

## Chapter 2: Structure of Atom

NCERT · Class 11 · Chemistry

### Textbook exercises

#### 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.

#### Solution

Ans: Mass of one electron =

$$9.10939 \times 10^{-31} kg$$

Number of electrons that weigh

$$9.10939 \times 10^{-31} kg$$

= 1

Number of electrons that will weigh 1 g =

$$1 \times 10^{-3} kg$$

$$\frac{1}{9.10939 \times 10^{-31} kg} (1 \times 10^{-3} kg) = 0.1098 \times 10^{-3+31}$$

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

Ans: Mass of one electron =

$$9.10939 \times 10^{-31} kg$$

Mass of one mole of electron =

$$(6.022 \times 10^{23}) \times (9.10939 \times 10^{-31} kg)$$

=

$$5.48 \times 10^{-7} kg$$

Charge on one electron =

$$1.6022 \times 10^{-19} coulomb$$

Charge on one mole of electron =

$$(1.6022 \times 10^{-19})(6.022 \times 10^{23})$$

=

$$9.65 \times 10^4 C$$

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## Question 2

(i). Calculate the Total Number of Electrons Present in One Mole of Methane.

(ii). a) Find the Total Number of Neutrons in 7mg of  $^{14}C$ . (Assume the Mass of a Neutron =  $1.675 \times 10^{-27} kg$ )

(b) Find the Total Mass of Neutrons in 7mg of  $^{14}C$ . (Assume the Mass of a Neutron =  $1.675 \times 10^{-27} kg$ )

(iii). (a) Find the Total Number of Protons in 34mg of  $NH_3$  at STP.

(b) Find the total mass of protons in 34mg of  $NH_3$  at STP. Will the answer change if the temperature and pressure changed?

## Solution

Ans: Number of electrons present in 1 molecule of methane (



)=

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

Number of electrons present in 1 mole i.e.,

$$6.023 \times 10^{23}$$

molecules of methane =

$$6.023 \times 10^{23} \times 10 = 6.023 \times 10^{24}$$

Ans: Number of atoms of



in 1 mole =

$$6.023 \times 10^{23}$$

Since 1 atom of



contains (14-6) i.e., 8 neutrons, the number of neutrons in 14g of



is

$$(6.023 \times 10^{23}) \times 8$$

Number of neutrons in 7 mg

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

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

Ans: Mass of one neutron =

$$1.67493 \times 10^{-27} \text{ kg}$$

Mass of total neutrons in 7g of

$$^{14}\text{C}$$

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

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

Ans: 1 mole of



= {1(14) + 3(1)} g of



= 17g of



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

Molecules of



Total number of protons present in 1 molecule of



$$= [1(7) + 1(3)] = 10$$

Number of protons in

$$6.023 \times 10^{23}$$

molecules of



$$= (6.023 \times 10^{23})(10) = 6.023 \times 10^{24}$$

17g of



contains

$$(6.023 \times 10^{24})$$

protons.

Number of protons in 34mg of



$$= \frac{6.022 \times 10^{24} \times 34mg}{1700mg}$$

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

Ans: Mass of one proton =

$$1.67493 \times 10^{-27} kg$$

Total mass of protons in 34mg of



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

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

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

### Question 3

How Many Neutrons and Protons are There in the Following Nuclei?



### Solution

Ans:  ${}^{13}_6C$  :

Atomic mass =13

Atomic number = Number of Protons = 6

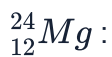
Number of neutrons = (Atomic Mass)-(Atomic Number)=13-6=7

${}^{16}_8O$  :

Atomic mass =16

Atomic number = Number of Protons = 8

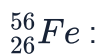
Number of neutrons = (Atomic Mass)-(Atomic Number) =16-8=8



Atomic mass = 24

Atomic number = Number of Protons = 12

Number of neutrons = (Atomic Mass)-(Atomic Number)=24-12 =12



Atomic mass = 56

Atomic number = Number of Protons = 26

Number of neutrons = (Atomic Mass)-(Atomic Number) = 56-26=30



Atomic mass = 88

Atomic number = Number of Protons = 38

Number of neutrons = (Atomic Mass)-(Atomic Number)= 88-38=50

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#### Question 4

Write the Complete Symbol for the Atom with the Given Atomic Number (Z) and Atomic mass (A)

a. Z = 17, A = 35

b. Z = 92, A = 233

c. Z = 4, A = 9

#### Solution

Ans:  ${}_{17}^{35}\text{Cl}$

Ans:  ${}_{92}^{233}\text{U}$

Ans:  ${}_4^9\text{Be}$

### Question 5

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

### Solution

Ans:

From the expression,

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

We get,

$$\nu = \frac{c}{\lambda} \text{ ----- (i)}$$

Where,

$\nu$  = frequency of yellow light

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

$\lambda$  = wavelength of yellow light =  $580 \text{ nm} = 580 \times 10^{-9} \text{ m}$

Substituting the values in the expression (i)

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

Thus, frequency of yellow light emitted from sodium lamp

$$= 5.17 \times 10^{14} \text{ S}^{-1}$$

Wave number of yellow light,

$$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{580 \times 10^{-9}m} = 1.72 \times 10^6 m^{-1}$$

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### Question 6

(i). Find energy of each of the photons which correspond to light of frequency  $3 \times 10^{15} Hz$

(ii). Find energy of each of the photons which have wavelength of  $0.50A^\circ$

### Solution

Ans: Energy (E) of a photon is given by the expression,

$$E = h\nu$$

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

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

Substituting the values in the given expression of Energy,

$$E = (6.626 \times 10^{-34})(3 \times 10^{15}) = 1.988 \times 10^{-18} J$$

Ans: Energy (E) of a photon having wavelength ( $\lambda$ ) is given by the expression,

