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:

1 electron weighs $9.109*10^{-31}$ kg. Therefore, number of electrons that weigh 1 g (10^{-3} kg) =

10^{-3}kg/$9.109*10^{-31}$ kg = $1.098*10^{27}$ electrons

(ii)

Mass of one mole of electrons = N_A* mass of one electron

= $(6.022*10^{23})*(9.109*10^{-31}$ kg) = $5.48*10^{-7}$ kg

Charge on one mole of electrons = N_A* charge of one electron

= $(6.022*10^{23})*(1.6022*10^{-19}$ C) = $9.65*10^{4}$ C

Q.2. (i) Calculate the total number of electrons present in one mole of methane.
(ii) Find (a) the total number and (b) the total mass of neutrons in 7 mg of ^{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 $10*N_A = 6.022*10^{24}$ electrons.

(ii) Number of neutrons in 14g (1 mol) of $^{14}C = 8*N_A = 4.817*10^{24}$ neutrons.

Number of neutrons in 7 mg (0.007g) = $(0.007/14)*4.817*10^{24} = 2.409*10^{21}$ neutrons.

Mass of neutrons in 7 mg of $^{14}C = (1.67493*10^{-27}kg)*(2.409*10^{21}) = 4.03*10^{-6}kg$

(iii) Molar mass of NH_3 = 17g

Number of protons in 1 molecule of NH_3 = 7+3 = 10

Therefore, 1 mole (17 grams) of NH_3 contains $10*N_A = 6.022*10^{24}$ protons.

34 mg of NH_3 contains $(34/1700)*6.022*10^{24}$ protons = $1.204*10^{22}$ protons.

Total mass accounted for by protons in 34 mg of $NH_3 = (1.67493*10^{-27}kg)*(1.204*10^{22}) = 2.017*10^{-5}kg$.

These values remain constant regardless of any change in temperature and pressure (since these factors do not affect the number of protons in the atom and the mass of each proton).

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 of carbon-13 = 13

Atomic number of carbon = Number of protons in one carbon atom = 6

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

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

Mass number of oxygen-16 = 16

Atomic number of oxygen = Number of protons = 8

Therefore, No. neutrons = Mass number – Atomic number = 16 – 8 = 8

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

Mass number = 24

Atomic number = No. protons = 12

No. neutrons = Mass number – Atomic number = 24 – 12 = 12

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

Mass nubmer = 56

Atomic number of iron = No. protons in iron= 26

No. neutrons = Mass number – Atomic number = 56 – 26 = 30

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

Mass number = 88

Atomic number = No. protons = 38

No. 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) $^{35}_{17}\mathrm{C}$

(II) $^{233}_{92}\mathrm{U}$

(III) $^{9}_{4}\mathrm{Be}$

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,

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

Here, ν denotes the frequency of the yellow light

c denotes the speed of light ($3 \times 10^8 \, m/s$)

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

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:

$5.17 \times \ 10^{14} \, s^{-1}$

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

Q.6. Find the energy of each of the photons which
(i) correspond to light of frequency 3×10^{15} 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 $6.626 \times \ 10^{-34} \, Js \ \ \nu$ (frequency of the light) =

$3 \times \ 10^{15}$ 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 of a photon whose wavelength is (λ) is:

E= $hc\nu$

Where,

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

c (speed of light) = $3 \times \ 10^{8} \, m/s$

Substituting these values in the equation for 'E':

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

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

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 (ν) = $\frac{1}{Period}$ $\frac{1}{Period}$

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

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

Wavelength of the light wave(λ) = $c\nu$

Where,

c denotes the speed of light, $3\times\ 10^{8}\ m/s$

Substituting the value of 'c' in the previous expression for λ:

$$\lambda = \frac{3\times10^{8}}{5.0\times10^{9}} = 6.0\times\ 10^{-2}m$$

Wave number $(\bar{\nu})$ of light = $\frac{1}{\lambda} = \frac{1}{6.0\times10^{-2}} = 1.66\times\ 10^{1}\ m^{-1} = 16.66\,m^{-1}$

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

Ans: Energy of one photon (E) = $h\nu$

Energy of 'n' photons (E_n) = $nh\nu\ \Rightarrow\ n = \frac{E_n\lambda}{hc}$

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

c denotes the speed of light in vacuum = $3 \times\ 10^{8}\, m/s$

h is Planck's constant, whose value is $6.626 \times\ 10^{-34}\, Js$

Substituting these values in the expression for 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 $2.012 \times\ 10^{16}$

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 $6.626 \times\ 10^{-34}\, Js$

c denotes the speed of light = $3 \times\ 10^{8}\, m/s$ λ = wavelength of the photon = $4 \times\ 10^{-7}\, m/s$

Substituting these values in the expression for E:

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

Therefore, energy of the photon = $4.97 \times\ 10^{-19}\, J$

(ii)

The kinetic energy of the emission E_k can be calculated as follows:

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

$= (E - W)eV$

$= (\frac{4.9695\times10^{-19}}{1.6020\times10^{-19}})eV - 2.13\,eV$

$= (3.1020 - 2.13)eV$

$= 0.9720\,eV$

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

(iii)

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

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

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

Where $(h\nu - h\nu_0)$ is the K.El of the emission (in Joules) and 'm' denotes the mass of the photoelectron.

