Q.1. (i) Calculate the number of electrons which will together weigh one gram.
(ii) Calculate the mass and charge of one mole of electrons.
Ans:

(i) Mass of 1 electron = 9.108×10^{-28} g

Hence,

1 g = $1/(9.108 \times 10^{-28})$ = 1.098×10^{27} electrons

(ii) Mass of one mole of electron = $9.108 \times 10^{-28} \times 6.022 \times 10^{23}$

We get,

= 5.48×10^{-4} g

Charge on one mole of electron = 1.6×10^{-19} C x 6.022×10^{23}

We get,

= 9.63×10^4 C

Q.2. (i) Calculate the total number of electrons present in one mole of methane.
(ii) Find (a) the total number and (b) the total mass of neutrons in 7 mg of ^{14}C.
(Assume that mass of a neutron = 1.675×10^{-27} kg).
(iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3
at STP.
Will the answer change if the temperature and pressure are changed?

Ans:

(i) 1 molecule of methane contains 10 electrons (6 from carbon, 4 from hydrogen)

Therefore, 1 mole of methane contains = $10 \times 6.022 \times 10^{23}$ = 6.022×10^{24} electrons.

(ii) (a)1 g atom of ^{14}C = 14 g

Since 14 g = 14000 mg have $6.022 \times 10^{23} \times 8$ neutrons

Therefore, 7 mg will have neutrons = $(6.022 \times 10^{23} \times 8)/14000 \times 7$ = 2.4088×10^{22}

(b) mass of 1 neutron = 1.675×10^{-27}kg

Hence, mass of 2.4088×10^{21} neutrons = $2.4088 \times 10^{21} \times 1.67 \times 10^{-27}$ = 4.0347×10^{-6} kg

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

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

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

34 mg of NH^3 have molecules = $\frac{6.022 \times 10^{23}}{(17 \times 10^3 mg)} \times (34mg)$

$= 1.2044 \times 10^{20} molecules$

Step II. Calculation of total number and mass of protons

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

No. of protons present in 12.044 x 10^{20} molecules of NH_3 = 12.044 x 10^{20} x 10

= 1.2044 x 10^{22} protons

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

Mass of 1.2044 x 10^{22} Protons = (1.67 x 10^{-27} kg) x 1.2044 x 10^{22}

= 2.01 x 10^{-5} kg

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

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

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

Ans:

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

Mass number = 13

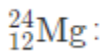Atomic number = Number of protons = 6

Therefore, total number of neutrons in 1 carbon atom = Mass number – Atomic number = 13 – 6 = 7

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

Mass number = 16

Atomic number = Number of protons = 8

Therefore, number of neutrons = Mass number – Atomic number = 16 – 8 = 8

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

Mass number = 24

Atomic number = Number of protons = 12

Number of neutrons = Mass number – Atomic number = 24 – 12 = 12

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

Mass number = 56

Atomic number = Number of protons = 26

Number of neutrons = Mass number – Atomic number = 56 – 26 = 30

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

Mass number = 88

Atomic number = Number of protons = 38

Number of neutrons = Mass number – Atomic number = 88 – 38 = 50

Q.4. Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)

(I)Z = 17, A = 35

(II)Z = 92, A = 233

(III)Z = 4, A = 9

Ans:

(I) The atom with atomic number 17 and mass number 35 is chlorine. Therefore, its complete symbol is

$^{35}_{17}Cl$

(II) The atom with atomic number 92 and mass number 233 is uranium. Therefore, its complete symbol is $^{233}_{92}U$

(III) The atom with atomic number 4 and mass number 9 is beryllium. Therefore, its complete symbol is

$^{9}_{4}Be$

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

Ans: Rearranging the expression,

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

the following expression can be obtained,

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

Here, ν denotes the frequency of the yellow light

c denotes the speed of light (3×10^8 m/s)

λ denotes the wavelength of the yellow light (580 nm, 580×10^{-9} m/s)

Substituting these values in eq. (1):

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

Therefore, the frequency of the yellow light which is emitted by the sodium lamp is:

5.17×10^{14} s^{-1}

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

The wave number of the yellow light is

Q.6. Find the energy of each of the photons which
(i) correspond to light of frequency 3×10¹⁵ Hz.
(ii) have a wavelength of 0.50 Å.

Ans:

(i)

The energy of a photon (E) can be calculated by using the following expression:

E= $h\nu$

Where, 'h' denotes Planck's constant, which is equal to 6.626×10^{-34} J s

ν (frequency of the light) = 3×10^{15} Hz

Substituting these values in the expression for the energy of a photon, E:

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

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

(ii)

The energy (E) of a photon having wavelength (λ) is given by the expression:

E=hc/λ

Where,

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

c (speed of light) = 3 x 10⁸ m/s

Substituting these values in the equation of 'E':

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

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

Q.7. Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0 × 10⁻¹⁰ s.

Ans: Frequency of the light wave $(\nu) = \frac{1}{Period}$

$$\frac{1}{period} = \frac{1}{2.0 \times 10^{-10} s}$$

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

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

Where,

c denotes the speed of light, 3 x 10⁸ m/s

Substituting the values in the given expression of λ:

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

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

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

Ans: Energy of photon $(E) = \frac{hc}{\lambda}$

Where, λ is the wavelength of the photons = $4000 \times 10^{-12} m = 4 \times 10^{-9} m$

c denotes the speed of light in vacuum = 3 x 10⁸ m/s

h is Planck's constant, whose value is 6.626 x 10⁻³⁴ J s

Substituting these values, we get,

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

= 4.969 x 10⁻¹⁷ J

Hence, 1J energy of photons = $\frac{1}{4.969 \times 10^{-17}} = 2.012 \times 10^{16} photons$

Q.9. A photon of wavelength 4 × 10⁻⁷ m strikes on metal surface, the work function of the metal is 2.13 eV. Calculate (i) the energy of the photon (eV), (ii) the kinetic energy of the emission, and (iii) the velocity of the photoelectron (1 eV= 1.6020 × 10⁻¹⁹ J).