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

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

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

$\lambda$  = wavelength of yellow light =  $0.50A^\circ = 0.5 \times 10^{-10} m$

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{0.5 \times 10^{-10}} J$$

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

### Question 7

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

### Solution

Ans: Frequency ( $\nu$ ) of light =  $\frac{1}{\text{period}} = \frac{1}{2 \times 10^{-10} s} = 5 \times 10^9 s^{-1}$

Wavelength of light,

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

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

Substitution the value in the given expression  $\lambda$ ,

$$\lambda = \frac{3 \times 10^8}{5 \times 10^9} = 6.0 \times 10^{-2} m$$

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

**Question 8**

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

**Solution**

Ans: Energy (E) of a photon is given by the expression,

$$E = h\nu$$

Energy of 'n' photons,

$$E_n = nh\nu$$

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

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

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

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

Substituting the values in the given expression '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  $2.012 \times 10^{16}$

**Question 9**

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

(i). Calculate the energy of photon (eV)

(ii). Calculate the kinetic energy of the emission.

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

### Solution

Ans: Energy of photon,  $E = h\nu = \frac{hc}{\lambda}$

Where,  $c$  = velocity of light in vacuum =  $3 \times 10^8 m/s$

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

$\lambda$  = wavelength of yellow light =  $4 \times 10^{-7} m$

Substituting the values in the given expression of  $E$ ,

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

Hence, the energy of photon is  $4.9695 \times 10^{-19} J$

Ans: the kinetic energy of emission is given by,

$$E_k = h\nu - h\nu_0 = (E - W)eV$$

$$= \left[ \frac{4.9696 \times 10^{-19}}{1.6020 \times 10^{-19}} \right] - 2.13eV = 0.9720eV$$

Hence, the kinetic energy of emission is 0.97eV

Ans: The velocity of a photoelectron ( $v$ ) can be calculated by the expression,

$$\frac{1}{2}mv^2 = h\nu - h\nu_0$$

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

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:

$$v = \sqrt{\frac{2(0.9720 \times 1.6020 \times 10^{-19})}{9.10939 \times 10^{-31} \text{ kg}}} \text{ J}$$

$$\Rightarrow v = \sqrt{0.3418 \times 10^{12} \text{ m}^2 \text{ s}^{-2}} = 5.84 \times 10^5 \text{ ms}^{-1}$$

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

### Question 10

Electromagnetic radiation of wavelength 242 nm is just sufficient to ionize the sodium atom. Calculate the ionization energy of sodium in  $\text{kJmol}^{-1}$

### Solution

Ans: Energy of sodium,  $E = \frac{N_A hc}{\lambda}$

$$E = \frac{(6.023 \times 10^{23} \text{ mol}^{-1})(6.626 \times 10^{-34} \text{ Js})(3 \times 10^8 \text{ ms}^{-1})}{242 \times 10^{-9} \text{ m}} = 4.947 \times 10^5 \text{ Jmol}^{-1}$$

Hence, the ionization energy of sodium is  $494 \text{ kJmol}^{-1}$

**Question 11**

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

**Solution**

Ans: power of Bulb,  $P = 25 \text{ Watt} = 25 J s^{-1}$

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

Substituting the values in the given expression of E,

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

Rate of emission of quanta per second,

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

**Question 12**

Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength  $6800 \text{Å}$ . Calculate threshold frequency ( $\nu_0$ ) and the work function ( $w_0$ ) of the metal.

**Solution**

Ans: Threshold wavelength of radian ( $\lambda_0$ ) =  $6800 \text{Å} = 6800 \times 10^{10} m$

Threshold frequency of metal ( $\nu_0$ ) =  $\frac{c}{\lambda_0} = \frac{3 \times 10^8 m s^{-1}}{6.8 \times 10^{-7} m} = 4.41 \times 10^{14} s^{-1}$

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

Hence, the work function of the metal is,

$$w_0 = h\nu_0 = (6.626 \times 10^{-34} \text{ Js})(4.41 \times 10^{14} \text{ s}^{-1}) = 2.922 \times 10^{-19} \text{ J}$$

### Question 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  $n = 2$ ?

### Solution

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

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

Substituting the values in the given expression of E,

$$E = 2.18 \times 10^{-18} \left[ \frac{1}{n_i^2} - \frac{1}{n_f^2} \right] = 2.18 \times 10^{-18} \left[ \frac{1}{4^2} - \frac{1}{2^2} \right] = -4.0875 \times 10^{-19} \text{ J}$$

The negative sign indicates the energy of emission.

Wavelength of light emitted,

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

Substituting the values in the given expression of

$$\lambda$$

:

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

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

$$\lambda = 486 nm$$

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### Question 14

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

### Solution

Ans: The expression of energy is given by,

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

Where,  $Z$  = atomic number of atom

$n$  = principal quantum number

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

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

$$\Rightarrow \Delta E = \left[ \left( \frac{(-2.18 \times 10^{-18} J)(1^2)}{\infty^2} \right) - \left( \frac{(-2.18 \times 10^{-18} J)(1^2)}{5^2} \right) \right] = 2.18 \times 10^{-18} J$$

Hence, the energy required for ionization from

$$n_1 = 5$$

to

$$n_2 = \infty$$

is

$$8.72 \times 10^{20} J$$

The energy required for ionization from

$$n_1 = 1$$

to

$$n_2 = \infty$$

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

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

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

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### Question 15

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

### Solution

Ans: 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 of  $(5 + 4 + 3 + 2 + 1)$  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  $\frac{n(n-1)}{2}$

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

### Question 16

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

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

### Solution

Ans: Energy associated with the fifth orbit of hydrogen atom is calculated as:

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

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

Ans: Radius of Bohr's  $n^{th}$  orbit for hydrogen atom is given by,

$$r_n = (0.0529nm)n^2$$

For  $n = 5$ ,

$$r_5 = (0.0529nm) \times 5^2$$

$$r_5 = 1.3225nm$$

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**Question 17**

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

**Solution**

Ans: For the Balmer series,  $n_i = 2$ . Thus, the expression of wave number ( $\bar{\nu}$ ) is given by

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

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:

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

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

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

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

### Solution

Ans: Energy (E) of the  $n^{\text{th}}$  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:

$$\Delta E = E_5 - E_1$$

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

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

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

$$= \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{(2.0928 \times 10^{-18})} = 9.498 \times 10^{-8} m$$

### Question 19

The electron energy in hydrogen atom is given by  $E_n = \frac{(-2.18 \times 10^{-18})}{n^2} J$ . Calculate the energy required to remove an electron completely from the  $n = 2$  orbit. What is the longest wavelength of light in cm that can be used to cause this transition?

