CHAPTER

2

Structure of Atom

CLASS 11TH

NCERT EXERCISE AND SOLUTIONS - CHEMISTRY

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

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

ANSWER:-

(i) Mass of one electron = 9.10939×10^{-31} kg

$\therefore$ Number of electrons that weigh 9.10939×10^{-31} kg = 1

Number of electrons that will weigh 1g = $(1 \times 10^{-3}$ kg)

$$= \frac{1}{9.10939 \times 10^{-31}} \times \left(1 \times 10^{-3} \text{kg}\right)$$

$$= 0.1098 \times 10^{-3+31}$$

$$= 0.1098 \times 10^{28}$$

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

(ii) Mass of one mole of electron = $(6.022 \times 10^{23}) \times (9.10939 \times 10^{-31}$ kg)

= 5.48×10^{-7} kg

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

Charge on one mole of electron = $(1.6022 \times 10^{-19}$ C)(6.022×10^{23})

= 9.65×10^{4} C

Q. 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 14C. (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 NH3 at STP. Will the answer change if the temperature and pressure are changed?

ANSWER:-

(i) Number of electrons present in 1 molecule of methane (CH_4)

$\{1(6) + 4(1)\} = 10$

Number of electrons present in 1 mole i.e., 6.023×10^{23} molecules of methane

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

(ii) Number of atoms of ^{14}C in 1 mole = 6.023×10^{23}

Since 1 atom of contains (14^{-6}) i.e., 8 neutrons, the number of neutrons in 14 g of ^{14} C is $(6.023 \times 10^{23}) \times 8$.

Pr, 14 g of ^{14}C contains $(6.022 \times 10^{23} \times 8)$ neutrons.

Number of neutrons in 7 mg

$$= \frac{6.022\times10^{23}\times8\times7\text{mg}}{1400\text{mg}}$$

$$= 2.4092\times10^{21}$$

(a) 1mole of $NH_3 = \{1(14) + 3(1)\}$g of NH_3

$= 17$g of NH_3

$= 6.022 \times 10^{23}$ molecules of NH_3

Total number of protons present in 1 molecule of NH_3

$= \{1(7) + 3(1)\}$

$= 10$

Number of protons in 6.023×10^{23} molecules of NH_3

$= (6.023 \times 10_{23})(10)$

$= 6.023 \times 10^{24}$

⇒ 17 Of NH_3 contains (6.023×10^{23}) protons.

Number of protons in 34 mg Of NH_3

$$= \frac{6.022\times10^{23}\times34\text{ mg}}{17000\text{ mg}}$$

$= 1.2046 \times 10^{21}$

(b) Mass of one proton = 1.67493×10^{-27} kg

Total mass of protons in 34 mg of

$= (1.67493 \times 10^{-27}\text{ kg})(1.2046 \times 10^{22})$

$= 2.0176 \times 10^{-5}$ kg

The number of protons, electrons, and neutrons in an atom is independent of temperature and pressure conditions. Hence, the obtained values will remain unchanged if the temperature and pressure is changed

Q. 3. **How many neutrons and protons are there in the following nuclei?**

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

ANSWER:-

$^{13}_{6}C -$

Atomic mass =13

Atomic number Number of protons = 6

Number of neutrons = (Atomic mass - Atomic number)

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$= 13 - 6 = 7$

$^{16}_{8}O -$

Atomic mass = 16

Number of protons = 8

Number of neutrons = (Atomic mass - Atomic number)

= 16 - 8 = 8

$^{24}_{12}Mg$ -

Atomic mass = 24

Atomic number = Number of protons = 12

Number of neutrons (Atomic mass - Atomic number)

= 24 - 12 = 12

^{56}Fe

Atomic mass = 56

Atomic number = Number of protons = 26

Number of neutrons = (Atomic mass - Atomic number)

= 56 - 26 = 30

^{88}Sr-

Atomic mass = 88

Atomic number = Number of protons = 38

Number of neutrons = (Atomic mass - Atomic number)

= 88 - 38 = 50

Q. 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.**

ANSWER:-

(i) $^{35}_{17}Cl$ **(ii)** $^{233}_{92}U$ **(iii)** $^{9}_{4}Be$

Q. 5. **Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency (v) and wavenumber of the yellow light.**

ANSWER:-

From the expression,

$$\lambda = \frac{c}{v}$$

We get,

$$v = \frac{c}{\lambda}$$

Where,

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v = frequency of yellow light
c = velocity of light in vacuum = 3 × 10^8 m/s
λ = wavelength of yellow light = 580 nm = 580 × 10^{-9} m
Substituting the values in expression (i):

$$v = \frac{3\times10^{8}}{580\times10^{-9}} = 5.17\times10^{14}\,s^{-1}$$

Thus, frequency of yellow light emitted from the sodium lamp

$$= 5.17\times10^{14}\,s^{-1}$$

Wave number of yellow light, $\bar{v} = \frac{1}{\lambda}$

$$= \frac{1}{580\times10^{-9}} = 1.72\times10^{6}\,m^{-1}$$

Q. 6. **Find energy of each of the photons which (i) correspond to light of frequency 3×10^{15} Hz. (ii) have wavelength of 0.50 Å**

ANSWER:-

(i) E = hv

Where,

$h =$ Planck's constant $= 6.626\times10^{-34}$ Js

$v =$ frequency of light $= 3\times10^{15}$ Hz

Substituting the values in the given expression of E:

$$E = \left(6.626\times10^{-34}\right)\left(3\times10^{15}\right)$$

$$E = 1.988\times10^{-18}\,J$$

(ii) Energy (E) of a photon having wavelength (λ) is given by the expression $E = \frac{hc}{\lambda}$

h = Planck's constant $= 6.626\times10^{-34}$ Js

c = velocity of light in vacuum = 3 × 10^8 m/s

Substituting the values in the given expression of E:

$$E = \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{0.50\times10^{-10}} = 3.976\times10^{-15}\,J$$

$$\therefore E = 3.98\times10^{-15}\,J$$

Q. 7. **Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0 × 10^{-10} s.**

ANSWER:-

$$\text{Frequency of light} = \frac{1}{\text{Period}}$$

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$$= \frac{1}{2.0\times10^{-10}s} = 5.0\times10^{9}s^{-1}$$