Ans:

(i)

Energy of the photon (E)= $h\nu = \frac{hc}{\lambda}$

Where, h denotes Planck's constant, whose value is 6.626 x 10⁻³⁴ J s

c denotes the speed of light = 3 x 10⁸ m/s

λ = wavelength of the photon = $4 \times 10^{-7} m/s$

Substituting these values, we get,

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

Therefore, energy of the photon = 4.97 x 10⁻¹⁹ J

(ii)

The kinetic energy of the emission E_k is given by

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

$$= (E - W)eV$$

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

$$= (3.1020 - 2.13)eV$$

$$= 0.9720\,eV$$

Therefore, the kinetic energy of the emission = 0.97 eV.

(iii)

The velocity of the photoelectron (v) can be determined using the following expression:

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

$$\Rightarrow v = \sqrt{\frac{2(h\nu - h\nu_0)}{m}}$$

Where $(h\nu - h\nu_0)$ is the K.E of the emission (in Joules) and 'm' denotes the mass of the photoelectron.
Substituting the values in the given expression of v:

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

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

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

Therefore, the velocity of the photoelectron is 5.84 x 10^5 ms^{-1}

Q.10. Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJ mol⁻¹.

Ans.
Ionization energy (E) of sodium = $\frac{N_A hc}{\lambda}$

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

= 4.947 x 10⁵ J mol⁻¹

= 494.7 x 10³ J mol⁻¹

= 494 kJ mol⁻¹

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

Ans: Power of the bulb, P = 25 Watt = 25 Js⁻¹

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

Substituting these values in the expression of E:

$$E = \frac{(6.626\times10^{-34})(3\times10^8)}{(0.57\times10^{-6})} = 34.87 \times 10^{-20}\ J$$

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

Thus, the rate of emission of quanta (per second) = $E = \frac{25}{34.87\times10^{-20}} = 7.169 \times 10^{19}\ s^{-1}$

Q.12. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (v₀) and work function (W₀) of the metal.

Ans: Wavelength of the radiation, λ = 6800 Å = 6800 x 10⁻¹⁰ m

Threshold frequency of the metal (ν_0) = $\frac{c}{\lambda_0} = \frac{3\times10^8 ms^{-1}}{6.8\times10^{-7}m} = 4.41 \times 10^{14}\ s^{-1}$

Therefore, work function, (W₀) = $h\nu_0$

= (6.626 x 10⁻³⁴ J s) (4.41 x 10¹⁴ s⁻¹)

= 2.922 x 10⁻¹⁹ J

Q.13. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes the transition from an energy level with n = 4 to an energy level with n = 2?

Ans: The n_i = 4 and n_f = 2 transition results in a spectral line of the Balmer series. The energy involved in this transition can be calculated using the following expression:

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

Substituting these values in the expression of E:

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

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

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

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

Here, the -ve sign denotes the emitted energy.

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

Since E = hc/λ

Substituting these values in the expression of λ

$$\lambda = \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{(4.0875 \times 10^{-19})} = 4.8631 \times 10^{-7} \, m$$

λ = 486.31 x 10⁻⁹ m

= 486 nm

Q.14. How much energy is required to ionise a H atom if the electron occupies n = 5 orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from n =1 orbit).

Ans: The expression for the ionization energy is given by,

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

Where Z denotes the atomic number and n is the principal quantum number

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

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

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

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

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

Therefore, the required energy for the ionization of hydrogen from n = 5 to n = ∞ is 8.72 x 10^{-20} J

Energy required for n_1 = 1 to n = ∞

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

= 2.18 x 10^{-18} J

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

Hence, a lower amount of energy is required in order to ionize electrons in the 5th orbital of a hydrogen atom when compared to that in the ground state.

Q.15. What is the maximum number of emission lines when the excited electron of a H atom in n = 6 drops to the ground state?

Ans. A total number of 15 lines (5 + 4 + 3 + 2 + 1) will be obtained in this hydrogen emission spectrum.

The number of spectral lines emitted when an electron in the nth level drops down to the ground state is given by

$$\frac{n(n-1)}{2}$$

Since n = 6, total no. spectral lines = $\frac{6(6-1)}{2} = 15$

Hence, the maximum number of emission lines when the excited electron of a H atom in n = 6 drops to the ground state is 15

Q.16. (i) The energy associated with the first orbit in the hydrogen atom is –2.18 × 10⁻¹⁸ J atom⁻¹. What is the energy associated with the fifth orbit? (ii) Calculate the radius of Bohr's fifth orbit for the hydrogen atom.

Ans. The expression for the energy associated with nth orbit in hydrogen atom is

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

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

$$E_5 = \frac{-(2.18 \times 10^{-18})}{(5)^2} = \frac{-(2.18 \times 10^{-18})}{25} = -8.72 \times 10^{-20}\, J atom^{-1}$$

(ii) Radius of Bohr's n^{th} orbit for hydrogen atom is given by, r_n = (0.0529 nm) n^2

For, n = 5

r_5 = (0.0529 nm) (5^2)

r_5 = 1.3225 nm

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

Ans. For the Balmer series of the hydrogen emission spectrum, n_i = 2. Therefore, the expression for the wavenumber $(\bar{\nu})$ is:

$$\bar{\nu} = [\tfrac{1}{(2)^2} - \tfrac{1}{n_f^2}](1.097 \times 10^7\, m^{-1})$$

Since wave number $(\bar{\nu})$ is inversely proportional to the transition wavelength, the lowest possible value of $(\bar{\nu})$ corresponds to the longest wavelength transition.