### Solution

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

Given,

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

$$\Delta E = \left( \frac{(-2.18 \times 10^{-18})}{\infty^2} \right) - \left( \frac{(-2.18 \times 10^{-18})}{2^2} \right) = 0.545 \times 10^{-18} J$$

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

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

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

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

, here  $\lambda$  is the longest wavelength causing the transition.

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

$$\lambda = 3647 \text{Å}$$

### Question 20

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

### Solution

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

Where,  $\lambda$  = wavelength of moving particle

$m$  = mass of particle,  $v$  = velocity of particle,  $h$  = Planck's constant

Substituting the values in the expression of  $\lambda$ :

$$\lambda = \frac{6.626 \times 10^{-34} \text{Js}}{(9.10939 \times 10^{-31} \text{kg})(2.05 \times 10^7 \text{ms}^{-1})} = 3.548 \times 10^{-11} m$$

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

### Question 21

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

### Solution

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

Given, the K.E. of electrons is  $3.0 \times 10^{-25} J$ .

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

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

$$\Rightarrow v = \sqrt{\frac{2(3.0 \times 10^{-25} J)}{9.10939 \times 10^{-31} kg}} = \sqrt{6.5866 \times 10^4}$$

$v = 811.579 ms^{-1}$

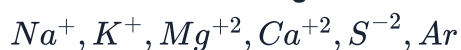
Substituting the value in the expression of  $\lambda$  :

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

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

### Question 22

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



### Solution

Ans: Isoelectronic species have the same number of electrons.

Number of electrons in sodium (Na) = 11

Number of electrons in ( $Na^+$ ) = 10

A positive charge denotes the loss of an electron.

Similarly,

Number of electrons in  $K^+ = 18$

Number of electrons in  $Mg^{+2} = 10$

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

A negative charge denotes the gain of an electron by a species.

Number of electrons in sulphur (S) = 16

$\therefore$  Number of electrons in  $S^{-2} = 18$

Number of electrons in argon (Ar) = 18

Hence, the following are isoelectronic species:

(1)  $Na^{+}$  and  $Mg^{+2}$  (10 electrons each)

(2)  $K^{+}$ ,  $Ca^{+2}$ ,  $S^{-2}$  and Ar (18 electrons each).

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### Question 23

(i). Write the Electronic Configurations of the Following ions:

(a)  $H^{-}$  ion

(b)  $Na^{+}$  ion

(c)  $O^{-2}$  ion

(d)  $F^{-}$  ion

(a)  $3s^1$

(b)  $2p^3$

(c)  $3p^5$

(a)  $[He]2s^1$

(b)  $[Ne]3s^23p^3$

(c)  $[Ar]4s^23d^1$

### Solution

Ans: The electronic configuration of H atom is  $1s^1$

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

∴ Electronic configuration of  $H^-$  is  $1s^2$

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

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

∴ Electronic configuration of  $Na^+$  is  $1s^2 2s^2 2p^6 3s^0$  or  $1s^2 2s^2 2p^6$

Ans: The electronic configuration of O atom is  $1s^2 2s^2 2p^4$

A negative charge on the species indicates the gain of two electrons by it.

∴ Electronic configuration of  $O^{2-}$  is  $1s^2 2s^2 2p^6$

Ans: The electronic configuration of F atom is  $1s^2 2s^2 2p^5$

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

∴ Electronic configuration of  $F^-$  is  $1s^2 2s^2 2p^6$

(ii). What are the Atomic Numbers of Elements Whose Outermost Electrons are Represented by

Ans: Completing the electron configuration of the element as  $1s^2 2s^2 2p^6 3s^1$

∴ Number of electrons present in the atom of the element

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

∴ Atomic number of the element = 11

Ans: Completing the electron configuration of the element as  $1s^2 2s^2 2p^3$

∴ Number of electrons present in the atom of the element

$$= 2 + 2 + 3 = 7$$

∴ Atomic number of the element = 7

Ans: Completing the electron configuration of the element as  $1s^2 2s^2 2p^6 3s^2 3p^5$

∴ Number of electrons present in the atom of the element

$$= 2 + 2 + 6 + 2 + 5 = 17$$

∴ Atomic number of the element = 17

(iii). Which Atoms are Indicated by the Following Configurations?

Ans: The electronic configuration of the element is  $[He]2s^1 = 1s^2 2s^1$

∴ Atomic number of the element = 3

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

Ans: The electronic configuration of the element is  $[Ne]3s^2 3p^3 = 1s^2 2s^2 2p^6 3s^2 3p^3$ .

∴ Atomic number of the element = 15

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

Ans: The electronic configuration of the element is  $[Ar]4s^2 3d^1 = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ .

∴ Atomic number of the element = 21

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

---

### Question 24

**What is the Lowest Value of n That Allows G Orbitals To Exist?**

#### Solution

Ans: 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).

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

---

### Question 25

**An Electron Is in One of the 3D Orbitals. Give the Possible Values of n, l and m, for This Electron.**

#### Solution

Ans: For the 3d orbital:

Principal quantum number ( $n$ ) = 3

Azimuthal quantum number ( $l$ ) = 2

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

---

### Question 26

(i). An Atom of an Element Contains 29 Electrons and 35 Neutrons. Deduce the Number of Protons.

