

NCERT SOLUTIONS

CLASS - 11th



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Class : 11th
Subject : Chemistry
Chapter : 2
Chapter Name : Structure Of Atom

Q2.1 (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 = $9.10939 \times 10^{-31} \text{ kg}$
 $\therefore$ Number of electrons that weigh $9.10939 \times 10^{-31} \text{ kg} = 1$
Number of electrons that will weigh 1g = $(1 \times 10^{-3} \text{ kg})$
 $= \frac{1}{9.10939 \times 10^{-31} \text{ kg}} \times (1 \times 10^{-3} \text{ kg})$
 $= 0.1098 \times 10^{-3+31}$
 $= 0.1098 \times 10^{28}$
 $= 1.098 \times 10^{27}$

(ii) Mass of one electron = $9.10939 \times 10^{-31} \text{ kg}$
Mass of one mole of electron = $9.10939 \times 10^{-31} \text{ kg}$
Mass of one mole of electron = $(6.022 \times 10^{23}) \times (9.10939 \times 10^{-31} \text{ kg})$
 $= 5.48 \times 10^{-7} \text{ kg}$
Charge on one electron = $1.6022 \times 10^{-19} \text{ coulomb}$
Charge on one mole of electron = $(1.6022 \times 10^{-19} \text{ C}) (6.022 \times 10^{23})$
 $= 9.65 \times 10^4 \text{ C}$

Page : 69 , Block Name : Exercise

Q2.2 (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 = $1.675 \times 10^{-27} \text{ kg}$).
(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) 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) (a) Number of atoms of ^{14}C in 1 mole = 6.023×10^{23}
since 1 atom of ^{14}C contains $(14 - 6)$ i.e., 8 neutrons, the number of neutrons in 14g of ^{14}C is $(6.023 \times 10^{23}) \times 8$. Or, 14g of ^{14}C contains $(6.022 \times 10^{23} \times 8)$ neutrons.

Number of neutrons in 7mg

$$= \frac{6.022 \times 10^{23} \times 8 \times 7 \text{mg}}{1400 \text{mg}}$$

$$= 2.4092 \times 10^{21}$$

(b) Mass of one neutron = $1.67493 \times 10^{-27} \text{kg}$

Mass of total neutron in 7g of ^{14}C

$$= (2.4092 \times 10^{21}) (1.67493 \times 10^{-27} \text{kg})$$

$$= 4.0352 \times 10^{-6} \text{kg}$$

(iii) (a) 1 mole of $\text{NH}_3 = \{1(14) + 3(1)\} \text{g of } \text{NH}_3$

$$= 17 \text{g of } \text{NH}_3$$

$$= 6.022 \times 10^{23} \text{ molecules of } \text{NH}_3$$

Total number of protons present in 1 molecule of NH_3

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

$$= 6.023 \times 10^{24}$$

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

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

$$= 6.023 \times 10^{24}$$

$\Rightarrow$ 17 g of NH_3 contains (6.023×10^{24}) protons

Number of protons in 34 mg of NH_3

$$= \frac{6.022 \times 10^{24} \times 34 \text{mg}}{17000 \text{mg}}$$

$$= 1.2046 \times 10^{22}$$

(b) Mass of one proton = $1.67493 \times 10^{-27} \text{kg}$

Total mass of protons in 34 mg of NH_3

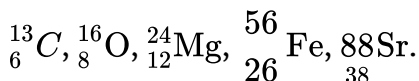
$$= (1.67493 \times 10^{-27} \text{kg}) (1.2046 \times 10^{22})$$

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

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.

Page : 69 , Block Name : Exercise

Q2.3 How many neutrons and protons are there in the following nuclei ?



Answer. ${}^{13}_6\text{C}$:

Atomic mass = 13

Atomic number = Number of protons = 6

Number of neutrons = (Atomic mass) - (Atomic number)

$$= 13 - 6 = 7$$

${}^{16}_8\text{O}$:

Atomic mass = 16

Atomic number = 8

Number of protons = 8

Number of neutrons = (Atomic mass)-(Atomic number)

$$= 16 - 8 = 8$$

$^{24}_{12}\text{Mg}$:

Atomic mass = 24

Atomic number = Number of protons = 12

Number of neutrons = (Atomic mass) - (Atomic number)

$$= 24 - 12 = 12$$

$^{56}_{26}\text{Fe}$:

Atomic mass = 56

Atomic number = Number of protons = 26

Number of neutrons = (Atomic mass) – (Atomic number)

$$= 56 - 26 = 30$$

$^{88}_{38}\text{Sr}$:

Atomic mass = 88

Atomic number = Number of protons = 38

Number of neutrons = (Atomic mass)-(Atomic number)

$$= 88 - 38 = 50$$

Page : 69 , Block Name : Exercise

Q2.4 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) $^{35}_{17}\text{Cl}$

(ii) $^{233}_{92}\text{U}$

(iii) ^9_4Be

Page : 69 , Block Name : Exercise

Q2.5 Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency (ν) and wavenumber ($\bar{\nu}$) of the yellow light.

Answer. From the expression,

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

We get,

$$\nu = \frac{c}{\lambda} \dots \dots \dots (1)$$

Where,

ν = frequency of yellow light

c = velocity of light in vacuum = $3 \times 10^8 \text{ m/s}$

λ = wavelength of yellow light = 580nm = 580×10^{-9} m

Substituting the values in expression(i):

$$v = \frac{3 \times 10^8}{580 \times 10^{-9}} = 5.17 \times 10^{14} \text{ s}^{-1}$$

Thus, frequency of yellow light emitted from the sodium lamp
= $5.17 \times 10^{14} \text{ s}^{-1}$

Wave number of yellow light, $\bar{\nu} = \frac{1}{\lambda}$
= $\frac{1}{580 \times 10^{-9}} = 1.72 \times 10^6 \text{ m}^{-1}$

Page : 69 , Block Name : Exercise

Q2.6 Find energy of each of the photons which

(i) correspond to light of frequency $3 \times 10^{15} \text{ Hz}$

(ii) have wavelength of 0.50 Å.

Answer. (i) Energy (E) of a photon is given by the expression,

$$E = h\nu$$

Where,

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

$$\nu = \text{frequency of light} = 3 \times 10^{15} \text{ Hz}$$

Substituting the values in the given expression of E:

$$E = (6.626 \times 10^{-34}) (3 \times 10^{15})$$

$$E = 1.988 \times 10^{-18} \text{ J}$$

(ii) Energy (E) of a photon having wavelength (λ) is given by the expression,

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

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

$$c = \text{velocity of light in vacuum} = 3 \times 10^8 \text{ m/s}$$

Substituting the values in the given expression of E:

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{0.50 \times 10^{-10}} = 3.976 \times 10^{-15} \text{ J}$$

$$\therefore E = 3.98 \times 10^{-15} \text{ J}$$

Page : 69 , Block Name : Exercise

Q2.7 Calculate the wavelength, frequency and wavenumber of a light wave whose period is $2.0 \times 10^{-10} \text{ s}$.