For $(\bar{\nu})$ to be of the lowest possible value, n_f should be minimum. In the Balmer series, transitions from n_i = 2 to n_f = 3 are allowed.

Hence, taking n_f = 3, we get:

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

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

$$= (1.097 \times 10^7)[\tfrac{9-4}{36}]$$

$$= (1.097 \times 10^7)[\tfrac{5}{36}]$$

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

Q.18. What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state? The ground state electron energy is –2.18 × 10⁻¹¹ ergs.

Ans. Step I. Calculation of energy required

The energy of electron (E_n) = $\frac{2.18 \times 10^{-11}}{n^2} ergs$ = $\frac{-2.18 \times 10^{-18} J}{n^2} (\because 1J = 10^7 ergs)$

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

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

$\therefore$ Energy required $(\Delta E) = E_5 - E_1 = (-\frac{2.18}{25} \times 10^{-18} J) - (-\frac{2.18}{1} \times 10^{-18} J)$

$= 2.18 \times 10^{-18}(1 - \frac{1}{25})J$

$= 2.18 \times 10^{-18} \times \frac{24}{25} = 2.09 \times 10^{-18} J$

Step II. Calculation of wavelength of light emitted

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

$\therefore \lambda = \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34} J-s) \times (3 \times 10^8 ms^{-1})}{(2.09 \times 10^{-18} J)}$

= 9.50 x 10⁻⁸ m

= 950 Å ($\because$ 1Å = 10⁻¹⁰m)

Q.19. The electron energy in hydrogen atom is given by E_n= (–2.18 × 10⁻¹⁸)/n² J. Calculate the energy required to remove an electron completely from the n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition?

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

Ans.

Required energy for the ionization from n = 2 is:

$\Delta E = E_\infty - E_2$

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

$$= [\frac{(2.18\times10^{-18})}{4} - 0]J$$

= 0.545 x 10⁻¹⁸ J

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

Now we know that,

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

Here, λ is the longest wavelength causing the transition

$$E = \frac{(6.626\times10^{-34})(3\times10^8)}{(5.45\times10^{-19})} = 3.647 \times 10^{-7} m$$

= 3647 x 10⁻¹⁰

= 3647 Å

Q.20. Calculate the wavelength of an electron moving with a velocity of 2.05 × 10⁷ m s⁻¹

Ans. As per de Broglie's equation,

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

Where, λ denotes the wavelength of the moving particle

m is the mass of the particle

v denotes the velocity of the particle

h is Planck's constant

Substituting these values we get,

$$\lambda = \frac{(6.626\times10^{-34})Js}{(9.10939\times10^{-31}kg)(2.05\times10^7ms^{-1})}$$

= 3.548 x 10⁻¹¹ m

Therefore, the wavelength of an electron moving with a velocity of 2.05 x 10⁷ ms⁻¹ is 3.548 x 10⁻¹¹ m

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

Ans.

The wavelength of an electron can be found by de Broglie's equation:

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

Given the K.E. of electron 3.0×10^{-25} J which is equal to $\frac{1}{2}m\nu^2$

Hence we get,

$$\frac{1}{2}m\nu^2 = 3.0 \times 10^{-25} J$$

$$\nu = \sqrt{\frac{2KE}{m}} = \sqrt{\frac{2(3.0 \times 10^{-25} J)}{9.1 \times 10^{-31} kg}} = 811.579 m/s$$

Hence the wavelength is given by,

$$\lambda = \frac{6.626 \times 10^{-34} Js}{(9.1 \times 10^{-31} kg)(811.579 m/s)}$$

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

Q. 22. Which of the following are isoelectronic species i.e., those having the same number of electrons?
Na^+ , K^+ , Mg^{2+}, Ca^{2+}, S^{2-}, Ar.

Ans:

Number of electrons present in different species are:

$_{11}Na^+ = 11 - 1 = 10$; $_{19}K^+ = 19 - 1 = 18$;

$_{12}Mg^{2+} = 12 - 2 = 10$;

$_{20}Ca^{2+} = 20 - 2 = 18$; $_{16}S^{2-} = 16 + 2 = 18$

Ar = 18

Hence isoelectronic species are

(i) Na^+ and Mg^{2+}

(having number of electrons = 10)

(ii) K^+, Ca^{2+}, S^{2-} and Ar (having number of electrons = 18)

Q.23. (I)Write the electronic configurations of the following ions:

(a)H^-

(b)Na^+

(c)O^{2-}

(d)F^-

(II) What are the atomic numbers of elements whose outermost electrons are represented by

(a) $3s^1$

(b) $2p^3$ and

(c) $3p^5$?

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

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

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

(c)[Ar] $4s^2\ 3d^1$.

Ans:

(I) (a) H^- ion

The electronic configuration of the Hydrogen atom (in its ground state) $1s^1$. The single negative charge on this atom indicates that it has gained an electron. Thus, the electronic configuration of $H^- = 1s^2$

(b) Na^+ ion

Electronic configuration of Na = $1s^2\ 2s^2\ 2p^6\ 3s^1$. Here, the +ve charge indicates the loss of an electron.

∴Electronic configuration of $Na^+ = 1s^2\ 2s^2\ 2p^6$

(c) O^{2-} ion

Electronic configuration of 0xygen = $1s^2\ 2s^2\ 2p^4$. The '-2' charge suggests that it has gained 2 electrons.

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

(d) F^- ion

Electronic configuration of Fluorine = $1s^2 2s^2 2p^5$. The negative charge indicates that it has gained one electron

∴ Electronic configuration of F^- ion = $1s^2\ 2s^2\ 2p^6$

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

Complete electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^1$.