(II). An Atom of an Element Contains 29 Electrons and 35 Neutrons. Deduce the Electronic Configuration of the Given Element.

### Solution

Ans: For an atom to be neutral, the number of protons is equal to the number of electrons.

∴ Number of protons in the atom of the given element = 29

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

The name of the element is Copper  ${}_{29}^{35}\text{Cu}$

---

### Question 27

Give the Number of Electrons in the Species  $H_2^+$ ,  $H_2$  and  $O_2^+$

### Solution

Ans:  $H_2^+$ :

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

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

$H_2$ :

Number of electrons in  $H_2 = 1 + 1 = 2$

$O_2^+$ :

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

$\therefore$  Number of electrons in  $O_2^+ = 16 - 1 = 15$

---

### 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 electrons for 3d orbital.

### Solution

Ans:  $n = 3$  (Given)

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  $l$ ,  $m_l$  can have  $(2l + 1)$  values.

$\therefore$  For  $l = 0$ ,  $m = 0$

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

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

$\therefore$  For  $n = 3$

$l = 0, 1, 2$

$m_0 = 0$

$m_1 = -1, 0, 1$

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

Ans: For 3d orbital,  $l = 2$ .

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

$\therefore$  For  $l = 2$

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

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

1p, 2s, 2p and 3f

Ans: 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  $(n - 1)$ .

Therefore 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.

---

### Question 29

Using s, p, d notations, describes the orbital with the following quantum numbers.

**A.**

$$n = 1, l = 0;$$

**B.**

$$n = 3; l = 1$$

**C.**

$$n = 4; l = 2;$$

**D.**

$$n = 4; l = 3.$$

### Solution

Ans:  $n = 1, l = 0$  (Given) the orbital is 1s.

Ans: For  $n = 3$  and  $l = 1$  the orbital is 3p.

Ans: For  $n = 4$  and  $l = 2$  the orbital is 4d.

Ans: For  $n = 4$  and  $l = 3$  the orbital is 4f.

### Question 30

Explain, Giving Reasons, Which of the Following Sets of Quantum Numbers are Not

|          |                           |                           |                              |                                        |
|----------|---------------------------|---------------------------|------------------------------|----------------------------------------|
| <b>A</b> | <b><math>n = 0</math></b> | <b><math>l = 0</math></b> | <b><math>m_l = 0</math></b>  | <b><math>m_s = +\frac{1}{2}</math></b> |
| <b>B</b> | <b><math>n = 1</math></b> | <b><math>l = 0</math></b> | <b><math>m_l = 0</math></b>  | <b><math>m_s = -\frac{1}{2}</math></b> |
| <b>C</b> | <b><math>n = 1</math></b> | <b><math>l = 1</math></b> | <b><math>m_l = 0</math></b>  | <b><math>m_s = +\frac{1}{2}</math></b> |
| <b>D</b> | <b><math>n = 2</math></b> | <b><math>l = 1</math></b> | <b><math>m_l = 0</math></b>  | <b><math>m_s = -\frac{1}{2}</math></b> |
| <b>E</b> | <b><math>n = 3</math></b> | <b><math>l = 3</math></b> | <b><math>m_l = -3</math></b> | <b><math>m_s = +\frac{1}{2}</math></b> |
| <b>F</b> | <b><math>n = 3</math></b> | <b><math>l = 1</math></b> | <b><math>m_l = 0</math></b>  | <b><math>m_s = +\frac{1}{2}</math></b> |

### Solution

Ans:

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

**Question 31**

**How Many Electrons in an Atom May Have the Following Quantum Numbers?**

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

b)  $n = 3, l = 0$

**Solution**

Ans:

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

∴ For  $n = 4$ ,

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

The given element has a fully filled orbital as  $1s^2 2s^2 3s^2 3p^6 4s^2 3d^{10}$

Hence, all the electrons are paired.

$$\therefore \text{Number of electrons (having } n = 4 \text{ and } m_s = -\frac{1}{2} \text{)} = 16$$

Ans:  $n = 3, l = 0$  indicates that the electrons are present in the 3s orbital. Therefore, the number of electrons having  $n = 3$  and  $l = 0$  is 2.

**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.**

**Solution**

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

$$mvr = n \frac{h}{2\pi} \text{ ----- (i)}$$

Where,  $n = 1, 2, 3 \dots$

According to de Broglie's equation:

$$\lambda = \frac{h}{mv} \text{ Or } mv = \frac{h}{\lambda} \text{ ----- (ii)}$$

Substituting the value of 'mv' from expression (ii) in expression (i):

$$\frac{hr}{\lambda} = n \frac{h}{2\pi} \text{ ----- (iii)}$$

Since  $2\pi r$  represents the circumference of the Bohr orbit (r), it is proved by equation (iii) 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.

### Question 33

**What Transition in the Hydrogen Spectrum Would have the Same Wavelength as the Balmer Transition  $n = 4$  to  $n = 2$  of  $He^+$  spectrum?**

### Solution

Ans: for  $He^+$  ion, the wave number  $\bar{\nu}$  associated with the Balmer transition,  $n=4$  to  $n=2$  given by:

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

Where,  $n_1 = 2$  and  $n_2 = 4$ ,  $Z$  = atomic number of Helium

$$\bar{\nu} = \frac{1}{\lambda} = R2^2 \left( \frac{1}{2^2} - \frac{1}{4^2} \right) = \frac{3R}{4}$$

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

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

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

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

----- (1)

By hit and trial method, the equality given by equation (1) is true only when  $n_1 = 1$  and  $n_2 = 2$

The transition for  $n_2 = 2$  to  $n_1 = 1$  in hydrogen spectrum would have the same wavelength as Balmer transition  $n = 4$  to  $n = 2$  of  $He^+$  spectrum.

