

CH2 – STRUCTURE OF ATOM**Question 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 an electron = 9.1×10^{-28} g

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

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

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

Mass of 1 electron = 9.1×10^{-31} kg

Mass of 6.022×10^{23} electrons = $(9.1 \times 10^{-31} \text{ kg}) \times (6.022 \times 10^{23}) = 5.48 \times 10^{-7}$ kg

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

Charge on one mole electrons = $1.602 \times 10^{-19} \times 6.022 \times 10^{23} = 9.65 \times 10^4$ coulombs

Question 2.

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

(ii) Find (a) the total number and (b) the total mass of neutrons in 7 mg of ^{14}C . (Assume that 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 ?

Answer:

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

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

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

= 6.022×10^{24} electrons

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

Gram atomic mass of carbon (C-14) = 14 g = 14×10^3 mg

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

$$7 \text{ mg of carbon (C-14) have atoms} = \frac{6.022 \times 10^{23}}{(14 \times 10^3 \text{ mg})} \times (7 \text{ mg}) = 3.011 \times 10^{20} \text{ atoms.}$$

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

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

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

= 2.408×10^{21} neutrons

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

Mass of 2.408×10^{21} neutrons = $(1.675 \times 10^{-27} \text{ kg}) \times 2.408 \times 10^{21}$
= 4.033×10^{-6} kg.

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

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

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

$$\begin{aligned} 34 \text{ mg of } \text{NH}_3 \text{ have molecules} &= \frac{6.022 \times 10^{23}}{(17 \times 10^3 \text{ mg})} \times (34 \text{ mg}) \\ &= 1.2044 \times 10^{20} \text{ molecules.} \end{aligned}$$

Step II. Calculation of total number and mass of protons

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

No. of protons present in 1.2044×10^{20} molecules of NH_3 = $1.2044 \times 10^{20} \times 10$

= 1.2044×10^{22} protons

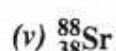
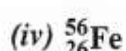
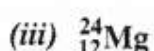
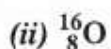
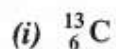
Mass of one proton = 1.67×10^{-27} kg

Mass of 1.2044×10^{22} protons = $(1.67 \times 10^{-27} \text{ kg}) \times 1.2044 \times 10^{22}$
= 2.01×10^{-5} kg.

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

Question 3.

How many protons and neutrons are present in the following nuclei



Answer:

- | | |
|---|-------------------------------------|
| (i) ${}^{13}_6\text{C}$; Atomic no. (Z) = 6 | Mass no. (A) = 13 |
| No. of protons (p) = 6 | No. of neutrons (n) = 13 - 6 = 7 |
| (ii) ${}^{16}_8\text{O}$; Atomic no. (Z) = 8 | Mass no. (A) = 16 |
| No. of protons (p) = 8 | No. of neutrons (n) = 16 - 8 = 8 |
| (iii) ${}^{24}_{12}\text{Mg}$; Atomic no. (Z) = 12 | Mass no. (A) = 24 |
| No. of protons (p) = 12 | No. of neutrons (n) = 24 - 12 = 12 |
| (iv) ${}^{56}_{26}\text{Fe}$; Atomic no (Z) = 26 | Mass no. (A) = 56 |
| No. of protons (p) = 26 | No. of neutrons (n) = 56 - 26 = 30. |
| (v) ${}^{88}_{38}\text{Sr}$; Atomic no (Z) = 38 | Mass no. (A) = 88 |
| No. of protons (p) = 38 | No. of neutrons (n) = 50. |

Question 4.

Write the complete symbol for the atom (X) 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) ${}^{35}_{17}\text{X}$ | (ii) ${}^{233}_{92}\text{X}$ | (iii) ${}^9_4\text{X}$ |
| (X = Cl) | (X = U) | (X = Be) |

Question 5.

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

Answer:

Step I. Calculation of frequency of yellow light

We know that

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

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

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

Step II. Calculation of wave number of yellow light

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

Question 6.

Calculate the energy of each of the photons which

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

(ii) have wavelength of 0.50 Å.

Answer:

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

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

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

Energy of photon (E) = $h\nu = hc\lambda$

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

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

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

Question 7.

Calculate the wavelength, frequency, and wavenumber of lightwave whose period is 2.0×10^{-10} s.

Answer:

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

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

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

Question 8.