Total number of electrons in the atom = 2 + 2 + 6 + 1 = 11

∴ The element's atomic number is 11

(b) $2p^3$

Complete electronic configuration: $1s^2\ 2s^2\ 2p^3$.

Total number of electrons in the atom =2 + 2 + 3 = 7

∴The element's atomic number is 7

(c) 3p ^{5}

Complete electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$.

Total number of electrons in the atom = 2 + 2 + 6 + 2 + 5 = 17

∴ The element's atomic number is 17

(III)(a)[He] 2s ^{1}

Complete electronic configuration: $1s^2\ 2s^1$.

∴ The element's atomic number is 3. The element is lithium (Li)

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

Complete electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$.

∴ The element's atomic number is 15. The element is phosphorus (P).

(c)[Ar] 4s ^{2} 3d^1

Complete electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$.

∴ The element's atomic number is 21. The element is scandium (Sc).

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

Ans. For g-orbitals, l = 4.

For any given value of 'n', the possible values of 'l' range from 0 to (n-1).

∴ For l = 4 (g orbital), least value of n = 5.

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

Ans: For the 3d orbital:

Possible values of the Principal quantum number (n) = 3

Possible values of the Azimuthal quantum number (l) = 2

Possible values of the Magnetic quantum number (m$_l$) = – 2, – 1, 0, 1, 2

Q.26. An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons and (ii) the electronic configuration of the element.

Ans:

(i)

In a neutral atom, number of protons = number of electrons.

∴ Number of protons present in the atoms of the element = 29

(ii)

The electronic configuration of this element (atomic number 29) is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$. The element is copper (Cu).

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

Ans: No. electrons present in H_2 = 1 + 1 = 2.

∴ Number of electrons in H_2^+ = 2 − 1 = 1

Number of electrons in H_2 = 1 + 1 = 2

Number of electrons in O_2 = 8 + 8 = 16.

∴ Number of electrons in O_2^+ = 16 − 1 = 15

Q.28. (I)An atomic orbital has n = 3. What are the possible values of l and m_l ?

(II)List the quantum numbers (m_l and l) of electrons for 3d orbital.

(III) Which of the following orbitals are possible? 1p, 2s, 2p and 3f

Ans.

(I)

The possible values of 'l' range from 0 to (n − 1). Thus, for n = 3, the possible values of l are 0, 1, and 2.

The total number of possible values for m_l = (2l + 1). Its values range from -l to l.

For n = 3 and l = 0, 1, 2:

$m_0 = 0$

$m_1 = -1, 0, 1$

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

(II)

For 3d orbitals, n = 3 and l = 2. For l = 2 , possible values of m_2 = −2, −1, 0, 1, 2

(III)

It is possible for the 2s and 2p orbitals to exist. The 1p and 3f cannot exist.

For the 1p orbital, n=1 and l=1, which is not possible since the value of l must always be lower than that of n.

Similarly, for the 3f orbital, n =3 and l = 3, which is not possible.

Q.29. Using s, p and 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.

Ans:

(a)n = 1, l = 0 implies a 1s orbital.

(b)n = 3 and l = 1 implies a 3p orbital.

(c)n = 4 and l = 2 implies a 4d orbital.

(d)n = 4 and l = 3 implies a 4f orbital.

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

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

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

Ans. (a) Not possible. The value of n cannot be 0.

(b) Possible.

(c) Not possible. The value of l cannot be equal to that of n.

(d) Possible.

(e) Not possible. For n = 3, l cannot be 3

(f) Possible.

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

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

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

Ans. (a) If n is the principal quantum number, the total number of electrons in the atom = $2n^2$

∴ For n = 4, Total number of electrons = $2(4)^2$ = 32 and half of them have m_s = -1/2

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

(b) n = 3, l = 0 indicates the 3s orbital. Therefore, number of electrons with n = 3 and l = 0 is 2.

Q.32. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.

Ans. Hydrogen atom have only one electron. As per Bohr's postulates, the angular momentum of this electron is:

$$mvr = n\frac{h}{2\pi} \quad ....(1)$$

Where, n = 1, 2, 3, …

As per de Broglie's equation:

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

$$or \ mv = \frac{h}{\lambda} \quad ....(2)$$

Substituting the value of mv from equation (2) in equation (1) we get,

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

$$or \ 2\pi r = n\lambda \quad .....(3)$$

But '2πr' is the Bohr orbit's circumference. Therefore, equation (3) proves that the Bohr orbit's circumference for the hydrogen atom is an integral multiple of the de Broglie wavelength of the electron, which is revolving around the orbit.

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

Ans. The wave number associated with the Balmer transition for the He⁺ ion (n = 4 to n = 2) is given by:

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

Where, $n_1 = 2$ $n_2 = 4$

Z = atomic number of helium

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

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

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

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

As per the question, the desired transition in the hydrogen spectrum must have the same wavelength as that of He$^+$ spectrum.

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

$$\Rightarrow \left[\tfrac{1}{n_1^2} - \tfrac{1}{n_2^2}\right] = \tfrac{3}{4} \quad \ldots\ldots.(1)$$

This equality is true only when the value of $n_1 = 1$ and $n_2 = 2$.

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

Q.34. Calculate the energy required for the process

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

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

Ans. The energy associated with hydrogen-like species is:

$$E_n = -2.18 \times 10^{-18}\left(\tfrac{Z^2}{n^2}\right)J$$
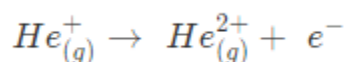
For the ground state of the hydrogen atom,

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

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

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

For the given process

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

An electron is removed from n = 1 to n = ∞.

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

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

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

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

Q.35. If the diameter of a carbon atom is 0.15 nm, calculate the number of carbon atoms which can be placed side by side in a straight line across the length of the scale of length 20 cm long.