Wavelength (λ) of light $= \frac{c}{v}$

Where,

c = velocity of light in vacumm = 3 × 10^8 m/s

Substituting the value in the given expression of λ:

$$\lambda = \frac{3\times10^{8}}{5.0\times10^{9}} = 6.0\times10^{-2}\text{m}$$

Wave number ($\overline{v}$) of light $= \frac{1}{\lambda} = \frac{1}{6.0\times10^{-2}} = 1.66\times10^{-1} = 16.66\text{m}$

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

ANSWER:-

Energy (E) of a photon hv

Energy (En) of 'n'photons = nhv

$$\Rightarrow n = \frac{E_n\lambda}{n_n}$$

Where,

λ = wavelength of light = 4000 pm = 4000 × 10^{-12}) m

c = velocity of light in vacuum = 3 × 10^8 m/s

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

Substituting the values in the given expression of n:

$$\text{n} = \frac{(1)\times\left(4000\times10^{-12}\right)}{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)} = 2.012\times10^{16}$$

Hence, the number of photons with a wavelength of 4000 pm and energy of 1 J are 2.012 × 10^{16}

Q. 9. **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).**

ANSWER:-

(i) Energy (E) of a photon $= hv = \frac{hc}{\lambda}$

Where,

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

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c = velocity of light in vacuum = 3 × 10^8 m/s

λ = wavelength of photon = 4 × 10^{-7} m

Substituting the values in the given expression of E:

$$E = \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{(4\times10)^{-7}} = 4.9695\times10^{-19}\,J$$

Hence, the energy of the photon is 4.97 × 10^{-19} J.

(ii) The kinetic energy of emission E_k is given by

= $hv-hv_o$

= (E-V)eV

$$=\left(\frac{4.9695\times10^{-19}}{1.6020\times100^{-19}}\right)ev-2.13eV$$

= (3.1020 – 2.13) eV

= 0.9720 eV

Hence, the kinetic energy of emission is 0.97 eV.

(iii) The velocity of a photoelectron (v) can be calculated by the expression,

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

$$\Rightarrow v=\sqrt{\frac{2\left(hv-hv_0\right)}{m}}$$

where ($hv-hv_0$) is the kinetic energy of emission in Joules and 'm' is the mass of the photoelectron.

Substituting the values in the given expression of v:

$$v=\sqrt{\frac{2\times\left(0.9720\times1.6020\times10^{-19}\right)J}{9.10939\times10^{-31}kg}}$$

$$=\sqrt{0.3418\times10^{12}m^2s^{-2}}$$

v = 5.84 × 10^5 ms^{-1}

Hence, the velocity of the photoelectron is 5.84 × 10^5 ms^{-1}.

Q. 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}

ANSWER:-

$$\text{Energy of sodium}(E)=\frac{N_A hc}{\lambda}$$

$$=\frac{\left(6.023\times10^{23}mol^{-1}\right)\left(6.626\times10^{-34}Js\right)\left(3\times10^{8}ms^{-1}\right)}{242\times10^{-9}}$$

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$$= 4.947 \times 10^5\,\text{Jmol}^{-1}$$
$$= 494.7 \times 10^3\,\text{Jmol}^{-1}$$
$$= 494\text{kJmol}^{-1}$$

Q. 11. **A 25 watt bulb emits monochromatic yellow light of wavelength of 0.57µm. Calculate the rate of emission of quanta per second.**

ANSWER:-

Power of bulb, P = 25 Watt = 25Js^{-1}

Energy of one photon, $E = h\nu = \dfrac{hc}{\lambda}$

Substituting the values in the given expression of :

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{0.57 \times 10^{-6}} = 34..87 \times 10^{-20}\,\text{J}$$

E = 34.87 × 10^{-20} J

Rate of emission of quanta per second

$$= \frac{25}{34.87 \times 10^{-20}} = 7.169 \times 10^{19}\,\text{s}^{-1}$$

Q. 12. **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.**

ANSWER:-

Threshold wavelength of radian $(\lambda_0) = 6800\text{Å} = 6800 \times 10^{-10}\,\text{m}$ Threshold frequency (v_0) of the metal

$$\frac{c}{\lambda_0} = \frac{3 \times 10^8\,\text{ms}^{-1}}{6.8 \times 10^{-7}} = 4.41 \times 10^{14}\,\text{s}^{-1}$$

Thus, the threshold frequency (v_0) of the metal is 4.41 × 10^{14} s^{-1}.

Hence, work function (W_O) of the metal = hv_0

= (6.626 × 10^{-34} Js)(4.41 × 10^{14} s^{-1})

= 2.922 × 10^{-19} J)

Q. 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?**

ANSWER:-

The $n_i = 4$ to $n_f = 2$ transition will give rise to a spectral line of the Balmer series. The energy involved in the transition is given by the relation,

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

Substituting the values in the given expression of *E*:

$$E = 2.18 \times 10^{-18}\left[\frac{1}{4^2} - \frac{1}{2^2}\right]$$

$$= 2.18 \times 10^{18}\left[\frac{1-4}{16}\right]$$

$$= 2.18 \times 10^{-18} \times \left(-\frac{3}{16}\right)$$

$E = -(4.0875 \times 10^{-19}\ \text{J})$

The negative sign indicates the energy of emission.

Wavelength of light emitted $(\lambda) = \dfrac{hc}{E}$

$$\left(\text{since } E = \frac{hc}{\lambda}\right)$$

Substituting the values in the given expression of λ:

$$\lambda = \frac{\left(6.626 \times 10^{-34}\right)\left(3 \times 10^{8}\right)}{4.08875 \times 10^{-19}}$$

$\lambda = 4.8631 \times 10^{-7}\text{m}$

$= 486.3 \times 10^{-9}\ \text{m}$

$= 486\ \text{nm}$

Q. 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).**

ANSWER:-

The expression of energy is given by,

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

Where,

Z = atomic number of the atom

n = principal quantum number

For ionization from $n_1 = 5$ to $n_2 = \infty$

$$\text{Ä}E = E_{\infty} - E_5$$

The expression of energy is given by,

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$$E_n = \frac{-\left(2.18\times10^{-18}\right)Z^2}{n^2}$$