What is the number of photons of light with wavelength 4000 pm which provide 1 Joule of energy?

Answer:

Energy of photon (E) = $hc\lambda$

$$h = 6.626 \times 10^{-34} \text{ J s}, c = 3 \times 10^8 \text{ m s}^{-1}, \lambda = 4000 \text{ pm} = 4000 \times 10^{-12} = 4 \times 10^{-9} \text{ m}$$

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

Now, $4.965 \times 10^{-17} \text{ J}$ is the energy of photon = 1

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

Question 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, (ii) the kinetic energy of the emission (iii) the velocity of the photoelectron. (Given $1 \text{ eV} = 1.6020 \times 10^{-19} \text{ J}$).

Answer:

(i) The energy of the photon.

$$\begin{aligned} \text{Energy (E)} &= \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^8 \text{ m s}^{-1})}{(4 \times 10^{-7} \text{ m})} = 4.97 \times 10^{-19} \text{ J} \\ &= \frac{(1 \text{ eV})}{(1.602 \times 10^{-19} \text{ J})} \times (4.97 \times 10^{-19} \text{ J}) = 3.1 \text{ eV} \end{aligned}$$

(ii) Kinetic energy of emission

$$\begin{aligned} \text{Kinetic energy of emission} &= E - \text{work function (i.e. kinetic energy of emitted electrons)} \\ &= (3.1 - 2.13) = 0.97 \text{ eV} \end{aligned}$$

(iii) Velocity of photoelectron

$$\begin{aligned} \text{KE of emission} &= \frac{1}{2}mv^2 = 0.97 \text{ eV} \\ &= 0.97 \times 1.602 \times 10^{-19} \text{ J} = 0.97 \times 1.602 \times 10^{-19} \text{ kg m}^2 \text{ s}^{-2} \\ \text{or } v^2 &= \frac{2 \times 0.97 \times 1.602 \times 10^{-19} (\text{kg m}^2 \text{ s}^{-2})}{(9.1 \times 10^{-31} \text{ kg})} = 0.34 \times 10^{12} \text{ m}^2 \text{ s}^{-2} \\ \text{or } v &= (0.34 \times 10^{12} \text{ m}^2 \text{ s}^{-2})^{1/2} = 0.583 \times 10^6 \text{ m s}^{-1} = 5.83 \times 10^5 \text{ m s}^{-1}. \end{aligned}$$

Question 10.

Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in k-J mol⁻¹.

Answer:

$$\begin{aligned} E &= hc/\lambda \\ \lambda &= 242 \text{ nm} = 242 \times 10^{-9} \text{ m}, \quad c = 3 \times 10^8 \text{ m s}^{-1}, \quad h = 6.626 \times 10^{-34} \text{ Js} \\ \therefore E &= \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^8 \text{ m s}^{-1})}{(242 \times 10^{-9} \text{ m})} = 0.0821 \times 10^{-17} \text{ J} \end{aligned}$$

$$\text{Ionisation energy per mol (E)} = \frac{(0.0821 \times 10^{-17} \text{ J}) \times (6.022 \times 10^{23} \text{ mol}^{-1})}{1000} = 494 \text{ kJ mol}^{-1}$$

Question 11.

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

Answer:Energy of one photon (E) = $h\nu = hc/\lambda$

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

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

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

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

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

Question 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:

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

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

Question 13.

What is the wavelength of the light emitted when the electron in a hydrogen atom undergoes transition from the energy level with $n = 4$ to energy level $n = 2$? What is the colour corresponding to this wavelength? (Given $R_H = 109678 \text{ cm}^{-1}$)

Answer:

According to Balmer formula,

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

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

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

Question 14.

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

Answer:

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

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

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

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

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

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

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

On comparing : $\frac{IE_1}{IE_5} = \frac{(2.18 \times 10^{-18} \text{ J atom}^{-1})}{(8.72 \times 10^{-20} \text{ J atom}^{-1})} = 25$

The energy required to remove an electron from first orbit in a hydrogen atom is 25 times the energy needed to remove an electron from fifth orbit.

Question 15.