Ans. 1 m = 100 cm

1 cm = 10^{-2} m

Length of the scale = 20 cm = 20 x 10^{-2} m

Diameter of one carbon atom = 0.15 nm = 0.15 x 10^{-9} m

Space occupied by one carbon atom 0.15 x 10^{-9} m.

∴ Number of carbon atoms that can be placed in a straight line $\frac{20 \times 10^{-2}m}{0.15 \times 10^{-9}m}$

= 133.33 x 10^7

= 1.33 x 10^9

Q.36. 2 x 10^8 m atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

Ans. Length of the arrangement = 2.4 cm

Number of carbon atoms present = 2×10^8

The diameter of the carbon atom = $\dfrac{2.4 \times 10^{-2} m}{2 \times 10^8 m}$

= 1.2 x 10⁻¹⁰ m

∴ Radius of carbon atom = $\dfrac{Diameter}{2}$

$= \dfrac{1.2 \times 10^{-10} m}{2}$

= 6.0 x 10⁻¹¹ m

Q.37. The diameter of the zinc atom is 2.6Å. Calculate (a) radius of zinc atom in pm and (b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

Ans. (a) Given

Diameter of zinc atom = 2.6 Å

= 2.6 x 10⁻¹⁰ m

= 260 x 10⁻¹² m

= 260 pm

Radius of zinc atom in pm = (260 pm)/2 = 130 pm

(b) Given

Length of the arrangement = 1.6 cm

= 1.6 x 10⁻² m

Diameter of a zinc atom = 2.6 Å = 2.6 x 10⁻¹⁰ m

∴ Number of zinc atom present in the arrangement

= (1.6 x 10⁻² m)/(2.6 x 10⁻¹⁰ m)

= 0.6153 x 10⁸ m

= 6.153 x 10⁷ m

Q.38. A certain particle carries 2.5 x 10⁻¹⁶ C of static electric charge. Calculate the number of electrons present in it.

Ans.

Charge carried by one electron = 1.6022 x 10⁻¹⁹ C

Therefore, number of electrons present in a particle carrying $2.5 \times 10^{-16} C = \frac{1}{1.6022 \times 10^{-19} C}(2.5 \times 10^{-16} C)$

= 1.560 x 10³ C

= 1560 C

Q.39. In Milikan's experiment, the 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 x 10⁻¹⁸ C, calculate the number of electrons present on it.

Ans.

Charge on the oil drop = 1.282 x 10⁻¹⁸ C

Charge on one electron = 1.6022 x 10⁻¹⁹ C

Therefore, number of electrons present on the oil drop

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

= 0.8001 x 10¹

= 8.0

Q.40. In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the α-particles. If the thin foil of light atoms like Aluminium etc. is used, what difference would be observed from the above results?

Ans.

Rutherford used alpha rays scattering to reveal the structure of the atom in 1911. The nucleus of heavy atoms is big and carries a lot of positive charge. As a result, when certain alpha particles hit the nucleus, they are easily deflected back. A number of alpha particles are also deflected at tiny angles due to the nucleus's substantial positive charge. If light atoms are used, their nuclei will be light and their nuclei will have a modest positive charge. As a result, the number of particles deflected back and those deflected at an angle will be insignificant.

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

Ans.

The general convention followed while representing elements along with their atomic masses (A), and their atomic numbers (Z) is $^{A}_{Z}X$.

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

^{79}Br is an acceptable representation but ^{35}Br is not since the atomic numbers of elements are constant but mass numbers are not (due to the existence isotopes).

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

Ans.

Given, Mass number of the element = 81, which implies that (number of protons + number of neutrons) = 81

Let the number of protons in the element be x.

Therefore, number of neutrons in the element = x + 31.7% of x

= x + 0.317 x

= 1.317 x

⇒ x + 1.317x = 81

2.317x = 81

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

= 34.95

$$x \simeq 35$$

Therefore, total number of protons = 35, which implies that atomic number = 35.

Element with atomic number 35 is Br

Therefore, the element is $^{81}_{35}\text{Br}$

Q.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.

Ans.

Let the number of electrons in the negatively charged ion be x.

Then, number of neutrons present = x + 11.1% of x = x + 0.111 x = 1.111 x

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

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

Therefore, number of protons present in the neutral atom = x – 1

Given, mass number of the ion = 37

Mass number = number of protons + number of neutrons

(x − 1) + 1.111x = 37

2.111x = 38

x = 18

∴ Number of protons = Atomic number = x − 1 = 18 − 1 = 17

Therefore, the symbol of the ion is $^{37}_{17}Cl^-$

Q.44. An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign the symbol to this ion.

Ans.

Let the total number of electrons present in the ion A^{3+} be x.

Therefore, total 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

Therefore, number of protons in neutral atom = x + 3

Mass number = number of protons + number of neutrons

The Mass number of the ion is 56 (Given)

Therefore, (x+3) + (1.304x) = 56

2.304x = 53

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

x = 23

Therefore, number of protons = x + 3 = 23 + 3 = 26

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

Q.45. Arrange the following type of radiations in increasing order of frequency: (a) radiation from microwave oven (b) amber light from traffic signal (c) radiation from FM radio (d) cosmic rays from outer space and (e) X-rays.

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 a wavelength is as follows:

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

Q.46. Nitrogen laser produces radiation at a wavelength of 337.1 nm. If the number of photons emitted is 5.6 x 10²⁴, calculate the power of this laser.

Ans.

Power of laser is, $E = \frac{Nhc}{\lambda}$

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):

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

= 0.3302 x 10⁷ J

= 3.3 x 10⁶ J

Hence, the power of the laser is 3.3 x 10⁶ J

Q.47. Neon gas is generally used in the signboards. If it emits strongly at 616 nm, calculate (a) the frequency of emission,

(b) distance travelled by this radiation in 30 s

(c) the energy of quantum and

(d) the number of quanta presents if it produces 2 J of energy.