Where,

Z = atomic number of the atom

n = principal quantum number

For ionization from $n_1 = 5$ to $n_2 = \infty$

$\Delta E = E_\infty - E_5$

$$= \left[\left\{\frac{-\left(2.18\times10^{-18} J\right)(1)^2}{(\infty)^2}\right\} - \left\{\frac{-\left(2.18\times10^{-18} J(1)^2\right)}{(5)^2}\right\}\right]$$

$$= \left(2.18\times10^{-18}\text{J}\right)\left(\frac{1}{(5)^2}\right)\left(\text{since } \frac{1}{\infty} = 0\right)$$

$= 0.0872 \times 10^{-18}$ J

Hence, the energy required for ionization from to $n = \infty$ is 8.72×10^{-20} J. Energy required for $n_1 = 1$ to $n = \infty$

Ä$E' = E_\infty - E_1$

$$= \left[\left\{\frac{-\left(2.18\times10^{-18}\right)(1)^2}{(\infty)^2}\right\} - \left\{\frac{-\left(2.18\times10^{-18}\right)(1)^2}{(1)^2}\right\}\right]$$

$$= \left(2.18\times10^{-18}\right)[1-0]$$

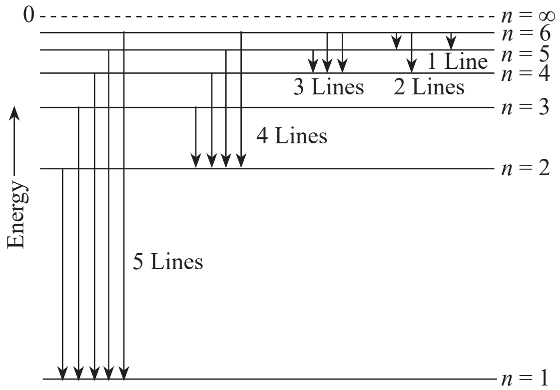
$$= 2.18\times10^{-18}\,\text{J}$$

Hence, less energy is required to ionize an electron in the orbital of hydrogen atom as compared to that in the ground state

Q. 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?**

ANSWER:-

When the excited electron of an H atom in drops to the ground state, the following transitions are possible:



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Hence, a total number of (5 + 4 + 3 + 2 + 1)15 lines will be obtained ir

The number of spectral lines produced when an electron in the nth level is given by $\frac{n(n-1)}{2}$

Given, n = 6

Number of spectral lines $= \frac{6(6-1)}{2} = 15$

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

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

ANSWER:-

(i) Energy associated with the fifth orbit of hydrogen atom is calculated as:

$$E_5 = \frac{-\left(2.18 \times 10^{-18}\right)}{(5)^2} = \frac{-2.18 \times 10^{-18}}{25}$$

$E_5 = -8.72 \times 10^{-20}$ J

Radius of Bohr's orbit for hydrogen atom is given by,

$r_n = (0.0529 \text{ nm})n^2$

For, n = 5

$r_5 = (0.052 \text{ nm})(5)^2$

$r_5 = 1.3225$ nm)

Q. 17. Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

ANSWER:-

For Balmer series $n_1 = 2$

If this line possesses longest wavelength

(i.e., lowest energy) then $n_2 = 3$

$$\bar{v} = \frac{1}{\lambda} = 109677\left[\frac{1}{2^2} - \frac{1}{3^2}\right]$$

$= 1.523 \times 10^4 \text{ cm}^{-1}$

$= 1.523 \times 10^6$ m

Q. 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 -2.18×10^{-11} ergs.

ANSWER:-

Energy (E) of the Bohr orbit of an atom is given by,

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$$E_n = \frac{-\left(2.18\times10^{-18}\right)Z^2}{n^2}$$

Where,

atomic number of the atom

Ground state energy = -2.18 × 10^{-11} ergs

= -2.18 × 10^{-11} × 10^{-7} J

= -2.18 × 10^{-18} J

Energy required to shift the electron from n = 1 to n = 5 is given as:

$\Delta E = E_5 - E_1$

$$= \frac{-\left(2.18\times10^{-18}\right)(1)^2}{(5)^2} - \left(-2.18\times10^{-18}\right)$$

$$= \left(2.18\times10^{-18}\right)\left[1-\frac{1}{25}\right]$$

$$= \left(2.18\times10^{-18}\right)\left(\frac{24}{25}\right)$$

$= 2.0928 \times 10^{-18}$ J

$$\text{Wavelength of emitted light} = \frac{hc}{E}$$

$$= \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{2.0928\times10^{-18}}$$

$$= 9.498\times10^{-8}\,\text{m}$$

Q. 19. The electron energy in hydrogen atom is given by En = (–2.18 × 10^{-18})/n² 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?

ANSWER:-

Given

$$E_n = -\frac{2.18\times10^{-18}}{n^2}J$$

Energy required for ionization from is given by,

$$\ddot{A}E = E_\infty - E_2$$

$$= \left[\left(\frac{-2.18\times10^{-18}}{(\infty)^2}\right) - \left(\frac{-218\times10^{-18}}{(2)^2}\right)\right]\text{J}$$

$$= \left[\frac{2.18\times10^{-18}}{4} - 0\right]\text{J}$$

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$$= 0.545 \times 10^{-18}\,\text{J}$$

$$\text{ÄE} = 5.45 \times 10^{-19}\,\text{J}$$

$$\lambda = \frac{\text{hc}}{\text{ÄE}}$$

Here, is the longest wavelength causing the transition.

$$\lambda = \frac{\left(6.626 \times 10^{-34}\right)\left(3 \times 10^{8}\right)}{5.45 \times 10^{-19}} = 3.647 \times 10^{-7}\,\text{m}$$

$$= 3647 \times 10^{-10}\,\text{m}$$

$$= 3647\text{Å}$$

Q. 20. **Calculate the wavelength of an electron moving with a velocity of 2.05 × 107 m s^{-1}**

ANSWER:-

The wavelength of an electron is obtained from de-Broglie equation.