Answer. Frequency (ν) of light = 1/period

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

$$\text{Wavelength } (\lambda) \text{ of light} = \frac{c}{\nu}$$

Where,

$$c = \text{velocity of light in vacuum} = 3 \times 10^8 \text{ m/s}$$

Substituting the value in the given expression of λ :

$$\lambda = \frac{3 \times 10^8}{5.0 \times 10^9} = 6.0 \times 10^{-2} \text{ m}$$

Wave number ($\bar{\nu}$) of light = $\frac{1}{\lambda} = \frac{1}{6.0 \times 10^{-2}} = 1.66 \times 10^1 \text{ m}^{-1} = 16.66 \text{ m}$

Page : 69 , Block Name : Exercise

Q2.8 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 = $h\nu$

Energy (E_n) of 'n' photons = $n h \nu$

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

Where,

$$\lambda = \text{wavelength of light} = 4000 \text{ pm} = 4000 \times 10^{-12} \text{ m}$$

$$c = \text{velocity of light in vacuum} = 3 \times 10^8 \text{ m/s}$$

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

Substituting the values in the given expression of n:

$$n = \frac{(1) \times (4000 \times 10^{-12})}{(6.626 \times 10^{-34}) (3 \times 10^8)} = 2.012 \times 10^{16}$$

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

Page : 69 , Block Name : Exercise

Q2.9 A photon of wavelength $4 \times 10^{-7} \text{ m}$ strikes on metal surface, the work function of the metal being 2.13eV. Calculate (i) the energy of the photon (eV), (ii) the kinetic energy of the emission, and (iii) the velocity of the photoelectron ($1 \text{ eV} = 1.6020 \times 10^{-19} \text{ J}$).

Answer. (i) Energy (E) of a photon = $h\nu = \frac{hc}{\lambda}$

Where,

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

$$c = \text{velocity of light in vacuum} = 3 \times 10^8 \text{ m/s}$$

$$\lambda = \text{wavelength of photon} = 4 \times 10^{-7} \text{ m}$$

Substituting the values in the given expression of E:

$$E = \frac{(6.626 \times 10^{-34}) (3 \times 10^8)}{4 \times 10^{-7}} = 4.9695 \times 10^{-19} \text{ J}$$

Hence, the energy of the photon is $4.97 \times 10^{-19} \text{ J}$

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

$$= h\nu - h\nu_0$$

$$= (E - W) \text{ eV}$$

$$= \left(\frac{4.9695 \times 10^{-19}}{1.6020 \times 10^{-19}} \right) \text{ eV} - 2.13 \text{ eV}$$

$$= (3.1020 - 2.13) \text{ eV}$$

$$= 0.9720 \text{ eV}$$

Hence, the kinetic energy of emission is 0.97 eV.

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

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

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

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:

$$v = \sqrt{\frac{2 \times (0.9720 \times 1.6020 \times 10^{-10}) \text{ J}}{9.10939 \times 10^{-31} \text{ kg}}}$$

$$= \sqrt{0.3418 \times 10^{12} \text{ m}^2 \text{ s}^{-2}}$$

$$v = 5.84 \times 10^5 \text{ ms}^{-1}$$

Hence, the velocity of the photoelectron is $5.84 \times 10^5 \text{ ms}^{-1}$.

Page : 69 , Block Name : Exercise

Q2.10 Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJmol^{-1} .

$$\begin{aligned} \text{Answer. Energy of sodium (E)} &= \frac{N_A hc}{\lambda} \\ &= \frac{(6.023 \times 10^{23} \text{ mol}^{-1})(6.626 \times 10^{-34} \text{ Js})(3 \times 10^8 \text{ ms}^{-1})}{242 \times 10^{-9} \text{ m}} \\ &= 4.947 \times 10^5 \text{ Jmol}^{-1} \\ &= 494.7 \times 10^3 \text{ Jmol}^{-1} \\ &= 494 \text{ kJmol}^{-1} \end{aligned}$$

Page : 69 , Block Name : Exercise

Q2.11 A 25 watt bulb emits monochromatic yellow light of wavelength of $0.57 \mu\text{m}$. Calculate the rate of emission of quanta per second.

Answer. Power of bulb, $P = 25 \text{ Watt} = 25 \text{ Js}^{-1}$

Energy of one photon, $E = hv = \frac{hc}{\lambda}$

Substituting the values in the given expression of E:

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{(0.57 \times 10^{-6})} = 34.87 \times 10^{-20} \text{ J}$$

$$E = 34.87 \times 10^{-20} \text{ J}$$

Rate of emission of quanta per second

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

Page : 69 , Block Name : Exercise

Q2.12 Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength $6800 \text{ \AA}$. Calculate threshold frequency (ν_0) and work function (w_0) of the metal.

Answer. Threshold wavelength of radian (λ_0) = $6800 \text{ \AA} = 6800 \times 10^{-10} \text{ m}$

Threshold frequency (ν_0) of the metal

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

Thus, the threshold frequency (ν_0) of the metal is $4.41 \times 10^{14} \text{ s}^{-1}$

Hence, work function (W_o) of the metal = $h\nu_0$

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

$$= 2.922 \times 10^{-19} \text{ J}$$

Page : 69 , Block Name : Exercise

Q2.13 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 $n_i = 4$ and $n_f = 2$ transition will give rise to a spectral line of the Balmer series. The energy involved in the transition is given by the relation,

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

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 (λ) = $\frac{hc}{E}$

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

Substituting the values in the given expression of λ :

$$\lambda = \frac{(6.626 \times 10^{-34}) (3 \times 10^8)}{4.0875 \times 10^{-19}}$$

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

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

$$= 486 \text{ nm}$$

Page : 69 , Block Name : Exercise

Q2.14 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,

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

Where,

Z = atomic number of the atom

n = principal quantum number

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

$$\begin{aligned}\Delta E &= E_{\infty} - E_5 \\ &= \left[\left\{ \frac{-(2.18 \times 10^{-18} \text{ J})(1)^2}{(\infty)^2} \right\} - \left\{ \frac{-(2.18 \times 10^{-18} \text{ J})(1)^2}{(5)^2} \right\} \right] \\ &= (2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{(5)^2} \right) \left(\text{since } \frac{1}{\infty} = 0 \right) \\ &= 0.0872 \times 10^{-18} \text{ J}\end{aligned}$$

$$\Delta E = 8.72 \times 10^{-30} \text{ J}$$

Hence, the energy required for ionization from $n = 5$ to $n = \infty$ is $8.72 \times 10^{-20} \text{ J}$

Energy required for $n_1 = 1$ to $n = \infty$

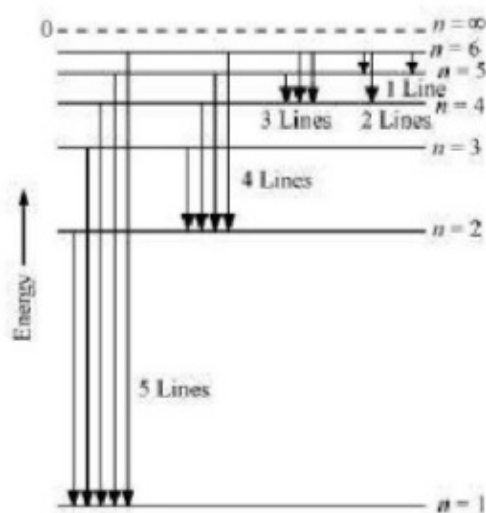
$$\begin{aligned}\Delta E &= E_{\infty} - E_1 \\ &= \left[\left\{ \frac{-(2.18 \times 10^{-18} \text{ J})(1)^2}{(\infty)^2} \right\} - \left\{ \frac{-(2.18 \times 10^{-18} \text{ J})(1)^2}{(1)^2} \right\} \right] \\ &= (2.18 \times 10^{-18}) [1 - 0] \\ &= 2.18 \times 10^{-18} \text{ J}\end{aligned}$$

Hence, less energy is required to ionize an electron in the 5th orbital of hydrogen atom as compared to that in the ground state.

Page : 70 , Block Name : Exercise

Q2.15 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 $n = 6$ drops to the ground state, the following transitions are possible:



Hence, a total number $(5 + 4 + 3 + 2 + 1)15$ lines will be obtained in the emission spectrum.