Ans.

Wavelength of the emitted radiation = 616 nm = 616 x 10⁻⁹ m (Given)

(a)Frequency of the emission (ν)

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

Where, c = speed of the radiation

λ = wavelength of the radiation

Substituting these values in the expression of (ν):

$$\nu = \frac{3 \times 10^8 \, m/s}{616 \times 10^{-9} m}$$

= 4.87 x 10⁸ x 10⁹ x 10⁻³ s⁻¹

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

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

(b) Speed of the radiation, c = 3 x 10⁸ ms⁻¹

Distance travelled by the radiation in a timespan of 30 s

= (3 x 10⁸ ms⁻¹) (30s)

= 9 x 10⁹ m

(c) Energy of one quantum (E) = hv

= (6.626 x 10⁻³⁴ J s) (4.87 x 10¹⁴ s⁻¹)

Energy of one quantum (E) = 32.27 x 10⁻²⁰ J

Therefore, 32.27 x 10⁻²⁰ J of energy is present in 1 quantum.

(d) Number of quanta in 2 J of energy = $\frac{2 J}{32.27 \times 10^{-20} J}$

= 6.19 x 10¹⁸

= 6.2 x 10¹⁸

Q.48. In astronomical observations, signals observed from the distant stars are generally weak. If the photon detector receives a total of 3.15 x 10⁻¹⁸ J from the radiations of 600 nm, calculate the number of photons received by the detector.

Ans.

From the expression of energy of one photon (E),

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

Where,

λ denotes the wavelength of the radiation

h is Planck's constant

c denotes the velocity of the radiation

Substituting these values in the expression of E:

$$E = \frac{(6.626 \times 10^{-34} Js)(3 \times 10^8 ms^{-1})}{(600 \times 10^{-9})} = 3.313 \times 10^{-19} J$$

Energy held by one photon = 3.313 x 10⁻¹⁹ J

Number of photons received with 3.15 x 10⁻¹⁸ J energy

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

= 9.5

≈ 10

Q.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 the duration of 2 ns and the number of photons emitted during the pulse source is 2.5 x 10¹⁵ J, calculate the energy of the source.

Ans.

Frequency of radiation (ν),

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

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

Energy (E) of source = Nhv

Where,

N is the number of photons emitted

h is Planck's constant

v denotes the frequency of the radiation

Substituting these values in the expression of (E):

E = (2.5 x 10¹⁵) (6.626 x 10⁻³⁴ J s) (5.0 x 10⁸ s⁻¹)

E = 8.282 x 10⁻¹⁰ J

Hence, the energy of the source (E) is 8.282 x 10⁻¹⁰ J

Q. 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.

Ans.

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

$$\therefore Frequency, \nu_1 = \frac{c}{\lambda_1} = \frac{3.0\times10^8 ms^{-1}}{589\times10^{-9}m}$$

= 5.093 x 10¹⁴ s⁻¹

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

$$\therefore Frequency(\nu_2) = \frac{c}{\lambda_2} = \frac{3.0\times10^8}{589.6\times10^{-9}} = 5.088 \times 10^{14}s^{-1}$$

The difference in energy $(\Delta E) = E_1 - E_2 = h(\nu_1 - \nu_2)$

= 6.626 x 10⁻³⁴ x (5.093 – 5.088) x 10¹⁴ = 3.313 x 10⁻²² J

Q.51. The work function for the 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 of 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

Ans.

Given, the work function of caesium (W_0) = 1.9 eV.

(a)From the $W_0 = \frac{hc}{\lambda_0}$ expression, we get:

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

Where,

λ_0 is the threshold wavelength

h is Planck's constant

c denotes the velocity of the radiation

Substituting these values in the expression of (λ_0):

$$\lambda_0 = \frac{(6.626\times10^{-34}Js)(3\times10^8 ms^{-1})}{(1.9\times1.602\times10^{-19}J)} = 6.53 \times 10^{-7}m$$

Therefore, threshold wavelength (λ_0) = 653 nm.

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

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

Where,

ν_0 is the threshold frequency

h denotes Planck's constant

Substituting these values in the expression of ν_0:

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

(1 eV = 1.602 x 10⁻¹⁹ J)

ν_0 = 4.593 x 10¹⁴ s⁻¹

Hence, threshold frequency of the radiation (ν_0) = 4.593 x 10¹⁴ s⁻¹

(c) Given, Wavelength used in the irradiation (λ) = 500 nm

Kinetic energy = h (ν – ν₀)

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

$$= (6.626 \times 10^{-34} Js)(3.0 \times 10^8 ms^{-1})\left(\frac{\lambda_0 - \lambda}{\lambda\lambda_0}\right)$$

$$= (1.9878 \times 10^{-26} Jm)\left[\frac{(653-500)10^{-9}m}{(653)(500)10^{-18}m^2}\right]$$

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

= 9.3149 x 10⁻²⁰ J

Kinetic energy held by the ejected photoelectron = 9.3149 x 10⁻²⁰ J

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

$$\nu = \sqrt{\frac{2(9.3149 \times 10^{-20} J)}{9.10939 \times 10^{-31}}}$$

$$= \sqrt{2.0451 \times 10^{11} m^2 s^{-2}}$$

= 4.52 x 10⁵ ms⁻¹

Q.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 × 10⁻⁵ (cm s⁻¹)	2.55	4.35	5.35

Ans.