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

Answer:

The maximum no. of emission lines = $n(n-1)/2 = 6(6-1)/2 = 3 \times 5 = 15$

The actual transitions which are taking place are as follows :

$n = 6 \text{ to } n = 1$	$n = 5 \text{ to } n = 1$	$n = 4 \text{ to } n = 1$	$n = 3 \text{ to } n = 1$	$n = 2 \text{ to } n = 1$
$6 \rightarrow 5$	$5 \rightarrow 4$	$4 \rightarrow 3$	$3 \rightarrow 2$	$2 \rightarrow 1$
$6 \rightarrow 4$	$5 \rightarrow 3$	$4 \rightarrow 2$	$3 \rightarrow 1$	
$6 \rightarrow 3$	$5 \rightarrow 2$	$4 \rightarrow 1$		
$6 \rightarrow 2$	$5 \rightarrow 1$			
$6 \rightarrow 1$				
(5 lines)	(4 lines)	(3 lines)	(2 lines)	(1 line)

Question 16.

(i) The energy associated with first orbit in hydrogen atom is $-2.17 \times 10^{-18} \text{ J atom}^{-1}$. What is the energy associated with the fifth orbit ?

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

Answer:

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

$$\frac{E_1}{E_2} = \left(\frac{n_2}{n_1} \right)^2 \quad \left[\because E_n \propto \frac{1}{n^2} \right]$$

According to available data : $n_1 = 1$, $E_1 = -2.17 \times 10^{-18} \text{ J atom}^{-1}$, $n_2 = 5$

$$\therefore \frac{(-2.17 \times 10^{-18} \text{ J atom}^{-1})}{E_2} = \left(\frac{5}{1} \right)^2 = 25$$

or $E_5 = \frac{(-2.17 \times 10^{-18} \text{ J atom}^{-1})}{25} = -8.77 \times 10^{-20} \text{ J atom}^{-1}$.

(ii) For hydrogen atom ; $r_n = 0.529 \times n^2 \text{ \AA}$
 $r_5 = 0.529 \times (5)^2 = 13.225 \text{ \AA} = 1.3225 \text{ nm}.$

Question 17.

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

Answer:

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

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

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

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

Answer:

Step I. Calculation of energy required

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

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

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

$$\begin{aligned} \therefore \text{Energy required } (\Delta E) &= E_5 - E_1 = \left(\frac{-2.18}{25} \times 10^{-18} \text{ J} \right) - \left(\frac{-2.18}{1} \times 10^{-18} \text{ J} \right) \\ &= 2.18 \times 10^{-18} \left(1 - \frac{1}{25} \right) \text{ J} \\ &= 2.18 \times 10^{-18} \times \frac{24}{25} = 2.09 \times 10^{-18} \text{ J} \end{aligned}$$

Step II. Calculation of wavelength of light emitted

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

$$\begin{aligned} \therefore \lambda &= \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34} \text{ J-s}) \times (3 \times 10^8 \text{ m s}^{-1})}{(2.09 \times 10^{-18} \text{ J})} \\ &= 9.50 \times 10^{-8} \text{ m} = 950 \text{ \AA}. \end{aligned} \quad (\because 1 \text{ \AA} = 10^{-10} \text{ m})$$

Question 19.

The electronic energy in hydrogen atom is given by $E_n (-2.18 \times 10^{-18} \text{ J}) / n^2$. 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:**Step I. Calculation of energy required**

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

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

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

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

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

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^8 \text{ ms}^{-1})}{(5.45 \times 10^{-19} \text{ J})}$$

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

Question 20.

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

Answer:

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

$$\text{Mass of electron (m)} = 9.1 \times 10^{-31} \text{ kg}$$

$$\text{Velocity of electron (v)} = 2.05 \times 10^7 \text{ m s}^{-1}$$

$$\text{Planck's constant (h)} = 6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}$$

$$\therefore \lambda = \frac{(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1})}{(9.1 \times 10^{-31} \text{ kg}) \times (2.05 \times 10^7 \text{ m s}^{-1})} = 3.55 \times 10^{-11} \text{ m.}$$

Question 21.

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

Answer:**Step I. Calculation of velocity of the electron**

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

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

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

Step II. Calculation of wavelength of the electron

According to de Broglie's equation,

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

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

Question 22.

Which of the following are iso-electronic species ?

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

Answer:

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

Question 23.

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

(ii) What are the atomic numbers of the elements whose outermost electronic configurations are represented by :

(a) $3s^1$ (b) $4p^3$ and (c) $3d^6$?