The number of spectral lines produced when an electron in the n^{th} level drops down to the ground state is given by $\frac{n(n-1)}{2}$

Given,

$$n = 6$$

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

- Q2.16 (i) The energy associated with the first orbit in the hydrogen atom is $-2.18 \times 10^{-18} \text{ J atom}^{-1}$. What is the energy associated with the fifth orbit?
 (ii) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

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

$$E_3 = \frac{-(2.18 \times 10^{-18})}{(5)^2} = \frac{-2.18 \times 10^{-18}}{25}$$

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

(ii) Radius of Bohr's n^{th} for hydrogen atom is given by,

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

For,

$$n = 5$$

$$r_5 = (0.0529 \text{ nm})(5)^2$$

$$r_5 = 1.3225 \text{ nm}$$

- Q2.17 Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

Answer. For the Balmer series, $n_i = 2$. Thus, the expression of wavenumber ($\bar{\nu}$) is given by,

$$\bar{\nu} = \left[\frac{1}{(2)^2} - \frac{1}{n_f^2} \right] (1,097 \times 10^7 \text{ m}^{-1})$$

Wave number ($\bar{\nu}$) is inversely proportional to wavelength of transition. Hence, for the longest wavelength transition, $\bar{\nu}$ has to be the smallest.

For $\bar{\nu}$ to be minimum, n_f should be minimum. For the Balmer series, a transition from $n_i = 2$ to $n_f = 3$ is allowed. Hence, taking $n_f = 3$, we get:

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

$$\bar{\nu} = (1.097 \times 10^7) \left[\frac{1}{4} - \frac{1}{9} \right]$$

$$= (1.097 \times 10^7) \left(\frac{9-4}{36} \right)$$

$$= (1.097 \times 10^7) \left(\frac{5}{36} \right)$$

$$\bar{\nu} = 1.5236 \times 10^6 \text{ m}^{-1}$$

- Q2.18 What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state? The ground state electron energy is $-2.18 \times 10^{-11} \text{ ergs}$.

Answer. Energy (E) on the n^{th} Bohr orbit of an atom is given by,

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

Where,

Z = atomic number of the atom

$$\text{Ground state energy} = -2.18 \times 10^{-11} \text{ ergs}$$

$$= -2.18 \times 10^{-11} \times 10^{-7} \text{ J}$$

$$= -2.18 \times 10^{-18} \text{ J}$$

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

$$\Delta E = E_5 - E_1$$

$$= \frac{-(2.18 \times 10^{-18})(1)^2}{(5)^2} - (-2.18 \times 10^{-18})$$

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

$$= (2.18 \times 10^{-18}) \left(\frac{24}{25} \right) = 2.0928 \times 10^{-18} \text{ J}$$

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

$$= \frac{(6.626 \times 10^{-34}) (3 \times 10^8)}{(2.0928 \times 10^{-18})}$$

$$= 9.498 \times 10^{-8} \text{ m}$$

Page : 70 , Block Name : Exercise

Q2.19 The electron energy in hydrogen atom is given by $E_m = (-2.18 \times 10^{-18}) / n^2 \text{ J}$. 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,

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

Energy required for ionization from $n = 2$ is given by,

$$\Delta E = E_\infty - E_2$$

$$= \left[\left(\frac{-2.18 \times 10^{-18}}{(\infty)^2} \right) - \left(\frac{-2.18 \times 10^{-18}}{(2)^2} \right) \right] \text{ J}$$

$$= \left[\frac{2.18 \times 10^{-18}}{4} - 0 \right] \text{ J}$$

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

$$\Delta E = 5.45 \times 10^{-19} \text{ J}$$

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

Here, λ is the longest wavelength causing the transition.

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

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

$$= 3647 \text{ \AA}$$

Page : 70 , Block Name : Exercise

Q2.20 Calculate the wavelength of an electron moving with a velocity of $2.05 \times 10^7 \text{ ms}^{-1}$.

Answer. According to de Broglie's equation,

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

Where,

λ = wavelength of moving particle

m = mass of particle

v = velocity of particle

h = Planck's constant

Substituting the values in the expression of λ :

$$\lambda = \frac{6.626 \times 10^{-34} \text{ Js}}{(9.10939 \times 10^{-31} \text{ kg})(2.05 \times 10^7 \text{ ms}^{-1})}$$

$$\lambda = 3.548 \times 10^{-11} \text{ m}$$

Hence, the wavelength of the electron moving with a velocity of $2.05 \times 10^7 \text{ ms}^{-1}$ is $3.548 \times 10^{-11} \text{ m}$.

Page : 70 , Block Name : Exercise

Q2.21 The mass of an electron is $9.1 \times 10^{-31} \text{ kg}$. If its K.E. is $3.0 \times 10^{-25} \text{ J}$. calculate its wavelength.

Answer. From de Broglie's equation,

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

Given,

Kinetic energy (K.E) of the electron = $3.0 \times 10^{-25} \text{ J}$

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

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

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

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

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

Substituting the value in the expression of λ :

$$\lambda = \frac{6.626 \times 10^{-34} \text{ Js}}{(9.10939 \times 10^{-31} \text{ kg})(811.579 \text{ ms}^{-1})}$$

$$\lambda = 8.9625 \times 10^{-7} \text{ m}$$

Hence, the wavelength of the electron is $8.9625 \times 10^{-7} \text{ m}$

Page : 70 , Block Name : Exercise

Q2.22 Which of the following are isoelectronic species i.e., those having the same number of electrons?

$\text{Na}^+, \text{K}^+, \text{Mg}^{2+}, \text{Ca}^{2+}, \text{s}^2, \text{Ar}$

Answer. Missing

Q2.23 (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) $[\text{He}]2s^1$

(b) $[\text{Ne}]3s^23p^3$

(c) $[\text{Ar}]4s^23d^1$

Answer. (i) (a) H^- ion

The electronic configuration of H atom is $1s^1$.

A negative charge on the species indicates the gain of an electron by it.

$\therefore$ Electronic configuration of $\text{H}^- = 1s^2$

(b) Na^+ ion

The electronic configuration of Na atom is $1s^22s^22p^63s^1$.

A positive charge on the species indicates the loss of an electron by it.

$\therefore$ Electronic configuration of $\text{Na}^+ = 1s^22s^22p^63s^0$ or $1s^22s^22p^6$

(c) O^{2-} ion

The electronic configuration of O atom is $1s^22s^22p^4$

A negative charge on the species indicates that two electrons are gained by it.

$\therefore$ Electronic configuration of O^{2-} ion = $1s^22s^22p^6$

(d) F^- ion

The electronic configuration of F atom is $1s^22s^22p^5$

A negative charge on the species indicates the gain of an electron by it.

$\therefore$ Electron configuration of F^- ion = $1s^22s^22p^6$

(ii) (a) $3s^1$

Completing the electron configuration of the element as

$1s^22s^22p^63s^1$

$\therefore$ Number of electrons present in the atom of the element

$= 2 + 2 + 6 + 1 = 11$

$\therefore$ Atomic number of the element = 11

(b) $2p^3$

Completing the electron configuration of the element as

$1s^22s^22p^3$

$\therefore$ Number of electrons present in the atom of the element = $2 + 2 + 3 = 7$

$\therefore$ Atomic number of the element = 7

(c) $3p^5$

Completing the electron configuration of the element as

$1s^2 2s^2 2p^5$.

$\therefore$ Number of electrons present in the atom of the element = $2 + 2 + 5 = 9$

$\therefore$ Atomic number of the element = 9

(iii) (a) $[\text{He}] 2s^1$

The electronic configuration of the element is $[\text{He}] 2s^1 = 1s^2 2s^1$.