### Question 34

Calculate the energy required for the process  $He^+_{(g)} \rightarrow He_{(g)}^{2+} + e^-$

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

### Solution

Ans: Energy associated with hydrogen-like species is given by,

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

For ground state of hydrogen atom,

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

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

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

For the given process,  $He^+_{(g)} \rightarrow He_{(g)}^{2+} + e^-$

An electron is removed from  $n = 1$  to  $n = \infty$ .

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

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

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

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

### Question 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 a Length of Scale of Length 20 cm Long.

### Solution

Ans: 1 m = 100 cm

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

$$\text{Length of the scale} = 20 \text{ cm} = 20 \times 10^{-2} m$$

$$\text{Diameter of a carbon atom} = 0.15 \text{ nm} = 0.15 \times 10^{-9} m$$

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

Number of carbon atoms that can be placed in a straight line =

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

### Question 36

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

### Solution

Ans: Length of the given arrangement = 2.4 cm

Number of carbon atoms present =  $2 \times 10^8$

$$\therefore \text{Diameter of carbon atom} = \frac{2.4 \times 10^{-2}}{2 \times 10^8} = 1.2 \times 10^{-10}m$$

$$\text{Radius of carbon atom} = \frac{\text{diameter}}{2} = \frac{1.2 \times 10^{-10}m}{2} = 6 \times 10^{-11}m$$

### Question 37

The diameter of the zinc atom is  $2.6 \text{ \AA}$ .

a) Calculate radius of zinc atom in pm

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

### Solution

$$\text{Ans: Radius of zinc atom} = \frac{\text{diameter}}{2} = \frac{2.6 \text{ \AA}}{2} = 1.3 \times 10^{-11}m = 13 \times 10^{-12}m = 13pm$$

$$\text{Ans: Length of the arrangement} = 1.6 \text{ cm} = 1.6 \times 10^{-2}m$$

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

Number of zinc atoms present in the arrangement =

$$\frac{1.6 \times 10^{-2}m}{2.6 \times 10^{-10}m} = 0.6153 \times 10^8m$$

---

**Question 38**

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

**Solution**

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

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

charge is carried by 1 electron.

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

$$= 1 \left( \frac{2.5 \times 10^{-16}C}{1.6022 \times 10^{-19}} \right) = 1.560 \times 10^3C = 1560C$$

---

**Question 39**

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

**Solution**

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

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

$$\therefore \text{Number of electrons present on the oil drop} = \frac{1.282 \times 10^{-18}C}{1.6022 \times 10^{-19}C} = 0.8001 \times 10^1 = 8.0$$

---

### Question 40

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

### Solution

Ans: 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  $\alpha$ -particles (positively charged).

---

### 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 acceptable. Answer briefly.

### Solution

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

Hence,  ${}^{79}_{35}\text{Br}$  is acceptable, but  ${}^{35}_{79}\text{Br}$  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.

---

### Question 42

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

### Solution

Ans: Let the number of protons in the element be  $x$ .

∴ Number of neutrons in the element =  $x + 0.317x = 1.317x$

According to the question,

Mass number of the element = 81

(Number of protons + number of neutrons) = 81

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

$$2.317x = 81$$

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

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  ${}_{35}^{81}\text{Br}$

---

### 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.

### Solution

Ans: Let the number of electrons in the ion carrying a negative charge be  $x$ .

Then,

$$\text{Number of neutrons present} = x + 0.111x = 1.111x$$

$$\text{Number of electrons in the neutral atom} = (x - 1)$$

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

$$\therefore \text{Number of protons in the neutral atom} = x - 1$$

Given,

$$\text{Mass number of the ion} = 37$$

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

$$2.111x = 38$$

$$x = 18$$

$$\text{Number of electrons} = 18; \text{Number of protons} = 18 - 1 = 17$$

$$\text{Atomic number of the ion} = 17; \text{Atom correspondence to ion} = Cl$$

$$\therefore \text{The symbol of the ion is } {}^{37}_{17}Cl^{-1}$$

### Question 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.

### Solution

Ans: Let the number of electrons present in ion  $A^{+3}$  be 'x'

Number of neutrons in it =  $x + 0.304x = 1.304x$

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

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

$$2.304x = 53$$

$$x = 23$$

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

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

---

### Question 45

Arrange the Following Type of Radiations in Increasing Order of Frequency:

1. radiation from microwave oven
2. amber light from traffic signal
3. radiation from FM radio
4. cosmic rays from outer space and
5. X-rays.

### Solution

Ans: The increasing order of frequency is as follows:

Radiation from FM radio < amber light < radiation from microwave oven < X- rays < cosmic rays

The increasing order of wavelength is as follows:

Cosmic rays < X-rays < radiation from microwave ovens < amber light < radiation of FM radio.

**Question 46**

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

**Solution**

Ans: Power of laser = Energy with which it emits photons

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

Where, N = number of photons emitted

h = Planck's constant

c = velocity of radiation

$\lambda$  = wavelength of radiation

Substituting the values in the given expression of E,

$$E = \frac{(5.6 \times 10^{24})(6.626 \times 10^{-34} Js)(3 \times 10^8 ms^{-1})}{(337.1 \times 10^{-9} m)} = 3.33 \times 10^6 J$$

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

**Question 47**

Neon Gas is Generally Used in Sign Boards. If it Emits Strongly at 616 nm,

**A.**

Calculate the frequency of emission.

**B.**

Calculate Distance Traveled by this Radiation in 30 s

**C.**

Calculate energy of quantum...

**D.**

Calculate number of quanta present if it produces 2 J of energy.