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

$$= \frac{6.626 \times 10^{-34}\,\text{Js}}{9.11 \times 10^{-31}\,\text{kg} \times 2.05 \times 10^{7}\,\text{m/s}}$$

$$= 3.55 \times 10^{-11}\,\text{m}$$

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

ANSWER:-

"From de Broglie's equation"

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

Given,

"Kinetic energy (K.E) of the electron " = 3.0 × 10^{-25} J

$$\text{Since } K.E = \frac{1}{2}\text{mv}^2$$

$$\therefore \text{ Velocity}(v) = \sqrt{\frac{2K.E}{m}}$$

$$= \sqrt{\frac{2\left(3.0 \times 10^{-25}\,\text{J}\right)}{9.10939 \times 10^{-31}}}$$

$$= \sqrt{6.5866 \times 10^{4}}$$

$$v = 811.579\text{ms}^{-1}$$

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Substituting the value in the expression of λ :

$$\lambda = \frac{6.626\times10^{-34}\,\text{Js}}{\left(9.10939\times10^{-31}\text{kg}\right)\left(811.579\text{ms}^{-1}\right)}$$

$$\lambda = 8.9625\times10^{-7}\,\text{m}$$

Hence, the wavelength of the electron is 8.9625×10^{-7} m

Q. 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.**

ANSWER:-

$\text{Na} \Rightarrow 1s^2, 2s^2 2p^6, 3s^1$

$\text{Na}^+ \Rightarrow 1s^2, 2s^2 2p^6$

similarly for $\text{Mg}^{2+} \Rightarrow 1s^2, 2s^2 2p^6$

$\text{Na}^+, \text{Mg}^{2+}$, are isoelectronic species. They contain 10 electrons each.

$\text{S}^{2-} \Rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6$

$\text{Ar} \Rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6$

$\text{K}^+ \Rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6$

$\text{Ca}^{2+} \Rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6$

Ar, S^{2-}, Ca^{2+} and K^+ are isoelectronic species. They contain 18 electrons each.

Q. 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) $2p^3$ and (c) $3p^5$?**

(iii) **Which atoms are indicated by the following configurations ? (a) [He] $2s^1$ (b) [Ne] $3s^2 3p^3$ (c) [Ar] $4s^2 3d^1$.**

ANSWER:-

(i) **(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) **(a)** $1s^2 2s^2 2p^6 3s^1 (Z = 11)$

(b) $1s^2 2s^2 2p^3 (Z = 7)$

(c) $1s^2 2s^2 2p^6 3s^2 3p^5 (Z = 17)$

(iii) **(a)** $1s^2 2s^1$, then atom is Li

(b) $1s^2 2s^2 2p^6 3s^2 3p^3$, the atom is P .

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$, the atom is Sc.

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

ANSWER:-

For g -orbitals, 1 = 4.

As for any value '*n*' of principal quantum number, the Azimuthal quantum number (1) can have a value from zero to (n − 1).

∴ For , minimum value of n = 5

Q. 25. **An electron is in one of the 3d orbitals. Give the possible values of n, l and m_l for this electron.**

ANSWER:-

For an electron in 3d orbital, the possible values of the quantum numbers are n = 3, l = 2, m_l = −2, −1, 0, +1, +2.

Q. 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.**

ANSWER:-

The number of protons is equal to the number of electrons which is 29.

It also corresponds to the atomic number.

The electronic configuration of the element will be $[Ar]^{18}3d^{10}4s^{1}$

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

ANSWER:-

H_2^+

Number of electrons present in hydrogen molecule (H_2) = 1 + 1 = 2

∴ Number of electrons in $H_2^+ = 2 - 1 = 1$

H_2 :

Number of electrons in H_2 = 1 + 1 = 2

O_2^+

Number of electrons present in oxygen molecule (O_2) = 8 + 8 = 16

∴ Number of electrons in $O_2^+ = 16 - 1 = 1$

Q. 28. **(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.**

ANSWER:-

(i) The possible values of l and m1 are :

l m_l

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0	0
1	−1, 0, +1
2	−2, −1, 0, +1, +2

(ii) The quantum numbers (m_l and l) of electrons for 3d orbital are l = 2, m_l = −2, −1, 0, +1, +2.

(iii) 2s, 2p orbitals are possible.

Q. 29. Using s, p, d 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**

ANSWER:-

n = no of principle shell

l = azimuthal quantum number

l = 0 for s orbital

l = 1 for p orbital

l = 2 for d orbital

l = 3 for f orbital

The orbital with the following quantum numbers are given below.

(a) n = 1, *l* = 0 : 1s orbital

(b) n = 3; *l* = 1 : 3p orbital

(c) n = 4; *l* = 2 : 4d orbital

(d) n = 4; *l* = 3 : 4f orbital

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

(a) n = 0, *l* = 0, m_l = 0, m_s = + ½

(b) n = 1, *l* = 0, m_l = 0, m_s = – ½

(c) n = 1, *l* = 1, m_l = 0, m_s = + ½

(d) n = 2, *l* = 1, m_l = 0, m_s = – ½

(e) n = 3, *l* = 3, m_l = –3, m_s = + ½

(f) n = 3, *l* = 1, m_l = 0, m_s = + ½

ANSWER:-

(a) As zero value for n is not possible, this set of quantum numbers is not possible.

(b) This set of quantum numbers is possible.

(c) As l = 1 is not possible for n = 1, this set of quantum numbers is not possible. For a given value of n, l can have values of 0,1,2...n − 1 only.

(d) This set of quantum numbers is possible.

(e) As l = 3 is not possible for n = 3, this set of quantum numbers is not possible. For a given value of n, l can have values of 0, 1, 2... n − 1 only.

(f) This set of quantum numbers is possible.

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Q. 31. **How many electrons in an atom may have the following quantum numbers?**

(a) $n = 4, m_s = -½$ **(b) $n = 3, l = 0$**

ANSWER:-

(a) Total number of electrons in an atom for a value of $n = 2n^2$

$\therefore$ For n = 4

Total number of electrons = $2(4)^2$

= 32

The given element has a fully filled orbital as

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

Hence , all the electrons are paired .

$\therefore$ Number of electrons (having n = 4 and m_s = 12) = 16

(b) n = 3, l = 0 indicates that the electrons are present in the 3s orbital. Therefore, the number of electrons having n = 3 and l = 0 is 2.

Q. 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.**

ANSWER:-

The expression for the angular momentum is $mvr = \frac{nh}{2\pi}$, or $2\pi r = \frac{nh}{mv}$ .... (i)

The wavelength is given by the De Broglie's equation.

$\lambda = \frac{h}{mv}$...... (ii)

From (i) and (ii), we have

$2\pi r = n\lambda$

Hence, the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie's wavelength associated with the electron revolving around the orbit.

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

ANSWER:-

We have to compare wavelength of transition in the H-spectrum with the Balmer transition n = 4 to n = 2 of He^+ spectrum.

