^{th} to 2^{nd} orbit, it belongs to Balmer Series. The wave number $\bar{v}$ of emitted radiation is given by

$$\bar{v} = 1.097 \times 10^7 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \text{m}^{-1}$$

$$n_1 = 2 \text{ and } n_2 = 5$$

$$\bar{v} = 1.097 \times 10^7 \left[\frac{1}{2^2} - \frac{1}{5^2}\right] \text{m}^{-1}$$

$$= 1.097 \times 10^7 \left(\frac{21}{100}\right) m^{-1}$$

$$\lambda = \frac{1}{\nu} = \frac{100}{1.097 \times 10^7 \times 21} m$$

$$= 4.34 \times 10^{-7}\ m = 434 \times 10^{-9}\ m = \mathbf{434\ nm.}$$

This wavelength lies in the visible region.

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 the microscope is $1.6 \times 10^6\ ms^{-1}$, calculate de Broglie wavelength associated with this electron.*

Ans. Given velocity, v of electron = $1.6 \times 10^6\ m\ s^{-1}$

de Broglie wavelength $\lambda = \frac{h}{mv}$

$$= \frac{6.626 \times 10^{-34} kg\ m^2\ s^{-1}}{(9.11 \times 10^{-31}\ kg)(1.6 \times 10^6\ m\ s^{-1})}$$

$$= 4.55 \times 10^{-10}\ m = \mathbf{455\ pm.}$$

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. Given:

de Broglie wavelength of neutron = 800 pm

$= 800 \times 10^{-12}$ m

Mass of neutron = 1.675×10^{-27} kg

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

or $$v = \frac{h}{m\lambda}$$

$$= \frac{6.626 \times 10^{-34}\ kg\ m^2\ s^{-1}}{(1.675 \times 10^{-27}\ kg)(800 \times 10^{-12}\ m)}$$

$(\because 1\ J = 1\ kg\ m^2\ s$

$$= \mathbf{4.94 \times 10^4\ m\ s^{-1}.}$$

Q59. *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.*

Ans. Given: Velocity of electron = $v = 2.19 \times 10^6$ m s^{-1}

de Broglie wavelength = $\lambda = \dfrac{h}{mv}$

$$= \frac{6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}}{(9.11 \times 10^{-31} \text{ kg}) \times (2.19 \times 10^6 \text{ m s}^{-1})}$$

$= 3.32 \times 10^{-10}$ m = **332 pm.**

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

Ans. Given: Velocity of hockey ball = Velocity of proton
$= 4.37 \times 10^5$ m s^{-1}

Mass of hockey ball = 0.1 kg

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

$= \mathbf{1.516 \times 10^{-28}}$ **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 $h/4\pi_m \times 0.05$ nm, is there any problem in defining this value?*

Ans. **Uncertainty in momentum**

Given: Uncertainty in position = 0.002 nm
$= 0.002 \times 10^{-9}$ m $= 2 \times 10^{-12}$ m

From Heisenberg's uncertainty principle,

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

or $$\Delta p \geq \frac{h}{4\pi.\Delta x} = \frac{6.626 \times 10^{-34} \text{kg m}^2\text{s}^{-1}}{4 \times 3.14 \times 2 \times 10^{-12} \text{m}}$$

$= 2.638 \times 10^{-23}$ kg m s^{-1}

∴ $\Delta p \geq \mathbf{2.638 \times 10^{-23}}$ **kg m s^{-1}**

Actual momentum

Given $$p = \frac{h}{4\pi \times 0.05 \text{ nm}} = \frac{6.626 \times 10^{-34} \text{kg m}^2\text{s}^{-1}}{4 \times 3.14 \times 0.05 \times 10^{-9} \text{m}}$$

$= 1.055 \times 10^{-24}$ kg m s^{-1}.

Since the uncertainty in momentum 2.638×10^{-23} is (26 times) greater than the actual momentum 1.055×10^{-24}, this momentum cannot be defined and the value is meaningless.

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

1. $n = 4, l = 2, m_l = -2, m_s = -1/2$
2. $n = 3, l = 2, m_l = 1, m_s = +1/2$
3. $n = 4, l = 1, m_l = 0, m_s = +1/2$
4. $n = 3, l = 2, m_l = -2, m_s = -1/2$
5. $n = 3, l = 1, m_l = -1, m_s = +1/2$
6. $n = 4, l = 1, m_l = 0, m_s = +1/2$

Ans. The energy of a multielectron atom is determined by $(n + l)$ values and n value if $(n + l)$ values are equal. Let us calculate these:

S.No.	n	l	$n + l$	Orbital
(*i*)	4	2	6	4*d*
(*ii*)	3	2	5	3*d*
(*iii*)	4	1	5	4*p*
(*iv*)	3	2	5	3*d*
(*v*)	3	1	4	3*p*
(*vi*)	4	1	5	4*p*

The order of increasing energy of these orbitals is $3p < 3d < 4p < 4d$.

The order of increasing energy of electrons (S. No.) is (*v*) < (*ii*) = (*iv*) < (*iii*) = (*vi*) < (*i*).

The combinations (*a*) (*ii*) and (*iv*) and (*b*) (*iii*) and (*vi*) have same energies.

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 electrons experiences the lowest effective nuclear charge?*

Ans. 4*p* electrons being farthest (largest value of *n*) would experience the lowest effective charge.

Q64. *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.*

Ans. (*i*) *Out of 2s and 3s:* 2*s* since it is closer to the nucleus (lower value of *n*)

(*ii*) *Out of 4d and 4f:* 4*d* because here *n* is same (4) for both but *l* is smaller for *d* orbitals.

(*iii*) *Out of 3d and 3p:* 3*p* (same *n* but lower *l*)

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. In both the atoms, the electron is present in the 3*p* orbital with same values of *n* and *l* but since silicon's atomic no. 14 is greater than that of aluminium (13), it has greater nuclear charge. Hence, the 3*p* electron of Si will experience more effective nuclear charge.

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

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

Ans. The number of unpaired electrons can be found from the electronic configurations of these elements:

(*a*) $P(Z = 15) = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^{\ 1}\ 3p_y^{\ 1}\ 3p_z^{\ 1}$

No. of unpaired electrons = 3

(*b*) $Si(Z = 14) = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1$

No. of unpaired electrons = 2

(*c*) $Cr(Z = 24) = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

No. of unpaired electrons = 6

(*d*) $Fe(Z = 26) = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$

No. of unpaired electrons = 4

(*e*) $Kr(Z = 36) = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

No. of unpaired electrons = 0

Same conclusion can be drawn for Kr even without writing the electronic configuration. Being noble gas, all orbitals are completely filled therefore, it has no unpaired electron.

Q67. *(a) How many subshells 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 $n = 4$, $l = 0, 1, 2, 3$ corresponding to 4*s*, 4*p*, 4*d*, 4*f*. Thus, four subshells are associated with $n = 4$ sub-shells.

(*b*) The number of orbitals in $= n^2 = 4^2 = 16$.

Each orbital will have one electron with $m_s = -\frac{1}{2}$.

Thus, there would be 16 electrons with

$$m_s = -\frac{1}{2} \text{ for } n = 4.$$

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