$\therefore$ Atomic number of the element = 3

Hence, the element with the electronic configuration $[\text{He}] 2s^1$ is lithium (Li)

(b) $[\text{Ne}] 3s^2 3p^3$

The electronic configuration of the element is $[\text{Ne}] 3s^2 3p^3 = 1s^2 2s^2 2p^6 3s^2 3p^3$

$\therefore$ Atomic number of the element = 15

Hence, the element with the electronic configuration $[\text{Ne}] 3s^2 3p^3$ is phosphorus (P).

(c) $[\text{Ar}] 4s^2 3d^2$

The electronic configuration of the element is $[\text{Ar}] 4s^2 3d^1 = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$

Atomic number of the element = 21

Hence, the element with the electronic configuration $[\text{Ar}] 4s^2 3d^1$ is scandium (Sc).

Page : 70 , Block Name : Exercise

Q2.24 What is the lowest value of n that allows g orbitals to exist?

Answer. For g-orbitals, $L = 4$.

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

$\therefore$ For $l = 4$, minimum value of $n = 5$.

Page : 70 , Block Name : Exercise

Q2.25 An electron is in one of the 3d orbitals. Give the possible values of n, l and m_l for this electron.

Answer. For the 3d orbital:

Principal quantum number (n) = 3

Azimuthal quantum number (l) = 2

Magnetic quantum number (m_l) = - 2, -1, 0, 1, 2.

Page : 70 , Block Name : Exercise

Q2.26 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. (i) For an atom to be neutral, the number of protons is equal to the number of electrons.

$\therefore$ Number of protons in the atom of the given element = 29

(ii) The electronic configuration of the atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

Q2.27 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) = $1 + 1 = 2$

$\therefore$ Number of electrons in $\text{H}_2^+ = 2 - 1 = 1$

H_2 :

Number of electrons in $\text{H}_2 = 1 + 1 = 2$

O_2^+

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

Number of electrons in $\text{O}_2^+ = 16 - 1 = 15$

Q2.28 (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

(i) $n = 3$ (Given)

Answer. For a given value of n , l can have values from 0 to $(n - 1)$.

$\therefore$ For $n = 3$

$l = 0, 1, 2$

For a given value of l , m , can have $(2l + 1)$ values.

For $l = 0, m = 0$

$l = 1, m = -1, 0, 1$

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

$\therefore$ For $n = 3$

$l = 0, 1, 2$

$m_0 = 0$

$m_1 = -1, 0, 1$

$m_2 = -2, -1, 0, 1, 2$

(ii) For 3d orbital, $l = 2$

For a given value of l , m can have $(2l + 1)$ values i.e., 5 values.

$\therefore$ For $l = 2$

$m_2 = -2, -1, 0, 1, 2$

(iii) Among the given orbitals only 2s and 2p are possible. 1p and 3f cannot exist.

For p-orbital, $l = 1$.

For a given value of n , l can have values from zero to $(n - 1)$.

For l is equal to 1, the minimum value of n is 2.

Similarly,

For f-orbital, $l = 4$.

For $l = 4$, the minimum value of n is 5.

Hence, $1p$ and $3f$ do not exist.

Page : 70 , Block Name : Exercise

Q2.29 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. (a) $n = 1$, $l = 0$ (Given)

The orbital is $1s$

(b) For $n = 3$ and $l = 1$

The orbital is $3p$

(c) For $n = 4$ and $l = 2$

The orbital is $4d$.

(d) For $n = 4$ and $l = 3$

The orbital $4f$.

Page : 70 , Block Name : Exercise

Q2.30 Explain, giving reasons, which of the following sets of quantum numbers are not possible.

(a) $n = 0$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$

(b) $n = 1$, $l = 0$, $m_l = 0$, $m_s = -\frac{1}{2}$

(c) $n = 1$, $l = 1$, $m_l = 0$, $m_s = +\frac{1}{2}$

(d) $n = 2$, $l = 1$, $m_l = 0$, $m_s = -\frac{1}{2}$

(e) $n = 3$, $l = 3$, $m_l = -3$, $m_s = +\frac{1}{2}$

(f) $n = 3$, $l = 1$, $m_l = 0$, $m_s = +\frac{1}{2}$

Answer. (a) The given set of quantum numbers is not possible because the value Of the principal quantum number (n) cannot be zero.

(b) The given set of quantum numbers is possible.

(c) The given set of quantum numbers is not possible.

For a given value of n , ' l ' can have values from zero to $(n - 1)$

For $n = 1$, $l = 0$ and not 1.

(d) The given set of quantum numbers is possible.

(e) The given set of quantum numbers is not possible.

For $n = 3$,

$l = 0$ to $(3 - 1)$

$l = 0$ to 2 i.e., 0, 1, 2

(f) The given set of quantum numbers is possible.

Page : 70 , Block Name : Exercise

Q2.31 How many electrons in an atom may have the following quantum numbers?

(a) $n = 4$, $m_s = -\frac{1}{2}$

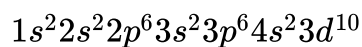
(b) $n = 3$, $l = 0$

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

$\therefore$ For $n = 4$

$$\begin{aligned}\text{Total number of electrons} &= 2(4)^2 \\ &= 32\end{aligned}$$

The given element has a fully filled orbital as



Hence, all the electrons are paired.

$\therefore$ Number of electrons (having $n = 4$ and $m_s = -\frac{1}{2}$) = 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.

Page : 71 , Block Name : Exercise

Q2.32 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. Since a hydrogen atom has only one electron, according to Bohr's postulate, the angular momentum of that electron is given by:

$$mvr = n \frac{h}{2\pi}, \dots \dots \dots (1)$$

Where,

$N = 1, 2, 3, \dots$

According to de Broglie's equation :

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

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

Substituting the value of 'mv' from expression (2) in expression (1):

$$\frac{hr}{\lambda} = n \frac{h}{2\pi}$$

$$\text{or } 2\pi r = n\lambda$$

Since ' $2\pi r$ ' represents the circumference of the Bohr orbit (r), it is proved by equation (3) that the circumference of the Bohr orbit of the hydrogen atom is an integral multiple of de Broglie's Wavelength associated with the electron revolving around the orbit.

Page : 71 , Block Name : Exercise

Q2.33 What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum ?

Answer. For He^+ ion, the wave number ($\bar{\nu}$) associated with the Balmer transition, $n = 4$ to $n = 2$ is given by :

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

Where,

$$n_1 = 2$$

$$n_2 = 4$$

Z = atomic number of helium

$$\bar{\nu} = \frac{1}{\lambda} = R(2)^2 \left(\frac{1}{4} - \frac{1}{16} \right)$$

$$= 4R \left(\frac{4-1}{16} \right)$$

$$\bar{\nu} = \frac{1}{\lambda} = \frac{3R}{4}$$

$$\Rightarrow \lambda = \frac{4}{3R}$$

According to the question, the desired transition for hydrogen will have the same wavelength as that of He^+

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

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

By hit and trial method, the equality given by equation (1) is true only when

$$n_1 = 1 \text{ and } n_2 = 2$$

$\therefore$ The transition for $n_2 = 2$ to $n = 1$ in hydrogen spectrum would have the same wavelength as Balmer transition $n = 4$ to $n = 2$ of (He) spectrum.

Page : 71 , Block Name : Exercise

Q2.34 Calculate the energy required for the process.