$\therefore \lambda_H = \lambda_{He+}$

$$\therefore R_H Z_H^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = R_H Z_{He^+}^2 \left[\frac{1}{2^2} - \frac{1}{4^2}\right]$$

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$$1\times\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]=4\times\left(\frac{1}{4}-\frac{1}{16}\right)$$

$$\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]=4\times\frac{4-1}{16}$$

$$\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]=\frac{3}{4}$$

If $n_1 = 1$, then $n_2 = 2, 3, \ldots$

For first line $n_2 = 2$, $n_1 = 1$

$$\left[\frac{1}{1^2}-\frac{1}{2^2}\right]=\frac{1}{1}-\frac{1}{4}=\frac{3}{4}$$

Hence, transition $n_2 = 2$ to $n_1 = 1$ will give spectrum of the same wavelength as that of Balmer transition, $n_2 = 4$ to $n_1 = 2$ in He^+.

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

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

ANSWER:-

Given that, The I.E for atom is $\frac{2\pi^2 me^4}{h^2} = 2.18\times10^{-18}\,J/$ atom

For He^+ $Z^2 = 2^2 = 4$

For the given process, the energy required $= 2^2\times\frac{2\pi^2 me^4}{h^2}$

$$= 4\times\frac{2\pi^2 me^4}{h^2}$$

$$= 4\times2.18\times10^{-18}$$

$$= 8.72\times10^{-18}\,J/\text{ion.}$$

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

ANSWER:-

The diameter of 1 C atom is 0.15 nm = 0.15×10^{-7} cm.

In 20 cm length, $\frac{20}{0.1.5\times10^{-7}} = 1.33\times10^{9}$ carbon atoms can be placed

Q. 36. **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.**

ANSWER:-

The length covered by 1 carbon atom corresponds to its diameter.

The diameter of carbon atom is $\frac{2.4}{2\times10^{8}} = 1.2\times10^{-8}$ cm.

The radius of carbon atom is $\frac{1.2\times10^{-8}}{2} = 0.6\times10^{-8}$ cm $= 0.06$nm.

Q. 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.**

ANSWER:-

(i) The diameter of zinc atom is 2.6Å$= 2.6\times10^{-10}$ m. The radius of Zn atom is

$$\frac{2.6\times10^{-10}}{2} = 1.3\times10^{-10}\,\text{m} = 130\times10^{-12}\,\text{m} = 130\text{pm}$$

(ii) The number of Zn atoms present on 1.6 cm of length are

$$\frac{1.6}{2.6\times10^{-8}} = 6.154\times10^{7}$$

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

ANSWER:-

One electron has charge 1.602×10^{-19} C.

The number of electrons present in the charged particle is

$$\frac{2.5\times10^{-16}}{1.602\times10^{-19}} = 1.56\times10^{3}$$

thus, the number of electrons present in the charged particle is $=1.56\times10^{3}$.

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

ANSWER:-

The static electric charge on the oil drop is, $q = 1.282 \times 10^{-18}$C.

We know that Charge $q = ne$

where n is the no. of electrons and is the charge of 1 electron, i.e.

1.602×10^{-19} C

The number of electrons present in the oil drop is, $n = \frac{1.282\times10^{-18}}{1.602\times10^{-19}} = 8$

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Q. 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?**

ANSWER:-

In Rutherford's experiment, if lighter atoms like aluminum were used as the target, the number of α-particles deflected backward or at large angles would be significantly smaller. This is because the lighter nuclei are smaller in size and carry a smaller positive charge. As a result, most α-particles would pass through the target atoms without significant deflection. Additionally, due to the smaller nuclear charge, the deflection angle of the α-particles would also be much smaller.

Q. 41. **Symbols $^{79}_{35}\mathrm{Br}$ and $^{79}\mathrm{Br}$ can be written, whereas symbols $^{35}_{79}\mathrm{Br}$ and $^{35}\mathrm{Br}$ are not acceptable. Answer briefly:**

ANSWER:-

In the symbol. $^{B}_{A}X$ of an element : A denotes the atomic number of the element Br denotes the mass number of the element.

The atomic number of the element can be identified from its symbol because no two elements can have the atomic number. However, the mass numbers have to be mentioned in order to identify the elements. Thus, Symbols. $^{79}_{35}\mathrm{Br}$ and. $^{79}\mathrm{Br}$ are accepted because atomic number of Br will remain 35 even if not mentioned. Symbol. $^{35}_{99}\mathrm{Br}$ is not accepted because atomic number of Br cannot be 79 (more than the mass number). Similarly, Symbol $^{35}\mathrm{Br}$ cannot be accepted because mass number has to be mentioned. This is needed to differentiate the isotopes of an element.

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

ANSWER:-

Let x be the number of protons.
The number of neutrons is x + 0.317x = 1.317x.
The mass number is .
It is equal to the sum of the number of protons and the number of neutrons.
81 = x + 1.317x = 2.317x
x = 35
Hence, the symbol is $^{81}_{35}\mathrm{Br}$.

Q. 43. **An ion with mass number 37 possesses one unit of negative charge. If the ion conatins 11.1% more neutrons than the electrons, find the symbol of the ion.**

ANSWER:-

The mass number (A) of the ion is 37.

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The ion has a negative charge of -1, which means it has one more electron than protons.

The ion contains 11.1% more neutrons than electrons**.**

Let:

x = number of electrons in the ion.

Since the ion has a -1 charge, the number of protons in the neutral atom will be $x - 1$.

The number of neutrons is given as $x + 11.1\%$ of x.

Number of neutrons = $x + 0.111 \cdot x = 1.111x$

The mass number (A) is the sum of the number of protons and neutrons:

A = Number of protons + Number of neutrons

Substituting the values we have:

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

$37 = x - 1 + 1.111x$

$37 = 2.111x - 1$

Adding 1 to both sides:

$38 = 2.111x$

$x = 38/2.111 \approx 18$

The number of protons in the neutral atom is $x - 1 = 18 - 1 = 17$.

The atomic number (Z) is 17, which corresponds to the element chlorine (Cl).

The ion has a mass number of 37 and an atomic number of 17. Therefore, the symbol of the ion is:

Cl^-

This indicates it is a chloride ion with a mass number of 37.

The symbol of the ion is Cl^- (Chlorine ion with mass number 37)

Q. 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.

ANSWER:-

The mass number (A) of the ion is 56.