Substituting thes values in the expression for v:

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

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

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

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

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

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

$= 4.947 \times 10^5 \, J \, mol^{-1}$

$= 494.7 \times 10^3 \, J \, mol^{-1}$

$= 494 \, kJ \, mol^{-1}$

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 = $25 \, J s^{-1}$

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

Substituting these values in the expression for E:

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

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

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

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 (v_0) and work function (W_0) of the metal.

Ans: Threshold wavelength of the radiation (λ_0) = 6800 Å= $6800 \times 10^{-10} \, m$

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

Therefore, threshold frequency (ν_0) of the metal = $h\nu_0$

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

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

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 to 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:

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

Substituting these values in the expression for E:

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

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

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

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

Here, the -ve sign denotes the emitted energy.

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

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

Substituting these values in the expression for $lambda$

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

$$\lambda = 486.31 \times 10^{-9} m$$

$$= 486 nm$$

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,

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

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

For the ionization from $n_t = 5$ 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 = ∞ = Energy required for n1 = 1 to n = ∞, is $8.72 \times 10^{-20}\ m$ $\Delta E = E_\infty - E_5$

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

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

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

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

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?

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

Total no. of spectral lines emitted when an electron initially in the ‘nth‘ level drops down to the ground state can be calculated using the following expression:

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

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

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

Ans:

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

$$E_5 = \frac{-(2.18\times10^{-18})}{(5)^2} = \frac{-(2.18\times10^{-18})}{25} = -8.72\times 10^{-20}\, Jatom^{-1}$$

(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:

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

Since wave number($\bar{\nu}$) is inversely proportional to the transition wavelength, the lowest possivle 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:

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

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

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

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

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

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.

The ground-state electron energy is –2.18 × 10^{-11} ergs.

Ans. Energy (E) associated with the n^{th} Bohr orbit of an atom is:

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

Where, Z denotes the atom's atomic number

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

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

The required energy for an electron shift from n = 1 to n = 5 is:

$\Delta E = E_5 - E_1$

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

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

$= (2.18 \times 10^{-18})[\frac{24}{25}]$

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

The wavelength of the emitted light = $\frac{hc}{E}$ $E = \frac{(6.626\times10^{-34})(3\times10^{8})}{(2.0928\times10^{-18})}$

$= 9.498 \times 10^{-8}\, m$ = 950 Å

Q.19. The electron energy in hydrogen atom is given by En= (–2.18 × 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?

Ans. $E_n = \frac{-(2.18\times10^{-18})}{(n)^2} J$

Required energy for the ionization from n = 2 is:

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

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

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

$$= 0.545 \times 10^{-18}\, J \;\; \Delta E = 5.45 \times 10^{-19}\, J \;\; \lambda = \frac{hc}{\Delta E}$$

If λ is the longest wavelength that can cause this transition,

$$E = \frac{(6.626\times10^{-34})(3\times10^{8})}{(0.57\times10^{-6})} = 34.87 \times 10^{-20}\, J \;\; E = \frac{(6.626\times10^{-34})(3\times10^{8})}{(5.45\times10^{-19})} = 3.647 \times 10^{-7}\, m$$

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

= 3647 Å

Q.20. Calculate the wavelength of an electron moving with a velocity of 2.05 × 10^7 m s^{-1}

Ans. As per de Broglie's equation,

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

Where, λ denotes thr 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 in the expression for λ:

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

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

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

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. As per de Broglie's equation,

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

Given, K.E of electron = 3.0×10^{-25} J

Since K.E.= $\frac{1}{2}mv^2 \therefore Velocity(v) = \sqrt{\frac{2K.E}{m}}$

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

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

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

Substituting these values in the expression for λ:

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

$= 8.9625 \times 10^{-7} m$ = 8967 Å

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

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

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. $\therefore$

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 species has gained one electron (accounted for by the -1 charge). $\therefore$ Electron 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 no. 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 no. 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 no. 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, no.protons = no.electrons. ∴ No. 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^+ and H_2 and O_2+

Ans: No. electrons present in H_2 = 1 + 1 = 2. ∴ Number of electrons in H_2^+ = 2 – 1 = 1

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

No. electrons 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) $n = 2, l = 0, m_l = 1, m_s = -\frac{1}{2}$

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

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

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) Not possible. Since, when l = 0, m_t cannot be 1.