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

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

Answer:

(i) (a) $1s^2$

(b) $1s^2 2s^2 2p^6$

(c) $1s^2 2s^2 2p^6$

(d) $1s^2 2s^2 2p^6$.

(ii) (a) Na ($Z = 11$) has outermost electronic configuration = $3s^1$

(b) N ($Z = 7$) has outermost electronic configuration = $2p^3$

(c) Fe ($Z = 26$) has outermost electronic configuration = $3d^6$

(iii) (a) Li

(b) P

(c) Sc

Question 24.

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

Answer:

The lowest value of l where 'g' orbital can be present = 4

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

Question 25.

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

Answer:

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

Question 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:

No. of protons in a neutral atom = No. of electrons = 29

Electronic configuration = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$.

Question 27.

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

Answer:

$H_2^+ = \text{one}$; $H_2 = \text{two}$; $O_2^+ = 15$

Question 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 electron in 3rd orbital.

(iii) Which of the following orbitals are possible ?

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

Answer:

(i) For $n = 3$; $l = 0, 1$ and 2 .

For $l = 0$; $m_l = 0$

For $l = 1$; $m_l = +1, 0, -1$

For $l = 2$; $m_l = +2, +1, 0, -1, -2$

(ii) For an electron in 3rd orbital ; $n = 3$; $l = 2$; m_l can have any of the values $-2, -1, 0,$

$+1, +2$.

(iii) 1p and 3f orbitals are not possible.

Question 29.

Using s, p and d notations, describe the orbitals with following quantum numbers :

(a) $n = 1, l = 0$

(b) $n = 4, l = 3$

(c) $n = 3, l = 1$

(d) $n = 4, l = 2$

Answer:

(a) 1s orbital

(b) 4f orbital

(c) 3p orbital

(d) 4d orbital

Question 30.

From the following sets of quantum numbers, state which are possible.

Explain why the others are not possible.

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

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

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

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

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

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

Answer:

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

(ii) The set of quantum numbers is possible.

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

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

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

(vi) The set of quantum numbers is possible.

Question 31.

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

(a) $n = 4 ; m_s = -1/2$

(b) $n = 3, l = 0$.

Answer:

(a) For $n = 4$

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

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

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

(b) For $n = 3$

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

Question 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:

According to Bohr's theory,

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

or $2\pi r = \frac{nh}{mv}$ or $mv = \frac{nh}{2\pi r}$

According to de Broglie equation,

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

Comparing (i) and (ii),

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

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

Question 33.

Calculate the number of atoms present in :

(i) 52 moles of He

(ii) 52 u of He

(iii) 52 g of He.

Answer:

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

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

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

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

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

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

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

Question 34.

Calculate the energy required for the process :



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

Answer:

The expression for the ionisation energy atom :

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

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

For He^+ ion ($Z = 2$), $E_n = 2.18 \times 10^{-18} \times (2)^2 = 8.72 \times 10^{-18} \text{ J atom}^{-1}$ (one electron species)

Question .35.

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

Answer:

$$\text{Length of scale} = 20 \text{ cm} = 20 \times 10^7 \text{ nm} = 2 \times 10^8 \text{ nm}$$

$$\text{Diameter of carbon atom} = 0.15 \text{ nm}$$

$$\therefore \text{Number of carbon atoms which can be placed side by side in the scale} = \frac{(2 \times 10^8 \text{ nm})}{(0.15 \text{ nm})} = 1.33 \times 10^9$$

Question 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 of the arrangement = 2.4 cm

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

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

$$\text{Radius of each carbon atom} = \frac{1}{2} (1.2 \times 10^{-8}) = 6.0 \times 10^{-9} \text{ cm} = 0.06 \text{ nm}$$

Radius of each carbon atom = $12(1.2 \times 10^{-8}) = 6.0 \times 10^{-9} \text{ cm} = 0.06 \text{ nm}$

Question 37.

The diameter of zinc atom is 2.6 Å. Calculate :

(a) the radius of zinc atom in pm

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

Answer:

(a) Radius of zinc atom = $1.3 \text{ Å} = 1.3 \times 10^{-10} \text{ m} = 130 \times 10^{-12} \text{ m} = 130 \text{ pm}$

(b) Length of the scale = 1.6 cm = $1.6 \times 10^{10} \text{ pm}$

Diameter of zinc atom = 260 pm

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

Question 38.