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

$Then\, h(\nu - \nu_0) = \frac{1}{2}mv^2 \,or\, hc = \left(\frac{1}{\lambda} - \frac{1}{\lambda_0}\right) = \frac{1}{2}mv^2$

Substituting the given results of the three experiments, we get

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

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

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

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

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

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

$or\, \lambda_0 - 450 = 2.619\lambda - 1309.5$

$or\, 1.619\lambda_0 = 859.5 \Rightarrow \lambda_0 = 531nm$

Substituting this value in equation (iii) we get,

$\frac{h \times (3 \times 10^8)}{10^{-9}}\left(\frac{1}{400} - \frac{1}{531}\right) = \frac{1}{2}(9.11 \times 10^{-31}) \times (5.20 \times 10^6)^2$

h = 6.66 x 10⁻³⁴ J s

Q.53. The ejection of the photoelectron from the silver metal in the photoelectric effect experiment can be stopped by applying the voltage of 0.35 V when the radiation 256.7 nm is used. Calculate the work function for silver metal.

Ans.

As per the law of conservation of energy, the energy associated with an incident photon (E) must be equal to the sum of its kinetic energy and the work function (W_0) of the radiation.

$E = W_0 + K.E$

$\Rightarrow W_0 = E - K.E$

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

Where,

c denotes the velocity of the radiation

h is Planck's constant

λ is the wavelength of the radiation

Substituting these values in the expression of E:

$$E = \frac{(6.626 \times 10^{-34} Js)(3 \times 10^8 ms^{-1})}{(256.7 \times 10^{-9} m)} = 7.744 \times 10^{-19} J$$

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

E = 4.83 eV

The potential that is applied to the silver is transformed into the kinetic energy (K.E) of the photoelectron.

Hence,

K.E = 0.35 V

K.E = 0.35 eV

Therefore, Work function, W_0 = E – K.E

= 4.83 eV – 0.35 eV

= 4.48 eV

Q.54. If the photon of the wavelength 150 pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of 1.5 x 10⁷ ms⁻¹, calculate the energy with which it is bound to the nucleus.

Ans.

Energy of incident photon (E) is given by,

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

$$E = \frac{(6.626\times10^{-34})(3\times10^8)}{(150\times10^{-12})} = 1.3252 \times 10^{-15} \, J$$

$$\simeq 13.252 \times 10^{-16} J$$

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

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

= 10.2480 x 10⁻¹⁷ J

=1.025 x 10⁻¹⁶ J

Therefore , the energy that binds the electron to the nucleus can be determined using the following formula:

= E – K.E

= 13.252 x 10⁻¹⁶ J – 1.025 x 10⁻¹⁶ J

$$= 12.227 \times 10^{-16} J = \frac{12.227\times10^{-16}}{1.602\times10^{-19}} eV$$

= 7.6 x 10³ eV

Q.55. Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as v = 3.29 x 10¹⁵ (Hz) [1/3² – 1/n ²] Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

Ans.

Here, $v = 3.29 \times 10^{15}(Hz)\left[\frac{1}{3^2} - \frac{1}{n^2}\right]$ .... (1)

Also, $\nu = \frac{c}{\lambda} = \frac{3.0\times10^8 ms^{-1}}{1285\times10^{-9}m}$ .... (2)

From equations (1) and (2) we have

$$3.29 \times 10^{15} \left[\frac{1}{3^2} - \frac{1}{n^2}\right] = \frac{3.0 \times 10^8}{1285 \times 10^{-9}}$$

$$or \frac{1}{9} - \frac{1}{n^2} = \frac{3.0 \times 10^8}{1285 \times 10^{-9} \times 3.29 \times 10^{15}} = 0.071$$

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

$$= 0.04 = \frac{1}{25}$$

or n^2 = 25 or n = 5

The radiation corresponding to 1285 nm lies in the infrared region

Q.56. Calculate the wavelength for the emission transition if it starts from the orbit having radius 1.3225 nm and ends at 211.6 pm. Name the series to which this transition belongs and the region of the spectrum.

Ans.

The radius of the nth orbit of hydrogen-like particles is given by,

$$r = \frac{0.529n^2}{Z} \text{ Å}$$

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

For radius (r_1) = 1.3225 nm

= 1.3225 x 10^{-9} m

= 1322.5 x 10^{-12} m

= 1322.5 pm

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

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

Similarly,

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

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

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

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

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

⇒ n_1 = 5 and n_2 = 2

Therefore, the electron transition is from the 5th orbit to the 2nd orbit and it, therefore, corresponds to the Balmer series. The wave number $(\bar{\nu})$ of the transition is:

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

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

= 2.303 x 10⁶ m⁻¹

Wavelength (λ) of the emitted radiation is:

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

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

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

= 434 nm

Q.57. Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and another type of material. If the velocity of the electron in this microscope is 1.6 x 10⁶ ms⁻¹, calculate de Broglie wavelength associated with this electron.

Ans.

As per de Broglie's equation,

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

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

= 4.55 x 10⁻¹⁰ m

$$\lambda = 455 pm$$

Therefore, de Broglie wavelength of the electron = 455 pm.

Q.58. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.

Ans.

From de Broglie's equation,

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

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

Where,

v denotes the velocity of the neutron

h is Planck's constant

m is the mass of the neutron

λ is the wavelength

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

$$\nu = \frac{(6.626 \times 10^{-34})}{(1.67493 \times 10^{-27})(800 \times 10^{-12} m)} = 4.94 \times 10^{2} J$$

= 494 ms⁻¹

Therefore, the velocity associated with the neutron is 494 ms⁻¹

Q.59. If the velocity of the electron in Bohr's first orbit is 2.19 x 10⁶ ms⁻¹, calculate the de Broglie wavelength associated with it.

Ans.