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

Answer. Energy associated with hydrogen-like species is given by,

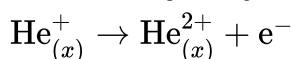
$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right) \text{ J}$$

For ground state of hydrogen atom

$$\Delta E = E_e - E_1$$

$$= 0 - \left[-2.18 \times 10^{-18} \left\{ \frac{(1)^2}{(1)^2} \right\} \right] \text{ J}$$

$$\Delta E = 2.18 \times 10^{-18} \text{ J}$$



An electron is removed from $n = 1$ to $n =$

$$\Delta E = E_m - E_1$$

$$= 0 - \left[-2.18 \times 10^{-18} \left\{ \frac{(2)^2}{(1)^2} \right\} \right]$$

$$\Delta E = 8.72 \times 10^{-18} \text{ J}$$

$\therefore$ The energy required for the process is $8.72 \times 10^{-18} \text{ J}$.

Page : 71 , Block Name : Exercise

Q2.35 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.

$$1\text{m} = 100\text{cm}$$

$$1\text{cm} = 10^{-2}\text{m}$$

$$\text{Length of the scale} = 20\text{cm}$$

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

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

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

$$\text{One carbon atom occupies } 0.15 \times 10^{-9}\text{m}.$$

$\therefore$ Number of carbon atoms that can be placed in a straight line

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

$$= 133.33 \times 10^7$$

$$= 1.33 \times 10^9$$

Page : 71 , Block Name : Exercise

Q2.36 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. Length of the given arrangement = 2.4 cm

Number of carbon atoms present = 2×10^8

$\therefore$ Diameter of carbon atom

$$= \frac{2.4 \times 10^{-1}\text{m}}{2 \times 10^8}$$

$$= 1.2 \times 10^{-16}\text{m}$$

$$\therefore \text{Radius of carbon atom} = \frac{\text{Diameter}}{2}$$

$$= 6.0 \times 10^{-11}\text{m}$$

Page : 71 , Block Name : Exercise

Q2.37 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.

$$\text{(a) Radius of zinc atom} = \frac{\text{Diameter}}{2}$$

$$= \frac{2.6\text{Å}}{2}$$

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

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

(b) Length of the arrangement = 1.6cm

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

Diameter of zinc atom = $2.6 \times 10^{-10} \text{m}$

∴ Number of zinc atoms present in the arrangement

$$= \frac{1.6 \times 10^{-10} \text{m}}{2.6 \times 10^{-10} \text{m}}$$

$$= 0.6153 \times 10^8$$

$$= 6.153 \times 10^7$$

Page : 71 , Block Name : Exercise

Q2.38 A certain particle carries $2.5 \times 10^{-16} \text{C}$ of static electric charge. Calculate the number of electrons present in it.

Answer. Charge on one electron = $1.6022 \times 10^{-19} \text{C}$

⇒ $1.6022 \times 10^{-19} \text{C}$ charge is carried by 1 electron.

∴ Number of electrons carrying a charge of $2.5 \times 10^{-16} \text{C}$

$$= \frac{1}{1.6022 \times 10^{-19} \text{C}} (2.5 \times 10^{-16} \text{C})$$

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

$$= 1560 \text{C}$$

Page : 71 , Block Name : Exercise

Q2.39 In Milikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the static electric charge on the oil drop is $-1.282 \times 10^{-18} \text{C}$ calculate the number of electrons present on it.

Answer.

Charge on the oil drop = $1.282 \times 10^{-18} \text{C}$

Charge on one electron = $1.6022 \times 10^{-19} \text{C}$

a Number of electrons present on the oil drop

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

$$= 0.8001 \times 10^{-18} \text{C}$$

$$= 8.0$$

Page : 71 , Block Name : Exercise

Q2.40 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. A thin foil of lighter atoms will not give the same results as given with the foil of heavier atoms.

Lighter atoms would be able to carry very little positive charge. Hence, they will not cause enough deflection of α -particles (positively charged).

Page : 71 , Block Name : Exercise

Q2.41 Symbols ${}^{79}_{35}\text{Br}$ and ${}^{79}\text{Br}$ can be written, whereas symbols ${}^{35}_{79}\text{Br}$ and ${}^{35}\text{Br}$ are not acceptable. Answer briefly.

Answer. The general convention of representing an element along with its atomic mass (A) and atomic number (Z) is ${}^A_Z\text{X}$.

Hence, ${}^{74}_{35}\text{Br}$ is acceptable ${}^{35}_{79}\text{Br}$ but is not acceptable.

${}^{79}\text{Br}$ can be written but ${}^{35}\text{Br}$ cannot be written because the atomic number of an element is constant, but the atomic mass of an element depends upon the relative abundance of its isotopes. Hence, it is necessary to mention the atomic mass of an element.

Page : 71 , Block Name : Exercise

Q2.42 An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

Answer. Let the number of protons in the element be x .

$\therefore$ Number of neutrons in the element

$$= x + 31.7\% \text{ of } x$$

$$= x + 0.317x$$

$$= 1.317x$$

According to the question,

Mass number of the element = 81

$$\therefore (\text{Number of protons} + \text{number of neutrons}) = 81$$

$$\Rightarrow x + 1.317x = 81$$

$$2.317x = 81$$

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

$$= 34.95$$

$$\therefore x = 35$$

Hence, the number of protons in the element i.e., x is 35.

Since the atomic number of an atom is defined as the number of protons present in its nucleus, the atomic number of the given element is 35.

The atomic symbol of the element is ${}^{81}_{35}\text{Br}$

Page : 71 , Block Name : Exercise

Q2.43 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 number of electrons in the ion carrying e negative charge be x. Then Number of neutrons present

$$= x + 11.1\% \text{ of } x$$

$$= x + 0.111x$$

$$= 1.111x$$

Number of electrons in the neutral atom = (x - 1)

(When an ion carries a negative charge, it carries an extra electron)

∴ Number of protons in the neutral atom = x - 1

Given, Mass number of the ion = 37

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

$$2.111x = 38$$

$$x = 18$$

The symbol of the ion is $^{17}\text{Cl}^-$

Page : 71 , Block Name : Exercise

Q2.44 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. Let the number of electrons present in ion

Number of neutrons in it = x + 30.4% of x = 1.304 x

Since the ion is tripositive,

Number of electrons in neutral atom = x + 3

Number of protons in neutral atom = x + 3

Given, Mass number of the ion = 56

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

$$2.304x = 53$$

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

$$x = 23$$

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

∴ The symbol of the ion $^{56}_{26}\text{Fe}^{3+}$

Page : 71 , Block Name : Exercise

Q2.45 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 as follows: Radiation from FM radio < amber light < radiation from microwave oven < X cosmic rays.

The increasing order of wavelength is as follows: rays < Cosmic rays < X-rays < radiation from microwave ovens < amber light < radiation of FM radio.