The ion has a charge of +3, which means it has 3 more protons than electrons.

The number of neutrons is 30.4% more than the number of electrons.

Let:

x = number of electrons

Since the ion has a +3 charge, the number of protons will be $x + 3$.

The number of neutrons will be $1.304x$ (since it is 30.4% more than x).

The mass number is the sum of the number of protons and neutrons:

A = Number of Protons + Number of Neutrons

Substituting the values we have:

$56 = (x + 3) + (1.304x)$

Combine like terms:

$56 = x + 3 + 1.304x$

$56 = 2.304x + 3$

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Subtract 3 from both sides:

$53 = 2.304x$

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

Number of electrons $x = 23$

Number of protons $= x + 3 = 23 + 3 = 26$

Number of neutrons $= 1.304x = 1.304 \times 23 \approx 30$

The atomic number (number of protons) is 26. The element with atomic number 26 is Iron (Fe).

Since the ion has a +3 charge, we write the symbol as: Fe^{3+}

Q. 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**

ANSWER:-

The increasing order of frequency is Cosmic rays > Xrays > amber colour > microwave > FM.

Q. 46. **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**

ANSWER:-

The power of the laser is,

$$E = \frac{Nhc}{\lambda}$$

where

E = power of the laser

N = No of photons = 5.6×10^{24}

h = Planks constant = 6.626×10^{-34}

c = speed of light = 3.0×10^{8}

$$E = \frac{5.6\times10^{24}\times6.62\times10^{-34}\times3\times10^{8}}{3.37.1\times10^{-9}} = 3.3\times10^{6}\,J/s = 3.3\times10^{6}\,W$$

Q. 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.**

ANSWER:-

Wavelength of radiation emitted = 616 nm = 616×10^{-9} m (Given)

(a) Frequency of emission (v)

$$v = \frac{c}{\lambda}$$

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Where,
c = velocity of radiation
λ = wavelength of radiation

$$v = \frac{3.0\times10^{8}\,\mathrm{m/s}}{616\times10^{-9}\,\mathrm{m}}$$

$$= 4.87\times10^{8}\times10^{9}\times10^{-3}\,\mathrm{s}^{-1}$$

$$v = 4.87\times10^{14}\,\mathrm{s}^{-1}$$

Frequency of emission (v) = 4.87×10^{14} s^{-1}

(b) Velocity of radiation, (c) = 3.0×10^{8} ms^{-1}

Distance travelled by this radiation in 30 s

= $(3.0 \times 10^{8}\ \mathrm{ms}^{-1})(30\ \mathrm{s}) = 9.0 \times 10^{9}$ m

(c) Energy of quantum (E) = hv

$(6.626 \times 10^{-34}\ \mathrm{Js})(4.87 \times 10^{14}\ \mathrm{s}^{-1})$

Energy of quantum (E)=32.27×10^{-20}) J

(d) Energy of one photon (quantum) = 32.27×10^{-20} J

Therefore, 32.27×10^{-20} J of energy is present in 1 quantum.

Number of quanta in 2 J of energy

$$= \frac{2J}{32.27\times10-20\mathrm{J}}$$

$$= 6.19\times10^{18}$$

$$= 6.2\times10^{18}$$

Q. 48. 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.

ANSWER:-

The energy of 1 photon,

$$\mathrm{E} = \mathrm{h}v = \frac{\mathrm{hc}}{\lambda} == \frac{6.626\times10^{-34}\times3.0\times10^{8}}{600\times10^{-9}} = 3.31\times10^{-19}\,\mathrm{J}$$

The number of photons is $\frac{3.15\times10^{-18}}{3.31\times10^{-19}} = 9.52 = 10$.

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

ANSWER:-

Time duration (t) = 2ns = 2×10^{-9}s

$$\text{Frequency } (v) = \frac{1}{t} = \frac{1}{2\times10^{-9}\,\mathrm{s}} = \frac{10^{9}}{2}\,\mathrm{s}^{-1}$$

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Energy of one photon, $E = hv = \left(6.626\times10^{-34}\,\text{Js}\right)\times\left(10^{9}/2\text{s}^{-1}\right) = 3.25\times10^{-25}\,\text{J}$

No. of photons 2.5×10^{5}

$\therefore$ Energy of source $= 3.3125\times10^{-25}\,\text{J}\times2.5\times10^{15} = 8.28\times10^{-10}\,\text{J}$

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

ANSWER:-

For λ_1 = 58912m

Frequency of transition $(v_1) = \dfrac{c}{\lambda_1}$

$$= \frac{3.0\times10^{8}\,\text{ms}^{-1}}{589\times10^{-9}\,\text{ma}}$$

Frequency oftransition $(v_1) = 5.093\times10^{14}\,\text{s}^{-1}$

Similarly, for λ_2=589.6 mm

Frequency of transition $(v_2) = \dfrac{c}{\lambda_2}$

$$= \frac{3.0\times10^{8}\,\text{ms}^{-1}}{589.6\times10^{-9}\,\text{m}}$$

Frequency of transition $(v_2) = 5.088\times10^{14}\,\text{s}^{-1}$

Energy difference (ΔE) between excited states = $E_1 - E_2$

Where,

E_2 = energy associated with λ_2

E_1 = energy associated with λ_1

$\Delta E = hv_1 - hv_2$

$= h(v_1 - v_2)$

$= (6.626 \times 10^{-34}\text{ Js})(5.093 \times 10^{14} - 5.088 \times 10^{14})\text{ s}^{-1})$

$= (6.626 \times 10^{-34})\text{ J})(5.0 \times 10^{-3}) \times 10^{14})$

$\Delta E = 3.31 \times 10^{-22}$ J

Q. 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.**

ANSWER:-

Given Data:

- Work function of cesium (W_o) = 1.9eV
- Wavelength of incident radiation (λ) = 500 nm = 500×10^{-9} m
- Planck's constant (h) = $6.626 \times 10^{-34})$ J·s

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- Speed of light (c) = 3.0×10^8 m/s

-1eV=1.602×10^{-19} J

(a) Calculate the threshold wavelength (λ_0):