A certain particle carries $2.5 \times 10^{-16} \text{ C}$ of static electric charge. Calculate the

number of electrons present in it.

Answer:

$$\begin{aligned}\text{Magnitude of charge } (q) &= 2.5 \times 10^{-16} \text{ C} \\ \text{Charge on one electron } (e) &= 1.602 \times 10^{-19} \text{ C} \\ \text{No. of electrons present} &= \frac{(2.5 \times 10^{-16} \text{ C})}{(1.602 \times 10^{-19} \text{ C})} = 1560\end{aligned}$$

Question 39.

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

Answer:

$$\begin{aligned}\text{Charge on oil droplet} &= -1.282 \times 10^{-18} \text{ C} \\ \text{Charge on an electron} &= -1.602 \times 10^{-19} \text{ C} \\ \text{Number of electrons} &= \frac{q}{e} = \frac{(-1.282 \times 10^{-18} \text{ C})}{(-1.602 \times 10^{-19} \text{ C})} = 8\end{aligned}$$

Question 40.

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

Answer:

We have studied that in Rutherford's experiment by using heavy metals like gold and platinum, a large number of α -particles suffered deflection while a very few had to retrace their path.

If a thin foil of lighter atoms like aluminium etc. be used in the Rutherford experiment, this means that the obstruction offered to the path of the fast moving α -particles will be comparatively quite less.

As a result, the number of α -particles deflected will be quite less and the particles which are deflected back will be negligible.

Question 41.

Symbols ${}^{79}_{35}\text{Br}$ and ${}^{79}\text{Br}$ can be written whereas symbols ${}^{35}_{79}\text{Br}$ and ${}_{35}\text{Br}$ are not accepted. Answer in brief.

Answer:

In the symbol ${}_Z^AX$ of an element :

A denotes the atomic number of the element

B 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 $^{79}35\text{Br}$ and ^{79}Br are accepted because atomic number of Br will remain 35 even if not mentioned. Symbol $^{35}79\text{Br}$ is not accepted because atomic number of Br cannot be 79 (more than the mass number = 35). Similarly, symbol ^{35}Br cannot be accepted because mass number has to be mentioned. This is needed to differentiate the isotopes of an element.

Question 42.

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

Answer:

An element can be identified by its atomic number only. Let us find the atomic number.

Let the number of protons = x

Number of neutrons = $x + [x \times 31.7100] = (x \times 0.317x)$

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

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

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

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

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

Question 43.

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

Answer:

Let the no. of electron in the ion = x

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

and the no. of neutrons = $x + x \times 11.1100 = 1.111x$

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

$(x - 1 + 1.111x)$

Given mass of the ion = 37

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

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

No. of electrons = 18 ; No. of protons = $18 - 1 = 17$

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

Symbol of the ion = $^{37}17\text{Cl}^-$

Question 44.

An ion with mass number 56 contains 3 units of positive charge and 30.4%

more neutrons than electrons. Assign symbol to the ion.

Answer:

Let the no. of electrons in the ion = x

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

and the no. of neutrons = $x + x \times 31.7100 = x + 0.304x$

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

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

$$\therefore 56 = (x + 3) + (x + 0.304x) \text{ or } 2.304x = 56 - 3 = 53$$

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

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

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

Question 45.

Arrange the following type of radiations in increasing order of wavelength :

- (a) radiation from microwave oven
- (b) amber light from traffic signal
- (c) radiation from FM radio
- (d) cosmic rays from outer space and
- (e) X-rays.

Answer:

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

Question 46.

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

Answer:

$$\text{Power of the laser (E)} = N h \nu = N h c / \lambda$$

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

Question 47.

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

- (a) frequency of emission (b) the distance travelled by this radiation in 30s
- (c) energy of quantum (d) number of quanta present if it produces 2 J of energy.