Page : 71 , Block Name : Exercise

Q2.46 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. Power of laser = Energy with which it emits photons = $(E = \frac{Mhc}{\lambda})$

Power Where,

N = number of photons emitted

h = Planck's constant

c = velocity of radiation

λ = wavelength of radiation

Substituting the values in the given expression Of Energy (E) :

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

$$= 0.3302 \times 10^7 \text{ J}$$

$$= 3.33 \times 10^6 \text{ J}$$

Hence, the power of the laser is $3.33 \times 10^6 \text{ J}$

Page : 71 , Block Name : Exercise

Q2.47 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 = $616 \times 10^{-9} \text{ m}$ (Given)

(a) Frequency of emission (v)

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

where,

c = velocity of radiation

λ = wavelength of radiation

Substituting the values in the given expression of

$$v = \frac{3.0 \times 10^8 \text{ m/s}}{616 \times 10^{-9} \text{ m}}$$

$$= 4.87 \times 10^8 \times 10^9 \times 10^{-3} \text{ s}^{-1}$$

$$V = 4.87 \times 10^{14} \text{ s}^{-1}$$

$$\text{Frequency of emission (v)} = 4.87 \times 10^{14} \text{ s}^{-1}$$

$$\text{(b) Velocity of radiation, (c) } = 3.0 \times 10^8 \text{ ms}^{-1}$$

Distance travelled by this radiation in 30s

$$= (3.0 \times 10^8 \text{ ms}^{-1}) (30 \text{ s})$$

$$= 9.0 \times 10^9 \text{ m}$$

$$\text{(c) Energy of quantum (E) } = h\nu$$

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

$$\text{Energy of quantum (E) } = 32.27 \times 10^{-20} \text{ J}$$

(d) Energy of one photon (quantum) = $32.27 \times 10^{-20} \text{ J}$

Therefore, $32.27 \times 10^{-20} \text{ J}$ of energy is present in 1 quantum.

Number of quanta in 2J of energy

$$\begin{aligned} &= \frac{2\text{J}}{32.27 \times 10^{-20} \text{ J}} \\ &= 6.19 \times 10^{18} \\ &= 6.2 \times 10^{18} \end{aligned}$$

Page : 71 , Block Name : Exercise

Q2.48 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. From the expression of energy of one photon (E),

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

Where,

λ = wavelength of radiation

h = Planck's constant

c = velocity of radiation

Substituting the values in the given expression of E:

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

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

Energy of one photon = $3.313 \times 10^{-19} \text{ J}$

Number of photons received with $3.15 \times 10^{-18} \text{ J}$ energy

$$= 9.15 \times 10^{-18} \text{ J}$$

$$\approx 10$$

Page : 72 , Block Name : Exercise

Q2.49 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. Frequency of radiation (ν),

$$\nu = \frac{1}{2.0 \times 10^{-9} \text{ s}}$$

$$\nu = 5.0 \times 10^8 \text{ s}^{-1}$$

Energy (E) of source $Nh\nu$

where, N = number of photons emitted

h = Planck's constant

ν = frequency of radiation Substituting the values in the given expression of (E) :

$$E = (2.5 \times 10^{15}) (6.626 \times 10^{-34} \text{ Js}) (5.0 \times 10^8)$$

$$E = 8.282 \times 10^{-10} \text{ J}$$

Hence, the energy of the source (E) is $8.282 \times 10^{-10} \text{ J}$.

Page : 72 , Block Name : Exercise

Q2.50 The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

Answer. Missing

Page : 72 , Block Name : Exercise

Q2.51 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. It is given that the work function (W_0) for caesium atom is 1.9 eV.

$$W_0 = \frac{hc}{\lambda_0}$$

(a) From the expression,

$$\lambda_0 = \frac{hc}{W_0}$$

Where,

λ_0 = threshold wavelength

h = Planck's constant

c = velocity of radiation

Substituting the values in the given expression of (λ_0):

$$\lambda_0 = \frac{(6.626 \times 10^{-34} \text{ Js}) (3.0 \times 10^8 \text{ ms}^{-1})}{1.9 \times 1.602 \times 10^{-19} \text{ J}}$$

$$\lambda_0 = 6.53 \times 10^{-7} \text{ m}$$

Hence, the threshold wavelength λ_0 is 653 nm.

(b) From the expression, $W_0 = h\nu_0$, we get:

$$\nu_0 = \frac{W_0}{h}$$

Where,

ν_0 = threshold frequency

h = Planck's constant

Substituting the values in the given expression of ν_0 .

$$\nu_0 = \frac{1.9 \times 1.602 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ Js}}$$

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

$$\nu_0 = 4.593 \times 10^{14} \text{ s}^{-1}$$

Hence, the threshold frequency of radiation (ν_0) is $4.593 \times 10^{14} \text{ s}^{-1}$

(c) According to the question:

Wavelength used in irradiation(λ) = 500 nm

Kinetic energy = $h(v - v_0)$

$$= hc \left(\frac{1}{\lambda} - \frac{1}{\lambda_0} \right)$$

$$= (6.626 \times 10^{-14} \text{ Js}) (3.0 \times 10^4 \text{ ms}^{-1}) \left(\frac{\lambda_0 - \lambda}{2\lambda_0} \right)$$

$$= (1.9878 \times 10^{-24} \text{ Jm}) \left[\frac{(653 - 500)10^{-9} \text{ m}}{(653)(500)10^{-18} \text{ m}^2} \right]$$

$$= \frac{(1.9878 \times 10^{-26}) (153 \times 10^9)}{(653)(500)} \text{ J}$$

$$= 9.3149 \times 10^{-20} \text{ J}$$

Kinetic energy of the ejected photoelectron = $9.3149 \times 10^{-20} \text{ J}$

since K. E = $\frac{1}{2}mv^2 = 9.3149 \times 10^{-20} \text{ J}$

$$v = 4.52 \times 10^5 \text{ ms}^{-1}$$

Hence, the velocity of the ejected photoelectron (v) is $4.52 \times 10^5 \text{ ms}^{-1}$.

Page : 72 , Block Name : Exercise

Q2.52 Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength and, (b) Planck's constant.

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

Answer.

(a) Assuming the threshold wavelength to be λ_0 nm ($= \lambda_0 \times 10^{-4} \text{ m}$), the kinetic energy of the radiation is given as:

$$h(v - v_0) = \frac{1}{2}mv^2$$

Three different equalities can be formed by the given value as:

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

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

$$\frac{hc}{10^{-9} \text{ m}} \left[\frac{1}{500} - \frac{1}{\lambda_0} \right] = \frac{1}{2}m (2.55 \times 10^{-3} \text{ ms}^{-1})^2$$

$$\frac{hc}{10^{-9} \text{ m}} \left[\frac{1}{450} - \frac{1}{\lambda_a} \right] = \frac{1}{2}m (3.45 \times 10^{-3} \text{ ms}^{-1})^2$$

$$\frac{hc}{10^{-9} \text{ m}} \left[\frac{1}{400} - \frac{1}{\lambda_v} \right] = \frac{1}{2}m (5.35 \times 10^{-3} \text{ ms}^{-1})^2$$

Dividing equation (3) by equation (1):

$$\frac{\left[\frac{\lambda_0 - 400}{400\lambda_0} \right]}{\left[\frac{\lambda_0 - 500}{500\lambda_0} \right]} = \frac{(5.35 \times 10^{-3} \text{ ms}^{-1})^2}{(2.55 \times 10^{-3} \text{ ms}^{-1})^2}$$

$$\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} = \left(\frac{5.35}{2.55} \right)^2 = \frac{28.6225}{6.5025}$$

$$\frac{5\lambda_0 - 2000}{4\lambda_{11} - 2000} = 4.40177$$

$$17.6070\lambda_0 - 5\lambda_0 = 8803.537 - 2000$$

$$\lambda_0 = \frac{6805.537}{12.607}$$

$$\lambda_0 = 539.8\text{nm}$$

$$\lambda_0 = 540\text{nm}$$

$\therefore$ Threshold wavelength (λ_0) = 540nm

Note: part (b) of the question is not done due to the incorrect values of velocity given in the question.