The threshold wavelength can be calculated using the formula:

$$\lambda_o = \frac{hc}{W_0}$$

First, we need to convert the work function from eV to Joules:

$$W_0 = 1.9\text{eV} \times 1.602 \times 10^{-19}\,\text{J/eV} = 3.046 \times 10^{-19}$$

$$\lambda_0 = \frac{(6.626 \times 10^{-34}\,\text{J}\cdot\text{s}) \times (3.0 \times 10^8\,\text{m/s})}{3.046 \times 10^{-19}\,\text{J}}$$

$$\lambda_0 = \frac{1.9878 \times 10^{-25}\,\text{J}\cdot\text{m}}{3.046 \times 10^{-19}\,\text{J}} \approx 6.53 \times 10^{-7}\,\text{m} = 653$$

(b) Calculate the threshold frequency (vo):

The threshold frequency can be calculated using the formula:

$$v_0 = \frac{W_0}{h}$$

$$v_0 = \frac{3.046 \times 10^{-19}\,\text{J}}{6.626 \times 10^{-34}\,\text{J}\cdot\text{s}} \approx 4.593 \times 10^{14}\,\text{s}^{-1}$$

Kinetic Energy of the Ejected Photoelectron:

The kinetic energy (KE) of the ejected photoelectron can be calculated using the formula:

KE = hν - W_0

First, we need to calculate the frequency (*v*) of the incident radiation:

$$v = \frac{c}{\lambda} = \frac{3.0 \times 10^8\,\text{m/s}}{500 \times 10^{-9}\,\text{m}} = 6.0 \times 10^{14}\,\text{s}^{-1}$$

$$KE = hv - W_0 = (6.626 \times 10^{-34}\,\text{J}\cdot\text{s})$$

$$\times (6.0 \times 10^{14}\,\text{s}^{-1}) - 3.046 \times 10^{-19}\,\text{J}$$

Calculating this gives:

$$KE = 3.976 \times 10^{-19}\,\text{J} - 3.046 \times 10^{-19}\,\text{J} = 9.30 \times 10^{-20}\,\text{J}$$

Velocity of the Ejected Photoelectron:

The kinetic energy can also be expressed in terms of the mass and velocity of the electron:

$$KE = \frac{1}{2}mv^2$$

Where *m* is the mass of the electron (9.109×10^{-31}) kg).

Rearranging for velocity (v):

$$v = \sqrt{\frac{2 \times KE}{m}}$$

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$$v=\sqrt{\frac{2\times 9.30\times 10^{-20}\,\text{J}}{9.109\times 10^{-31}\,\text{kg}}}$$

$v \approx 4.52 \times 10^5$ m/s

- (a) Threshold Wavelength (λ_0) = 653 nm
- (b) Threshold Frequency (v_o) = 4.593×10^{14} s^{-1}
- Kinetic Energy of ejected photoelectron = 9.30×10^{-20} J
- Velocity of ejected photoelectron = 4.52×10^5 m/s

Q. 52. Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength and (b) Plancks constant.

λ(nm)	500	450	400
v × 10^{-5} (cms^{-1})	2.55	4.35	5.35

ANSWER:-

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

Then $h\left(v-v_0\right)=\frac{1}{2}mv^2$ or $hc\left(\frac{1}{\lambda}-\right)=\frac{1}{2}mv^2$

Substituting the given results of the three experiments, we get

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

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

$$\frac{hc}{10^{-9}}\left(\frac{1}{400}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m\left(5.20\times 10^6\right)^2 \ldots(iii$$

Dividing equation (ii) by equation (i), we get

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

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

$$\text{or } \lambda_0-450=2.619\lambda-1309.5$$

$$\text{or} 1.619\lambda_0=859.5\Rightarrow\lambda_0=531\text{nm}$$

Subtituting this value of equation (iii), we get

$$\frac{h\times\left(3\times 10^8\right)}{10^{-9}}\left(\frac{1}{400}-\frac{1}{531}\right)=\frac{1}{2}\left(9.11\times 10^{-31}\right)$$

$$\times\left(5.20\times 10^6\right)^2$$

$$h=6.66\times 10^{-34}\,\text{Js}$$

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Q. 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.**

ANSWER:-

The energy of the incident radiation,

$$E = h\nu = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\times3.0\times10^{8}}{256.7\times10^{-9}} = 7.74\times10^{-19}\,J$$

Convert the unit in eV .

$$E = \frac{7.74\times10^{-19}}{1.602\times10^{-19}} = 4.83eV$$

The potential energy is equal to 0.35 eV .

The work function is 4.83eV - 0.35eV = 4.48eV.

Q. 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 × 10^7m s^{-1}, calculate the energy with which it is bound to the nucleus.**

ANSWER:-

The energy of the photon,

E=hc/λ

$$= \frac{6.626\times10^{-34}\times3\times10^{8}}{150\times10^{-12}}$$

$$=1.325\times10^{-15}\,J$$

The energy of the ejected electron is

$$\frac{1}{2}mv^2 = \frac{1}{2}\times9.11\times10^{-31}\times\left(1.5\times10^{7}\right)^2 = 1.025\times10^{-16}$$

The energy with which the electron was bound to the nucleus is:

Energy of photon - Energy ejected by electron

= 13.25 × 10^{-16})$^{-1}$.025 × 10^{-16} = 12.225 × 10^{-16} J

= 7.63 × 10^3 eV)

Q. 55. **Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as v = 3.29 × 1015 (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.**

ANSWER:-

“Wavelength of transition = 1285 nm = 1285 × 10^{-9} m(Given)

$$\nu = 3.29\times10^{15}\left(\frac{1}{3^2}-\frac{1}{n^2}\right) \text{ (Given)}$$

Since $v = \dfrac{c}{\lambda}$

$$= \frac{3.0\times10^{8}}{1285\times10^{-9}\,\text{m}}$$

$$v = 2.33\times10^{14}\,\text{s}^{-1}$$

Substituting the value of in the given expression,

$$3.29\times10^{15}\left(\frac{1}{9}-\frac{1}{\text{n}^2}\right) = 2.33\times10^{14}$$

$$\frac{1}{9}-\frac{1}{n^2} = \frac{2.33\times10^{14}}{3.29\times10^{15}}$$

$$\frac{1}{9}-0.7082\times10^{-1} = \frac{1}{\text{n}^2}$$

$$\Rightarrow \frac{1}{\text{n}^2} = 1.1\times10^{-1}-0.7082\times10^{-1}$$

$$\Rightarrow \frac{1}{\text{n}^2} = 1.1\times10^{-1}-0.7082\times10^{-1} \quad \frac{1}{\text{n}^2} = 4.029\times10^{-2}$$

$$\text{n} = \sqrt{\frac{1}{4.029\times10^{-2}}}$$

N = 4.98

$\text{N} \approx 5$

Hence, for the transition to be observed at .