Answer:

$$\begin{aligned}
 (a) \quad \text{Frequency of emission } (\nu) &= \frac{c}{\lambda} = \frac{(3.0 \times 10^8 \text{ m s}^{-1})}{(616 \times 10^{-9} \text{ m})} = 4.87 \times 10^{14} \text{ s}^{-1} \\
 (b) \quad \text{Velocity of radiation } (c) &= 3.0 \times 10^8 \text{ m s}^{-1} \\
 \text{Distance travelled in 30s} &= (3.0 \times 10^8 \text{ m s}^{-1}) \times (30 \text{ s}) = 9.0 \times 10^9 \text{ m} \\
 (c) \quad \text{Energy of quanta } (E) &= h\nu = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^8 \text{ m s}^{-1})}{(616 \times 10^{-9} \text{ m})} \\
 &= 32.27 \times 10^{-20} \text{ J} \\
 (d) \quad \text{Number of quanta present in 2 J of energy} \\
 &= \frac{\text{Total energy}}{\text{Energy per quanta}} = \frac{(2 \text{ J})}{(32.27 \times 10^{-20} \text{ J})} = 6.2 \times 10^{18}
 \end{aligned}$$

Question 48.

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

Answer:

$$\begin{aligned}
 \text{Time duration } (t) &= 2 \text{ ns} = 2 \times 10^{-9} \text{ s} \\
 \text{Frequency } (\nu) &= \frac{1}{t} = \frac{1}{2 \times 10^{-9} \text{ s}} = \frac{10^9}{2} \text{ s}^{-1} \\
 \text{Energy of one photon, } E &= h\nu = (6.626 \times 10^{-34} \text{ Js}) \times (10^9/2 \text{ s}^{-1}) = 3.25 \times 10^{-25} \text{ J} \\
 \text{No. of photons} &= 2.5 \times 10^5 \\
 \therefore \text{Energy of source} &= 3.3125 \times 10^{-25} \text{ J} \times 2.5 \times 10^{15} = 8.28 \times 10^{-10} \text{ J}
 \end{aligned}$$

Question 49.

Life times 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:

$$\text{Time duration } (t) = 2 \text{ ns} = 2 \times 10^{-9} \text{ s}$$

$$\text{Frequency } (\nu) = \frac{1}{t} = \frac{1}{2 \times 10^{-9} \text{ s}} = \frac{10^9}{2} \text{ s}^{-1}$$

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

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

$$\therefore \text{Energy of source} = 3.3125 \times 10^{-25} \text{ J} \times 2.5 \times 10^{15} = 8.28 \times 10^{-10} \text{ J}$$

Question 50.

The longest wavelength doublet absorption transition is observed at 589

nm and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

Answer:

$$\lambda_1 = 589 \text{ nm} = 589 \times 10^{-9} \text{ m}$$

$$\nu_1 = \frac{c}{\lambda_1} = \frac{3 \times 10^8 \text{ m s}^{-1}}{589 \times 10^{-9} \text{ m}} = \frac{3000}{589} \times 10^{14} \text{ s}^{-1} = 5.0934 \times 10^{14} \text{ s}^{-1}$$

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

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

$$\begin{aligned} \Delta E &= E_1 - E_2 = \frac{hc}{\lambda_1} - \frac{hc}{\lambda_2} = hc \left[\frac{1}{\lambda_1} - \frac{1}{\lambda_2} \right] \\ &= (6.626 \times 10^{-34} \text{ Js} \times 3 \times 10^8 \text{ m s}^{-1}) \left[\frac{1}{589 \times 10^{-9} \text{ m}} - \frac{1}{589.6 \times 10^{-9} \text{ m}} \right] \\ &= \frac{19.878 \times 10^{-34} \times 10^8}{10^{-9}} \left[\frac{589.6 - 589}{589.6 \times 589} \right] \text{ J} \\ &= \frac{19.878 \times 10^{-17} \times 0.6}{589.6 \times 589} \text{ J} = 3.43 \times 10^{-22} \text{ J} \end{aligned}$$

Question 51.

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

Answer:

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

$$\text{Threshold frequency } (\nu_0) = \frac{E_0}{h} = \frac{1.9 \times 1.602 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ Js}} = 0.459 \times 10^{15} \text{ s}^{-1} = 4.59 \times 10^{14} \text{ s}^{-1}$$

$$\text{Threshold wavelength } (\lambda_0) = \frac{c}{\nu_0} = \frac{3 \times 10^8 \text{ m s}^{-1}}{4.59 \times 10^{14} \text{ s}^{-1}} = 0.6536 \times 10^{-6} \text{ m} = 653.6 \text{ nm} \approx 654 \text{ nm}$$