$$\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} = \left(\frac{5.35}{2.55} \right)^2 = \frac{28.6225}{6.5025}$$

$$\frac{5\lambda_0 - 2000}{4} = 4.40177$$

$$17.6070\lambda_0 - 5\lambda_0 = 8803.537 - 2000$$

$$\lambda_0 = \frac{6805.537}{12.607}$$

$$\lambda_0 = 539,8\text{nm}$$

$$\lambda_0 = 540\text{nm}$$

Page : 72 , Block Name : Exercise

Q2.53 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. From the principle of conservation of the energy of an incident photon (E) is equal to the sum of the work function (W_0) of radiation and its kinetic energy (K.E) i.e.,

$$E = W_0 + K.E$$

$$\Rightarrow W_0 = E - K.E$$

Energy of incident photon (E) Where,

c = velocity of radiation

h = Planck's constant

λ = wavelength of radiation

Substituting the values in the given expression of E:

$$E = \frac{(6.626 \times 10^{-34} \text{Js}) (3.0 \times 10^8 \text{ms}^{-1})}{256.7 \times 10^{-9} \text{m}}$$

$$= 7.744 \times 10^{-19} \text{J}$$

$$= \frac{7.744 \times 10^{-19}}{1.602 \times 10^{-19}} \text{eV}$$

$$E = 4.83 \text{eV}$$

The potential applied to silver metal changes to kinetic energy (K.E) Of the photoelectron. Hence,

$$KE = 0.35V$$

$$K.E = 0.35eV$$

$$\therefore \text{Work function, } W_0 = E - K.E$$

$$= 4.83eV - 0.35eV$$

$$= 4.48eV$$

$$\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} = \left(\frac{5.35}{2.55} \right)^2 = \frac{28.6225}{6.5025}$$

$$\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} = 4.40177$$

$$17.6070\lambda_0 - 5\lambda_0 = 8803.537 - 2000$$

$$\lambda_0 = \frac{6805.537}{12.607}$$

$$\lambda_0 = 539.8nm$$

$$\lambda_0 = 540nm$$

Page : 72 , Block Name : Exercise

Q2.54 If the photon of the wavelength 150 pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of $1.5 \times 10^7 \text{ms}^{-1}$, calculate the energy with which it is bound to the nucleus.

Answer.

Energy of incident photon (E) is given by,

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

$$= \frac{(6.626 \times 10^{-34} \text{Js})(3.0 \times 10^8 \text{ms}^{-1})}{(150 \times 10^{-12} \text{m})}$$

$$= 1.3252 \times 10^{-15} \text{J}$$

$$= 13.252 \times 10^{-16} \text{J}$$

Energy of the electron ejected (K.E)

$$= \frac{1}{2} m_e v^2$$

$$= \frac{1}{2} (9.10939 \times 10^{-31} \text{kg}) (1.5 \times 10^7 \text{ms}^{-1})^2$$

$$= 10.2480 \times 10^{-17} \text{J}$$

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

Hence, the energy with which the electron is bound to the nucleus can be obtained as:

$$= E - K.E$$

$$= 13.252 \times 10^{-16} \text{J} - 1.025 \times 10^{-16} \text{J}$$

$$\begin{aligned}
&= 12.227 \times 10^{-16} \text{ J} \\
&= \frac{12.227 \times 10^{-16}}{1.602 \times 10^{-19}} \text{ eV} \\
&= 7.6 \times 10^3 \text{ eV} \\
\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} &= \left(\frac{5.35}{2.55} \right)^2 = \frac{28.6225}{6.5025} \\
\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} &= 4.40177 \\
17.6070\lambda_0 - 5\lambda_0 &= 8803.537 - 2000 \\
\lambda_0 &= \frac{6805.537}{12.607} \\
\lambda_0 &= 539.8 \text{ nm} \\
\lambda_0 &= 540 \text{ nm}
\end{aligned}$$

Page : 72 , Block Name : Exercise

Q2.55 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.

$$\begin{aligned}
&\text{Wavelength of transition} = 1285 \text{ nm} \\
\text{Answer. } &= 1285 \times 10^{-9} \text{ m (Given)} \\
\nu &= 3.29 \times 10^{15} \left(\frac{1}{3^2} - \frac{1}{n^2} \right) \text{ (Given)} \\
\nu &= \frac{c}{\lambda} \\
&\text{since} \\
&= \frac{3.0 \times 10^8 \text{ ms}^{-1}}{1285 \times 10^{-9} \text{ m}} \\
\nu &= 2.33 \times 10^{14} \text{ m} \\
&\text{Substituting the value of } \nu \text{ in the given expression,} \\
3.29 \times 10^{15} \left(\frac{1}{9} - \frac{1}{n^2} \right) &= 2.33 \times 10^{14} \\
\frac{1}{9} - \frac{1}{n^2} &= \frac{2.33 \times 10^{14}}{3.29 \times 10^{15}} \\
\frac{1}{9} - 0.7082 \times 10^{-1} &= \frac{1}{n^2} \\
\Rightarrow \frac{1}{n^2} &= 1.1 \times 10^{-1} - 0.7082 \times 10^{-1} \\
\frac{1}{n^2} &= 4.029 \times 10^{-2} \\
n &= \sqrt{\frac{1}{4.029 \times 10^{-2}}} \\
n &= 4.98 \\
n &\approx 5
\end{aligned}$$

Hence, for the transition to be observed at 1285 nm, $n = 5$

The spectrum lies in the infrared region.

Q2.56 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. The radius of the n^{th} orbit of hydrogen-like particles is given by,

$$r = \frac{0.529n^2}{Z} \text{ \AA}$$

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

For radius (r_1) = 1.3225nm

$$= 1.32225 \times 10^{-9} \text{ m}$$

$$= 1322.25 \times 10^{-12} \text{ m}$$

$$= 1322.25 \text{ pm}$$

Q2.57 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 $1.6 \times 10^6 \text{ ms}^{-1}$, calculate de Broglie wavelength associated with this electron.

Answer. From de Broglie's equation,

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

$$\lambda = \frac{6.626 \times 10^{-34} \text{ Js}}{(9.10939 \times 10^{-31} \text{ kg}) (1.6 \times 10^6 \text{ ms}^{-1})}$$

$$= 4.55 \times 10^{-10} \text{ m}$$

$$\lambda = 455 \text{ pm}$$

$\therefore$ de Broglie's wavelength associated with the electron is 455pm

Q2.58 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.

From de Broglie's equation,

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

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

Where,

v = velocity of particle (neutron)

h = Planck's constant

m = mass of particle (neutron)

λ = wavelength

Substituting the values in the expression of velocity (v) ,

$$v = \frac{6.626 \times 10^{-34} \text{ Js}}{(1.67493 \times 10^{-27} \text{ kg})(800 \times 10^{-12} \text{ m})}$$

$$n_1^{\frac{1}{2}} = \frac{r_1 Z}{52.9}$$

$$n_1^2 = \frac{1322.25 Z}{52.9}$$

Similarly,

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

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

$$\frac{n_1^2}{n_2^2} = 6.25$$

$$\frac{n_1}{n_2} = 2.5$$

$$\frac{n_1}{n_2} = \frac{25}{10} = \frac{5}{2}$$

$$\Rightarrow r_1 = 5 \text{ and } n_2 = 2$$

Thus, the transition is from the 5th orbit to the 2nd orbit. It belongs to the Balmer series.

Wave number ($\bar{\nu}$) for the transition is given by,

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

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

$$= 2.303 \times 10^6 \text{ m}^{-1}$$

$\therefore$ Wavelength (λ) associated with the emission transition is given by,

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

$$= \frac{1}{2.303 \times 10^6 \text{ m}^{-1}}$$

$$= 0.434 \times 10^{-6} \text{ m}$$

$$\lambda = 434 \text{ nm}$$

Page : 72 , Block Name : Exercise

Q2.59 If the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6 \text{ ms}^{-1}$. calculate the de Broglie wavelength associated with it.