### Solution

Ans: Wavelength of radiation emitted = 616 nm =  $616 \times 10^{-9} \text{ m}$  (Given)

Frequency of emission,  $\nu = \frac{c}{\lambda}$

Where, c = velocity of radiation

$\lambda$  = Wavelength of radiation

Substituting the values in the given expression of frequency of emission,

$$\nu = \frac{3.0 \times 10^8 \text{ m/s}}{616 \times 10^{-9} \text{ m}} = 4.87 \times 10^{14} \text{ s}^{-1}$$

Hence, Frequency of emission,  $\nu = 4.87 \times 10^{14} \text{ s}^{-1}$

Ans: velocity of radiation, c =  $3 \times 10^8 \text{ m/s}$

Distance travelled by this radiation in 30 s =  $(3 \times 10^8 \text{ m s}^{-1})(90 \text{ s}) = 9 \times 10^9 \text{ m}$

Ans: energy of Quantum,  $E = h\nu = (6.626 \times 10^{-34} \text{ Js})(4.87 \times 10^{14} \text{ s}^{-1}) = 32.27 \times 10^{-20} \text{ J}$

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

Ans: Energy of one photon (E) =  $32.27 \times 10^{-20} \text{ J}$

Therefore,  $32.27 \times 10^{-20} \text{ J}$  of energy is present in 1 quantum number of quanta in 2J of energy =

$$\frac{2 \text{ J}}{32.27 \times 10^{-20} \text{ J}} = 6.19 \times 10^{18}$$

### Question 48

In astronomical observations, signals observed from 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.

**Solution**

Ans: From the expression of energy of one photon,  $E = \frac{hc}{\lambda}$

Where,  $\lambda$  = wavelength of radiation

$h$  = Planck's constant

$c$  = velocity of radiation

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

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

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

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

$$= \frac{3.15 \times 10^{-18} J}{3.313 \times 10^{-19} J} = 9.5 \approx 10$$

.

Hence, the number of photons received by the detector = 10

**Question 49**

Lifetimes of the molecules in the excited states are often measured by using pulsed radiation source of duration nearly in the nanosecond range. If the radiation source has a duration of 2 ns and the number of photons emitted during the pulse source is  $2.5 \times 10^{15}$ , calculate the energy of the source.

**Solution**

Ans: frequency of radiation,  $\nu = \frac{1}{2.0 \times 10^{-9} s} = 5 \times 10^8 s^{-1}$

Energy of source,  $E = N h \nu$

Where,  $N$  = number of photons emitted

$h$  = Planck's constant

$\nu$  = frequency of radiation

Substituting the values in the given expression of  $E$ ,

$$E = (2.5 \times 10^{15})(6.626 \times 10^{-34} Js)(5.0 \times 10^8 s^{-1}) = 8.282 \times 10^{-10} J$$

Hence, the energy of source,  $E = 8.282 \times 10^{-10} J$

---

### Question 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.

### Solution

Ans: Given,

Wavelength associated with first transition,  $\lambda_1 = 589 nm = 589 \times 10^{-9} m$

Wavelength associated with second transition,  $\lambda_2 = 589.6 nm = 589.6 \times 10^{-9} m$

Frequency of first wavelength is,  $\nu_1 = \frac{c}{\lambda_1} = \frac{3 \times 10^8 ms^{-1}}{589 \times 10^{-9} m} = 5.093 \times 10^{14} s^{-1}$

And, frequency of second wavelength is,  $\nu_2 = \frac{c}{\lambda_2} = \frac{3 \times 10^8 ms^{-1}}{589.6 \times 10^{-9} m} = 5.088 \times 10^{14} s^{-1}$

Energy difference between two excited states is given as,

$$\Delta E = h\nu_1 - h\nu_2 = h(\nu_1 - \nu_2)$$

$$\Rightarrow \Delta E = (6.626 \times 10^{-34} Js)(5.093 \times 10^{14} s^{-1} - 5.088 \times 10^{14} s^{-1}) = 3.31 \times 10^{-22} J$$

**Question 51**

The Work Function for Cesium Atoms is 1.9 eV.

- Calculate the threshold wavelength
- Calculate 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.

**Solution**

Ans: It is given that the work function ( $W_0$ ) for caesium atom is 1.9 eV.

From the expression,  $W_0 = \frac{hc}{\lambda_0}$

We get,  $\lambda_0 = \frac{hc}{W_0}$

Where,  $\lambda_0$  = threshold wavelength

$h$  = Planck's constant

$c$  = velocity of radiation

Substituting the values in the given expression of  $\lambda_0$ :

$$\lambda_0 = \frac{(6.626 \times 10^{-34} \text{ Js})(3.0 \times 10^8 \text{ ms}^{-1})}{1.9 \times 1.602 \times 10^{-19} \text{ J}} = 6.53 \times 10^{-7} \text{ m}$$

Hence, threshold wavelength  $\lambda_0$  is 653 nm

Ans: From the expression,  $W_0 = h\nu_0$

we get,  $\nu_0 = \frac{h}{W_0}$

Where,  $\nu_0$  = threshold frequency

$h$  = Planck's constant

Substituting the values in the given expression of  $\nu_0$ ,

$$\nu_0 = \frac{1.9 \times 1.602 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ Js}} = 4.593 \times 10^{14} \text{ s}^{-1}$$

Hence, the threshold frequency of the radiation,  $\nu_0 = 4.593 \times 10^{14} \text{ s}^{-1}$

Ans: According to the question:

Wavelength used in irradiation,  $\lambda = 500 \text{ nm}$

$$\text{Kinetic energy} = h(v - v_0) = hc \left( \frac{1}{\lambda} - \frac{1}{\lambda_0} \right)$$

$$\Rightarrow K.E = (6.626 \times 10^{-34} Js)(3.0 \times 10^8 ms^{-1}) \left( \frac{1}{500 \times 10^{-9} m} - \frac{1}{653 \times 10^{-9} m} \right)$$

$$= \frac{(1.9878 \times 10^{-26})(153 \times 10^9)}{653 \times 500} = 9.3149 \times 10^{-20} J$$

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

$$\text{Since, kinetic energy} = \frac{1}{2}mv^2 = 9.3149 \times 10^{-20} J$$

$$v = \sqrt{\frac{2(9.3149 \times 10^{-20} J)}{9.10939 \times 10^{-31} kg}} = 4.52 \times 10^5 ms^{-1}$$

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

### Question 52

Following Results are Observed When Sodium Metal is Irradiated With Different Wavelengths.