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

$$\begin{aligned} \text{Kinetic energy } \left(\frac{1}{2} m v^2 \right) &= E - E_0 = hc \left[\frac{1}{\lambda} - \frac{1}{\lambda_0} \right] \\ &= \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3 \times 10^8 \text{ m s}^{-1})}{10^{-9} \text{ m}} \times \left[\frac{1}{500} - \frac{1}{654} \right] \\ &= \frac{6.626 \times 3 \times 154}{500 \times 654} \times 10^{-34+8+9} = 9.36 \times 10^{-20} \text{ J} \\ \text{Velocity } (v) &= \sqrt{\frac{2 \times 9.36 \times 10^{-20} \text{ J}}{m}} = \sqrt{\frac{2 \times 9.36 \times 10^{-20} \text{ kg m}^2 \text{ s}^{-2}}{9.1 \times 10^{-31} \text{ kg}}} \\ &= \sqrt{2.057 \times 10^{11} \text{ m}^2 \text{ s}^{-2}} = \sqrt{20.57 \times 10^{10} \text{ m}^2 \text{ s}^{-2}} = 4.5356 \times 10^5 \text{ m s}^{-1} \end{aligned}$$

Question 52.

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

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

Answer:

$$\begin{array}{lll} \lambda(\text{m}) & : & 500 \times 10^{-9} \quad 450 \times 10^{-9} \quad 400 \times 10^{-9} \\ \nu(\text{m s}^{-1}) & : & 2.55 \times 10^5 \quad 4.35 \times 10^5 \quad 5.35 \times 10^5 \end{array}$$

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

According to photoelectric effect :

$$h(\nu - \nu_0) = \frac{1}{2} m \nu^2$$

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

Substituting the results of three experiments in the above equation :

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

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

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

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

$$\frac{(\lambda_0 - 450)}{450 \times \lambda_0} \times \frac{500 \times \lambda_0}{(\lambda_0 - 500)} = \frac{(4.35 \times 10^5)^2}{(2.55 \times 10^5)^2}$$

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

$$\text{or} \quad (\lambda_0 - 450) = 2.619 (\lambda_0 - 500) = 2.619 \lambda_0 - 1309.5$$

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

Question 53.

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

Answer:

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

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

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

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

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

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

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

Question 54.

If the photon of the wavelength 150 pm strikes an atom, one of its inner bound electrons is ejected out with a velocity of $1.5 \times 10^7 \text{ m s}^{-1}$. Calculate the energy with which it is bound to the nucleus.

Answer:

$$\lambda = 150 \text{ pm} = 150 \times 10^{-12} \text{ m} = 1.5 \times 10^{-10} \text{ m} ; v = 1.5 \times 10^7 \text{ m s}^{-1}$$

$$\begin{aligned} \text{K.E.} &= \frac{1}{2} m v^2 = \frac{1}{2} \times 9.1 \times 10^{-31} \text{ kg} \times (1.5 \times 10^7 \text{ m s}^{-1})^2 \\ &= \frac{9.1 \times 1.5 \times 1.5}{2} \times 10^{-31+14} \text{ J} \\ &= 10.2375 \times 10^{-17} \text{ J} = 1.02375 \times 10^{-16} \text{ J} \end{aligned}$$

$$\begin{aligned} E &= \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js})(3 \times 10^8 \text{ m s}^{-1})}{(1.5 \times 10^{-10} \text{ m})} \\ &= \frac{6.626 \times 3}{1.5} \times 10^{-34+8+10} \\ &= \frac{6.626 \times 3}{1.5} \times 10^{-16} \text{ J} = 13.252 \times 10^{-16} \text{ J} \end{aligned}$$

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

$$E_0 = E - \text{K.E.} = (13.252 - 1.024) \times 10^{-16} \text{ J}$$

$$= 12.228 \times 10^{-16} \text{ J} = \frac{12.228 \times 10^{-16}}{1.602 \times 10^{-19}} = 7.63 \times 10^3 \text{ eV}$$

Question 55.

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

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

Answer:

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

$$\lambda = 1285 \text{ nm} = 1285 \times 10^{-9} \text{ m} = 1.285 \times 10^{-6} \text{ m}$$

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

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

$$\frac{2.3346}{32.9} = \frac{1}{3^2} - \frac{1}{n^2} \quad \text{or} \quad 0.071 = \frac{1}{9} - \frac{1}{n^2}$$

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

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

Paschen series lies in infrared region of the spectrum.