As per de Broglie's equation,

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

Where, λ is the wavelength of the electron

h is Planck's constant

m is the mass of the electron

v denotes the velocity of electron

Substituting these values in the expression of λ:

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

$$= \frac{(6.626 \times 10^{-34})}{9.103939 \times 10^{-31} kg (2.19 \times 10^6 ms^{-1})}$$

= 3.32 x 10⁻¹⁰ m

$$\lambda = 332 pm$$

Q.60. The velocity associated with a proton moving in a potential difference of 1000 V is 4.37 x 10⁵ ms⁻¹. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

Ans.

As per de Broglie's expression,

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

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

= 1.516 x 10⁻³⁸ m

Q.61. If the position of the electron is measured within an accuracy of ± 0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is h/4π_m × 0.05 nm, is there any problem in defining this value.

Ans.

As per Heisenberg's uncertainty principle, $\Delta x.\Delta p = h/4\pi$

$\Delta p = (1/\Delta x)(h/4)$

Where,

Δx = uncertainty in the position of the electron

Δp = uncertainty in the momentum of the electron

Substituting the given values in the expression for Heisenberg's uncertainty principle :

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

$$= \frac{1}{2 \times 10^{-12}m} \times \frac{6.626 \times 10^{-34}Js}{4 \times 3.14}$$

= 2.637 × 10⁻²³ Jsm⁻¹

Δp = 2.637 × 10⁻²³ kg.m.s⁻¹ (1 J = 1 kgm²s⁻²)

Hence, uncertainty in the momentum of the electron = 2.637 × 10⁻²³ 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 × 10⁻²⁴ kg.m.s⁻¹

The value cannot be defined because the magnitude of the actual momentum is much smaller than the uncertainty.

Q.62: The quantum numbers of six electrons are given below. Arrange them in order of increasing energies. If any of these combination(s) has/have the same energy lists:

1. n = 4, l = 2, ml = –2 , m$_s$ = –1/2
2. n = 3, l = 2, ml= 1 , m$_s$ = +1/2
3. n = 4, l = 1, ml = 0 , m$_s$ = +1/2
4. n = 3, l = 2, ml = –2 , m$_s$ = –1/2
5. n = 3, l = 1, ml = –1 , m$_s$= +1/2
6. n = 4, l = 1, ml = 0 , m$_s$ = +1/2

Ans.

Electrons 1, 2, 3, 4, 5, and 6 reside in the 4d, 3d, 4p, 3d, 3p, and 4p orbitals (respectively). Ranking these orbitals in the increasing order of energies: (3p) < (3d) < (4p) < (4d).

Q.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 electron experiences the lowest effective nuclear charge?

Ans.

The nuclear charge that is experienced by electrons (which are present in atoms containing multiple electrons) depends on the distance between its orbital and the nucleus of the atom. The greater the distance, the lower the effective nuclear charge. Among p-orbitals, 4p orbitals are the farthest from the nucleus of the bromine atom with (+35) charge. Hence, the electrons that reside in the 4p orbital are the ones to experience the lowest effective nuclear charge. These electrons are also shielded by electrons that are present in the 2p and 3p orbitals along with the s-orbitals.

Q.64. Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge?

(i) 2s and 3s,

(ii) 4d and 4f,

(iii) 3d and 3p

Ans.

The nuclear charge can be defined as the net positive charge that acts on an electron in the orbital of an atom that has more than 1 electrons. It is inversely proportional to the distance between the orbital and the nucleus.

(i)Electrons that reside in the 2s orbital are closer to the nucleus than those residing in the 3s orbital and will, therefore, experience greater nuclear charge.

(ii)4d orbital is closer to the nucleus than 4f orbital and will, therefore, experience greater nuclear charge.

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

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

Ans.

The nuclear charge can be defined as the net positive charge that acts on an electron in the orbital of an atom that has more than 1 electrons. The greater the atomic number, the greater the nuclear charge. Silicon holds 14 protons while aluminium holds only 13. Therefore, the nuclear charge of silicon is greater than that of aluminium, implying that the electrons in the 3p orbital of silicon will experience a more effective nuclear charge than aluminium.

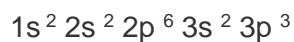**Q.66. Indicate the number of unpaired electrons in:**

(a)P

(b)Si

(c)Cr

(d)Fe

(e)Kr

Ans.

(a)Phosphorus (P):

The atomic number of phosphorus is 15

Electronic configuration of Phosphorus:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

This can be represented as follows:

From the diagram, it can be observed that phosphorus has three unpaired electrons.

(b) Silicon (Si):

The atomic number of Silicon is 14

Electronic configuration of Silicon:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

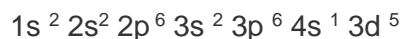This can be represented as follows:

From the diagram, it can be observed that silicon has two unpaired electrons.

(c) Chromium (Cr):

The atomic number of Cr is 24

Electronic configuration of Chromium:

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

This can be represented as follows:

From the diagram, it can be observed that chromium has six unpaired electrons.

(d) Iron (Fe):

The atomic number of iron is 26

Electronic configuration of Fe:

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

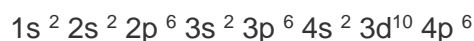This can be represented as follows:

From the diagram, it can be observed that iron has four unpaired electrons.

(e) Krypton (Kr):

The atomic number of Krypton is 36

Its electronic configuration is:

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

This can be represented as follows:

From the diagram, it can be observed that krypton has no unpaired electrons.

Q.67.

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

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

Ans.

(a)n = 4 (Given)

For a given value of 'n', the values of 'l' range from 0 to (n – 1).

Here, the possible values of l are 0, 1, 2, and 3

Therefore, 4 subshells are associated with n = 4, which are s, p, d and f.

(b) Number of orbitals in the nth shell = n^2

For n = 4

Therefore, number of orbitals when n = 4 is 16

Each orbital has 1 electron with m$_s$ value of -1/2.

Therefore, total number of electrons present in the subshell having m$_s$ value of ($-\frac{1}{2}$) is 16.