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

a) Calculate threshold wavelength

b) Calculate Planck's constant

### Solution

Ans: Assuming the threshold wavelength to be  $\lambda_0 nm$ , 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:

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

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

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

Similarly,

$$\frac{hc}{10^{-9}m} \left( \frac{1}{450} - \frac{1}{\lambda_0} \right) = \frac{1}{2}m(3.45 \times 10^3ms^{-1})^2 \text{ ----- (2)}$$

$$\frac{hc}{10^{-9}m} \left( \frac{1}{400} - \frac{1}{\lambda_0} \right) = \frac{1}{2}m(5.35 \times 10^3ms^{-1})^2 \text{ ----- (3)}$$

Dividing equation (3) by equation (1):

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

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

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

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

$$\lambda_0 = \frac{6805.537}{12.607} = 539.8nm = 540nm$$

Threshold wavelength = 540 nm

Ans: the question is not done due to the incorrect values of velocity given in the question.

### Question 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.**

### Solution

Ans: From the principle of conservation of energy, 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

we get,  $W_0 = E - K$ . E

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

Where,  $\lambda$  = wavelength of radiation

C = velocity of radiation

h = Planck's constant

Substituting the values in the given expression of E:

$$E = \frac{(6.626 \times 10^{-34} Js)(3.0 \times 10^8 ms^{-1})}{256.7 \times 10^{-9} m} = 7.744 \times 10^{-19} J$$

$$E = \frac{7.744 \times 10^{-19}}{1.602 \times 10^{-19} C} = 4.83 eV$$

The potential applied to silver metal changes to kinetic energy (K.E) of the photoelectron. Hence, K.E = 0.35 eV

Work function,  $W_0 = E - K.E$

$$W_0 = 4.83 eV - 0.35 eV = 4.48 eV$$

### Question 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  $1.5 \times 10^7 ms^{-1}$ , Calculate the energy with which it is bound to the nucleus

### Solution

Ans: Energy of incident photon (E) is given by,  $E = \frac{hc}{\lambda}$

$$E = \frac{(6.626 \times 10^{-34} Js)(3.0 \times 10^8 ms^{-1})}{(150 \times 10^{-12} m)} = 13.252 \times 10^{-16} J$$

Energy of the electron ejected (K.E) =  $\frac{1}{2} m_e v^2$

$$= \frac{1}{2} (9.10939 \times 10^{-31} kg)(1.5 \times 10^7 ms^{-1}) = 1.025 \times 10^{-16} J$$

Hence, the energy with which the electron is bound to the nucleus can be obtained as: = E – K.E

$$= 13.252 \times 10^{-16} J - 1.02 \times 10^{-16} J = 12.227 \times 10^{-16} J$$

$$= \frac{12.227 \times 10^{-16}}{1.602 \times 10^{-19} C} = 7.6 \times 10^3 eV$$

The energy with which it is bound to the nucleus is  $7.6 \times 10^3 eV$

### Question 55

**Emission Transitions in the Paschen Series end at orbit  $n = 3$  and start from orbit  $n$  and can be represented as  $\nu = 3.29 \times 10^5 (Hz) \left[ \frac{1}{3^2} - \frac{1}{n^2} \right]$  Calculate the value of  $n$  if the transition is observed at 1285 nm. Find the region of the spectrum.**

### Solution

Ans: Wavelength of transition = 1285 nm =  $1285 \times 10^{-9} m$   $\nu = 3.29 \times 10^5 (Hz) \left[ \frac{1}{3^2} - \frac{1}{n^2} \right]$

$$\text{Since, } \nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8 m s^{-1}}{1285 \times 10^{-9} m} = 2.33 \times 10^{14} s^{-1}$$

Substituting the value of  $\nu$  in the given expression,

$$2.33 \times 10^{14} = 3.29 \times 10^5 (Hz) \left[ \frac{1}{3^2} - \frac{1}{n^2} \right]$$

$$\Rightarrow \left( \frac{1}{9} - \frac{1}{n^2} \right) = \frac{2.33 \times 10^{14}}{3.29 \times 10^{15}}$$

$$\Rightarrow \frac{1}{n^2} = 1.1 \times 10^{-1} - 0.7082 \times 10^{-1}$$

$$\Rightarrow \frac{1}{n^2} = 4.029 \times 10^{-2}$$

$$n = \sqrt{\frac{1}{4.029 \times 10^{-2}}} = 4.98 \approx 5$$

Hence, for the transition to be observed at 1285 nm,  $n = 5$ . The spectrum lies in the infra-red region.

### 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.

### Solution

Ans: The radius of the  $n^{th}$  orbit of hydrogen-like particles is given by,

$$r = \frac{0.529n^2}{Z} A^0$$

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

For radius,  $r_1 = 1.3225 nm = 1.3225 \times 10^{-9} m = 1322.5 pm$

$$n_1^2 = \frac{r_1 Z}{52.9} = \frac{1322.5 Z}{52.9}$$

Similarly,

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

$$\frac{n_1^2}{n_2^2} = \frac{1322.5}{211.6} = 6.25$$

$$\frac{n_1}{n_2} = 2.5 = \frac{25}{10} = \frac{5}{2}$$

$$\Rightarrow n_1 = 5 \& n_2 = 2$$

Thus, the transition is from the  $5^{th}$  orbit to the  $2^{nd}$  orbit. It belongs to the Balmer series. Wave number ( $\bar{\nu}$ ) for the transition is given by,

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

$$= 1.097 \times 10^7 \left( \frac{21}{100} \right) = 2.303 \times 10^6 m^{-1}$$

Wavelength associated with the emission transition is given by,

$$\lambda = \frac{1}{\bar{\nu}} = \frac{1}{2.303 \times 10^6 m^{-1}} = 0.434 \times 10^6 m = 434 nm$$

**Question 57**

Dual Behavior of Matter Proposed by De Broglie Led to the Discovery of Electron Microscopes Often Used for the Highly Magnified Images of Biological Molecules and Other Types of Material. If the velocity of the electron in this microscope is  $1.6 \times 10^6 \text{ ms}^{-1}$ , calculate the de Broglie wavelength associated with this electron.