Question 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:

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

$$r_1 = 1.3225 \text{ nm} = 1322.5 \text{ pm} = \frac{52.9 n_1^2}{Z} \quad \dots (i)$$

$$r_2 = 211.6 \text{ pm} = \frac{52.9 n_2^2}{Z} \quad \dots (ii)$$

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

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

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

$$\begin{aligned} \text{Now, Wave number } (\bar{\nu}) &= (1.097 \times 10^7 \text{ m}^{-1}) \times \left(\frac{1}{2^2} - \frac{1}{5^2} \right) = 1.097 \times 10^7 \times \frac{21}{100} \text{ m}^{-1} \\ &= 23.037 \times 10^5 \text{ m}^{-1} \end{aligned}$$

$$\text{Wavelength } (\lambda) = \frac{1}{\bar{\nu}} = \frac{1}{(23.037 \times 10^5 \text{ m}^{-1})} = 434 \times 10^{-9} \text{ m} \quad \text{or} \quad 434 \text{ nm}$$

It lies in the visible region of light.

Question 57.

Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of

biological molecules and other type of material. If the velocity of the electron in this microscope is $1.6 \times 10^6 \text{ m s}^{-1}$, calculate de Broglie wavelength associated with this electron.

Answer:

$$\lambda = \frac{h}{m \cdot v} = \frac{(6.626 \times 10^{-34} \text{ Js})}{(9.1 \times 10^{-31} \text{ kg}) \times (1.6 \times 10^6 \text{ m s}^{-1})}$$

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

Question 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:

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

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

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

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

$$= 0.494 \times 10^3 \text{ m s}^{-1} = 494 \text{ m s}^{-1}.$$

Question 59.

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

Answer:

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

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

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

Question 60.

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

Answer:

$$v = 4.37 \times 10^9 \text{ m s}^{-1}; m = 0.1 \text{ kg}$$

$$\lambda = \frac{h}{mv} = \frac{(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1})}{(0.1 \text{ kg}) \times (4.37 \times 10^9 \text{ m s}^{-1})} = \frac{6.626}{0.437} \times 10^{-34-9} \text{ m}$$

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

Question 61.

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

Answer:

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

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

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

$$\text{Actual momentum } (p) = \frac{h}{4\pi \times 0.05 \text{ nm}} = \frac{(6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1})}{4 \times 3.142 \times (5 \times 10^{-11} \text{ m})} = 1.055 \times 10^{-24} \text{ kg m s}^{-1}$$

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

Question 62.

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

(i) $n = 4, l = 2, m_l = -2, m_s = -1/2$

(ii) $n = 3, l = 2, m_l = 1, m_s = +1/2$

(iii) $n = 4, l = 1, m_l = 0, m_s = +1/2$

(iv) $n = 3, l = 2, m_l = -2, m_s = -1/2$

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

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

Answer:

The electrons may be assigned to the following orbitals :

(i) 4d

(ii) 3d

(iii) 4p

(iv) 3d

(v) 3p

(vi) 4p.

The increasing order of energy is :

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

Question 63.

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

Answer:

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

Question 64.

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

Answer:

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

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

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

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

Question 65.

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

Answer:

Configuration of the two elements are :

Al ($Z = 13$) : $[\text{Ne}]103s^23p^1$; Si ($Z = 14$) : $[\text{Ne}] 103s^23p^2$

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

Question 66.

Indicate the number of unpaired electrons in :

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

Answer:

(a) P ($z=15$) : $[\text{Ne}]103s^23p^3$; No. of unpaired electrons = 3

(b) Si ($z=14$) : $[\text{Ne}]103s^23p^2$; No. of unpaired electrons = 2

(c) Cr ($z=24$): $[\text{Ar}]184s^13d^5$; No. of unpaired electrons = 6

(d) Fe ($z=26$): $[\text{Ar}]184s^23d^6$; No. of unpaired electrons = 4

(e) Kr ($z=36$) : $[\text{Ar}]184s^23d^{10}4p^6$; No. of unpaired electrons = Nil.

Question 67.

(a) How many sub-shells are associated with $n = 4$?

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

Answer:

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

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

Each orbital can have an electron with $m_s = -1/2$ -'. Total no. of electrons with $m_s = -1/2$ is 16.



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