Answer. According to de Broglie's equation,

Where,

λ = wavelength associated with the electron

h = Planck's constant

m = mass of electron

V = velocity of electron

Substituting the values in the expression of λ :

$$\begin{aligned} &= \frac{6.626 \times 10^{-34} \text{Js}}{(9.10939 \times 10^{-31} \text{kg}) (2.19 \times 10^6 \text{ms}^{-1})} \\ &= 3.32 \times 10^{-10} \text{m} = 3.32 \times 10^{-10} \text{m} \times \frac{100}{100} \\ &\lambda = 332 \times 10^{-12} \text{m} \end{aligned}$$

$\therefore$ Wavelength associated with the electron = 332pm

Page : 72 , Block Name : Exercise

Q2.60 The velocity associated with a proton moving in a potential difference of 1000 V is $4.37 \times 10^5 \text{ms}^{-1}$. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

Answer.

According to de Broglie's expression,

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

Substituting the values in the expression,

$$\lambda = \frac{6.626 \times 10^{-34} \text{Js}}{(0.1 \text{kg}) (4.37 \times 10^5 \text{ms}^{-1})}$$

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

Page : 72 , Block Name : Exercise

Q2.61 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 $h/4\pi_m \times 0.05 \text{nm}$, is there any problem in defining this value.

Answer. From Heisenberg's uncertainty principle,

$$\Delta x \times \Delta p = \frac{h}{4\pi} \Rightarrow \Delta p = \frac{1}{\Delta x} \cdot \frac{h}{4\pi}$$

Where,

Δx = uncertainty in position of the electron

Δp = uncertainty in momentum of the electron

Substituting the values in the expression of Δp :

$$\begin{aligned} \Delta p &= \frac{1}{0.002 \text{nm}} \times \frac{6.626 \times 10^{-34} \text{Js}}{4 \times (3.14)} \\ &= \frac{1}{2 \times 10^{-12} \text{m}} \times \frac{6.626 \times 10^{-34} \text{Js}}{4 \times 3.14} \\ &= 2.637 \times 10^{-23} \text{Jsm}^{-1} \end{aligned}$$

$$\Delta p = 2.637 \times 10^{-23} \text{kgms}^{-1} \left(1\text{J} = 1\text{kgms}^2\text{s}^{-1} \right)$$

$$\therefore \text{Uncertainty in the momentum of the electron} = 2.637 \times 10^{-23} \text{kgms}^{-1}$$

$$= \frac{h}{4\pi m \times 0.05 \text{nm}}$$

$$= \frac{6.626 \times 10^{-34} \text{Js}}{4 \times 3.14 \times 5.0 \times 10^{-11} \text{m}}$$

$$= 1 \times 3.14 \times 5.0 \times 10^{-11} \text{m}$$

Since the magnitude of the actual momentum is smaller than the uncertainty, the value cannot be defined.

Page : 72 , Block Name : Exercise

Q2.62 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. For $n = 4$ and $l = 2$, the orbital occupied is 4d.

For $n = 3$ and $l = 2$, the orbital occupied is 3d.

For $n = 4$ and $l = 1$, the orbital occupied is 4p.

Hence, the six electrons i.e., 1, 2, 3, 4, 5, and 6 are present in the 4d, 3d, 4p, 3d, 3p, and 4p orbitals respectively.

Therefore, the increasing order of energies is $5(3p) < 2(3d) = 4(3d) < 3(4p) = 6(4p) < 1(4d)$.

Page : 72 , Block Name : Exercise

Q2.63 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. Nuclear charge experienced by an electron (present in a multi-electron atom) is dependant upon the distance between the nucleus and the orbital, in which the electron is present. As the distance increases, the effective nuclear charge also decreases. Among p-orbitals, 4p orbitals are farthest from the nucleus of bromine atom with (+35) charge. Hence, the electrons in the 4p orbital will experience the lowest effective nuclear charge. These electrons are shielded by electrons present in the 20 and 30 orbitals along with the s-orbitals. Therefore, they will experience the longest nuclear charge.

Page : 73 , Block Name : Exercise

Q2.64 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. Nuclear charge is defined as the net positive charge experienced by an electron in the orbital of a multi-electron atom. The closer the orbital, the greater is the nuclear charge experienced by the electron (s) in it.

(i) The electron(s) present in the 2s orbital will experience greater nuclear charge (being closer to the nucleus) than the electron(s) in the 3s orbital.

(ii) 4d will experience greater nuclear charge than 4f since 4d is closer to the nucleus.

(iii) 3p will experience greater nuclear charge since it is closer to the nucleus than 3f.

Page : 73 , Block Name : Exercise

Q2.65 The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus ?

Answer. Nuclear charge is defined as the net positive charge experienced by an electron in a multi-electron atom.

The higher the atomic number, the higher is the nuclear charge. Silicon has 14 protons while aluminum has 13 protons. Hence, silicon has a larger nuclear charge of (+14) than aluminium, Which has a nuclear charge of (+13). Thus, the electrons in the 3p orbital of silicon will experience more effective nuclear charge than aluminium.

Page : 73 , Block Name : Exercise

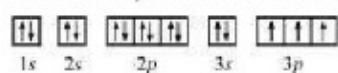
Q2.66 Indicate the number of unpaired electrons in : (a) P, (b) Si, (c) Cr, (d) Fe and (e) Kr.

Answer. (a) Phosphorus (P):

Atomic number = 15

The electronic configuration of p is: $1s^2 2s^2 2p^6 3s^2 3p^3$

The orbital picture of P can be represented as:



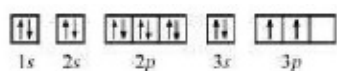
From the orbital picture, phosphorus has three unpaired electrons.

(b) Silicon (Si):

Atomic number = 14

The electronic configuration of Si is:

The orbital Picture of Si can be represented as:



From the orbital picture, silicon has two unpaired electrons.

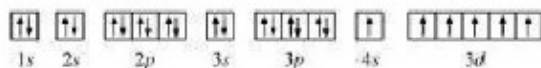
(c) Chromium (Cr):

Atomic number = 24

The electronic configuration of Cr is:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

The orbital picture of chromium is:

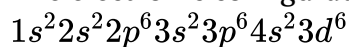


From the orbital picture, chromium has six unpaired electrons.

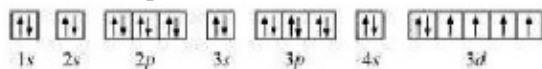
(d) Iron (Fe):

Atomic number = 26

The electronic configuration is:



The orbital picture of Iron is:

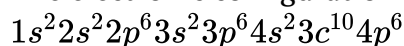


From the orbital picture, iron has four unpaired electrons.

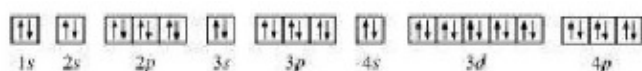
(e) Krypton (Kr):

Atomic number = 36

The electronic configuration is:



The orbital picture of krypton is:



Since all orbitals are fully occupied, there are no unpaired electrons in krypton.

Page : 73 , Block Name : Exercise

Q2.67 (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) $n = 4$ (Given)

For a given value of 'n', I can have values from zero to $(n - 1)$.

Thus, four subshells associated with $n = 4$, which are s, p, d and f.

(b) Number of orbitals in the n^{th} shell $= n^2$

For $n = 4$

Number Of orbitals = 16

If each orbital is taken fully, then it Will have 1 electron With m_s value of $-\frac{1}{2}$

$\therefore$ Number of electrons with m_s value of $\left(-\frac{1}{2}\right)$ is 16.

Page : 73 , Block Name : Exercise