**Solution**

Ans: From de Broglie's equation,

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ Js}}{(9.10939 \times 10^{-31} \text{ kg})(1.6 \times 10^6 \text{ ms}^{-1})} = 4.55 \times 10^{-10} \text{ m} = 455 \text{ pm}$$

de Broglie's wavelength associated with the electron is 455 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

**Solution**

Ans: From de Broglie's equation,

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

$$\text{Then, } v = \frac{h}{m\lambda}$$

Where,

v = velocity of particle (neutron)

h = Planck's constant

m = mass of particle (neutron)

$\lambda$  = wavelength

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

$$v = \frac{(6.626 \times 10^{-34}) \text{ Kg m}^2 \text{ s}^{-1}}{(1.675 \times 10^{-27})(8 \times 10^{-10} \text{ m})} = \frac{6.626 \times 10^3}{1.675 \times 8} = 4.94 \times 10^2 \text{ m s}^{-1} = 494 \text{ m s}^{-1}$$

Velocity associated with the neutron =  $494 \text{ m s}^{-1}$

### 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.

### Solution

Ans: From de Broglie's equation,

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

Where,

v = velocity of particle (neutron)

h = Planck's constant

m = mass of particle (neutron)

$\lambda$  = wavelength

Substituting the values in the expression of  $\lambda$ :

$$\lambda = \frac{(6.626 \times 10^{-34} \text{ Js})}{(9.10939 \times 10^{-31} \text{ kg})(2.19 \times 10^6 \text{ m s}^{-1})} = 3.32 \times 10^{-10} \text{ m} = 332 \times 10^{-12} \text{ m} = 332 \text{ pm}$$

Wavelength associated with the electron = 332 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.

**Solution**

Ans: From de Broglie's equation,

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

Substituting the values in the expression,

$$\lambda = \frac{(6.626 \times 10^{-34} \text{ Js})}{(0.1 \text{ kg})(4.37 \times 10^5 \text{ m s}^{-1})} = 1.516 \times 10^{-38} \text{ m}$$

**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  $\frac{h}{4\pi m \times 0.05 \text{ nm}}$ . Is there any problem in defining this value?

**Solution**

Ans: From Heisenberg's uncertainty principle,

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

Where,  $\Delta x$  = uncertainty in position of the electron

$\Delta p$  = uncertainty in momentum of the electron

Substituting the values in the expression of  $\Delta p$ ,

$$\Delta p = \frac{1}{0.002 \text{ nm}} \frac{6.626 \times 10^{-34} \text{ Js}}{4 \times 3.14}$$

$$\Rightarrow \Delta p = 2.637 \times 10^{-23} \text{ kg m s}^{-1}$$

$$\text{Actual momentum} = \frac{h}{4\pi m \times 0.05nm} = \frac{6.626 \times 10^{-34} Js}{4 \times 3.14 \times 5.0 \times 10^{-11} m} = 1.055 \times 10^{-24} kgms^{-1}$$

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

### Question 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.  $n = 4, l = 2, m_l = -2, m_s = -\frac{1}{2}$

2.  $n = 3, l = 2, m_l = 1, m_s = +\frac{1}{2}$

3.  $n = 4, l = 1, m_l = 0, m_s = +\frac{1}{2}$

4.  $n = 3, l = 2, m_l = -2, m_s = -\frac{1}{2}$

5.  $n = 3, l = 1, m_l = -1, m_s = +\frac{1}{2}$

6.  $n = 4, l = 1, m_l = 0, m_s = +\frac{1}{2}$

### Solution

Ans: 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  $5(3p) < 2(3d) = 4(3d) < 3(4p) = 6(4p) < 1(4d)$ .

### 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 the lowest effective nuclear charge?

## Solution

Ans: Nuclear charge experienced by an electron (present in a multi-electron atom) is dependent 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 atoms 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 2p and 3p orbitals along with the s-orbitals. Therefore, they will experience the lowest nuclear charge.

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## Question 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

## Solution

Ans: 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.

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.

Ans: 4d will experience greater nuclear charge than 4f since 4d is closer to the nucleus.

Ans: 3p will experience greater nuclear charge since it is closer to the nucleus than 3f.

---

## Question 65

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

## Solution

Ans: 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 aluminium 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 a more effective nuclear charge than aluminium.

---

## Question 66

Indicate the number of unpaired electrons

A.

Phosphorous (P)

B.

Silicon (Si)

C.

Chromium (Cr)

D.

Iron (Fe)

## Solution

Ans: 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.

Ans: Atomic number = 14

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

The orbital picture of Si can be represented as:

From the orbital picture, phosphorus has two unpaired electrons.

Ans: 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.

Ans: Atomic number = 26

The electronic configuration is:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

The orbital picture of iron is:

From the orbital picture, iron has four unpaired electrons.

e) Krypton (Kr)

Ans: Atomic number = 36

The electronic configuration is:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

The orbital picture of krypton is:

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

---

### Question 67

Answer the following.

- a) How many subshells are associated with  $n = 4$ ?
- b) How many electrons will be present in the sub-shells having  $m_s$  value of  $-\frac{1}{2}$  for  $n = 4$ ?

### Solution

Ans:  $n = 4$  (Given)

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

Hence,  $l = 0, 1, 2, 3$

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

Ans: 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  $-\frac{1}{2} = 16$