The spectrum lies in the infra-red region.

Q. 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.**

ANSWER:-

Radius of nth orbit of H-like particles

$$= \frac{0.529n^2}{Z}\text{Å} = \frac{52.9n^2}{Z}\,pm$$

$$r_1 = 1.3225\text{nm} = 1322.5\text{pm} = 52.9n_1^2$$

$$r_2 = 211.6\text{pm} = \frac{52.9n_2^2}{Z}$$

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

$$\text{or}\ \frac{n_1^2}{n_2^2} = 6.25\ \text{or}\ \frac{n_1}{n_2} = 2.5$$

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If $n_2 = 2$, $n_1 = 5$. Thus the transition is from 5 th orbit to orbit. It belongs to Balmer series.

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

$$= 1.097 \times \frac{21}{100} \times 10^7 \text{m}^{-1}$$

$$\text{or } \lambda = \frac{1}{\bar{v}} = \frac{100}{1.097 \times 21 \times 10^7} \text{m}$$

$$= 434 \times 10^{-9} \text{m} = 434 \text{nm}$$

It lies in the visible region.

Q. 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 1.6×10^6 ms^{-1}, calculate de Broglie wavelength associated with this electron.**

ANSWER:-

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{9.11 \times 10^{-31} \times 1.6 \times 10^6} = 4.55 \times 10^{-10} \text{m} = 455 \text{pm}$$

Q. 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.**

ANSWER:-

$$v = \frac{h}{m\lambda} = \frac{6.626 \times 10^{-34}}{1.675 \times 10^{-27} \times 800 \times 10^{-12}} = 4.94 \times 10^2 \text{m/s}$$

Q. 59. **If the velocity of the electron in Bohr's first orbit is 2.19×106 ms^{-1}, calculate the de Broglie wavelength associated with it.**

ANSWER:-

According to de Broglie's equation

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

where, de-Broglie wavelength

$h = 6.626 \times 10^{-34}$ J s = Planck's constant

m = mass of electron = 9.1×10^{-31}kg

velocity of electron, $v = 2.19 \times 10^6$ m

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$$\lambda = \frac{6.626\times10^{-34}}{9.1\times10^{-31}\times2.19\times10^{6}}$$

$$\lambda = 0.332\times10^{-9} = 3.32\times10^{-10}\,\text{m}$$

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

ANSWER:-

The associated wavelength is obtained from de-Broglie equation Here

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

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

Substituting the above values in debroglie equation, we get

$$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}}{0.1\times4.37\times10^{5}} = 1.516\times10^{-28}\,\text{m}$$

Q. 61. **If the position of the electron is measured within an accuracy of $\pm 0.002\text{nm}$, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is $\frac{h}{4\pi m}\times 0.05\text{nm}$ is there any problem in defining this value?**

ANSWER:-

By applying the uncertainty principle,

$$\Delta P = \frac{h}{4\pi\Delta x} = \frac{6.626\times10^{-34}}{4\times3.1416\times0.002\times10^{-9}} = 2.63\times10^{-23}\,\text{kgm/s}$$

The uncertainty in momentum is,

$$\frac{h}{4\pi m\times0.002\times10^{-9}} = \frac{h\times5\times10^{11}}{4\times\pi\times m}$$

This is much larger than the actual momentum, i.e, $\frac{h}{4\pi m}\times 0.05$nn cannot be defined.

Q. 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 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. **$n = 4, l = 1, m_l = 0$, $m_s = +1/2$**

ANSWER:-

1. The quantum numbers $n = 4, l = 2, m_l = -2, m_s = -1/2$ represents 4d orbital.
2. The quantum numbers $n = 3, l = 2, m_l = 1, m_s = +1/2$ represents 3d orbital.
3. The quantum numbers $n = 4, l = 1, m_l = 0, m_s = +1/2$ represents 4p orbital.

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4. The quantum numbers $n = 3, l = 2, m_l = -2, m_s = -1/2$ represents 3d orbital.
5. The quantum numbers $n = 3, l = 1, m_l = -1, m_s = +1/2$ represents 3p orbital.
6. The quantum numbers $n = 4, l = 1, m_l = 0, m_s = +1/2$ represents 4p orbital.
The order of increasing energies is (v) < (ii) = (iv) < (vi) = (iii) < (i).

Q. 63. **The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electron in 4p orbital. Which of these electron experiences the lowest effective nuclear charge?**

ANSWER:-

The distance between the nucleus and 4p electrons is maximum.
Hence, 4p electrons will experience the lowest effective nuclear charge.

Q. 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.**

ANSWER:-

The orbitals 2s, 4d, 3p will experience the largest effective nuclear charge as they are closest to the nucleus.

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

ANSWER:-

Silicon has greater nuclear charge (+14) than aluminium (+13). Hence, the unpaired 3p electron in case of silicon will experience more effective nuclear charge.

Q. 66. **Indicate the number of unpaired electrons in : (a) P, (b) Si, (c) Cr, (d) Fe and (e) Kr**

ANSWER:-

(a) P with atomic number 15 with electronic configuration $[Ne]3s^2 3p^3$ contains 3 unpaired electrons.
(b) Si with atomic number 16 with electronic configuration $[Ne]3s^2 3p^2$ contains 2 unpaired electrons.
(c) Cr with atomic number 24 with electronic configuration $[Ar]3d^5 4s^1$ contains 6 unpaired electrons.
(d) Fe with atomic number 26 with electronic configuration $[Ar]3d^6 4s^2$ contains 4 unpaired electrons.
(e) Kr with atomic number 36$[Ar]3d^{10} 4s^2 4p^6$ contains 0 unpaired electrons

Q. 67. **(a)** **How many subshells are associated with n = 4?**
(b) **How many electrons will be present in the subshells having m_s value of –1/2 for n = 4?**

ANSWER:-

(a) Four sub-shells are associated with n = 4. These include 4s,4p,4d and 4f.
(b) Sixteen electrons will be present in the sub-shells having m_s value of –1/2 for n = 4.
Also, sixteen electrons will be present in the sub-shells having m_s value of +1/2 for n = 4.

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