(e) Not possible. As when 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 = $2n^2$

$\therefore$ For n = 4, Total no. electrons = $2(4)^2 = 32$

Electron configuration for an atom with 32 electrons: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$.

Hence, all the electrons are paired.

$\therefore$ No. electrons (having n = 4 and $m_s = -\frac{1}{2}$) = 16

(b)n = 3, l = 0 indicates the 3s orbital. Therefore, no. 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 atoms 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:

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

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

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

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:

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

Where, n_1 = 2 n_2 = 4

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

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

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

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

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

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

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

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

The transition for n_2 = 2 to n = 1 in the hydrogen spectrum would, therefore, 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 $2.18 \times 10^{-18} J\, atom^{-1}$

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

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

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 process given by:

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

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

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

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

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

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

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

Space occupied by one carbon atom $0.15 \times 10^{-9} m$.

$\therefore$ No. carbon atoms that can be placed in a straight line $\frac{20 \times 10^{-2} m}{0.15 \times 10^{-9} m}$

$= 133.33 \times 10^{7}$

$= 1.33 \times 10^{9}$

Q.36. $2 \times 10^{8} m$ **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

No. carbon atoms present = 2 × 10^{8}

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

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

$= \frac{1.2\times10^{-10}m}{2}$

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

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

$= \frac{2.6}{2}$

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

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

(b) Length of the arrangement = 1.6 cm

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

Diameter of a zinc atom = $1.6\times\ 10^{-10}m$ $\therefore$ No. zinc atoms in the arrangement

$= \frac{1.6\times10^{-2}m}{2.6\times10^{-10}m}$

$= 0.6153\times\ 10^{8}m$

$= 6.153\times\ 10^{7}$

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 held by one electron = 1.6022 x 10^{-19} C ⇒ 1.6022 x 10^{-19} C charge is held by one electron

Therefore, No. electrons that carry a charge of $2.5 \times 10^{-16}C$ $\frac{1}{1.6022 \times 10^{-19}C}(2.5 \times 10^{-16}C)$

$= 1.560 \times 10^{3}C$

$= 1560\ C$

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 held by the oil drop = $1.282 \times 10^{-18}C$

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

Therefore, No. electrons present in the drop of oil

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

$= 0.8001 \times 10^{1}$

$= 8.0$

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.

The results obtained when a foil of heavy atoms will be different from the results obtained when relatively light atoms are used in the foil. The lighter the atom, the lower the magnitude of positive charge in its nucleus. Therefore, lighter atoms will not cause enough deflection of the positively charged α-particles.

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

Answer briefly.

Ans.

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

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

$^{79}\mathrm{Br}$ is an acceptable representation but $^{35}\mathrm{Br}$ 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.

Let the No. protons in the element be x.

Therefore, No. neutrons in the element = x + 31.7% of x

= x + 0.317 x

= 1.317 x

Given, Mass number of the element = 81, which implies that (No. protons + No. neutrons) = 81

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

$$2.317x = 81$$

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

$$= 34.95$$

$$x \simeq 35$$

Therefore, total no. protons = 35, which implies that atomic number = 35.

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

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 no. electrons in the negatively charged ion be x.

Then, no. neutrons present = x + 11.1% of x = x + 0.111 x = 1.111 x

No. electrons present in the neutral atom = (x – 1)

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

No. protons present in the neutral atom = x – 1

Given, mass number of the ion = 37

(x – 1) + 1.111x = 37

2.111x = 38

x = 18

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

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 no. electrons present in A^{3+} be x. Now, total no. neutrons in it = x + 30.4% of x = 1.304 x

Since the ion has a charge of +3 , $\Rightarrow$ no. electrons in neutral atom = x + 3

Therefore, no. protons in neutral atom = x + 3

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, no. protons = x + 3 = 23 + 3 = 26

The ion is $^{56}_{26}\mathrm{Fe}^{3+}$

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 ovens < amber light < radiation of 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 10^{24}, calculate the power of this laser.

Ans.

Power of laser = Energy with which it emits photons

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

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):

$$E = \frac{(5.6\times10^{24})(6.626\times10^{-34}Js)(3\times10^{8}ms^{-1})}{(337.1\times10^{-9}m)}$$

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

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

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

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 = $616 \times 10^{-9} m$ (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 for (ν):

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

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

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

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

(b) Speed of the radiation,c = $3 \times 10^{8} ms^{-1}$

Distance travelled by the radiation in a timespan of 30 s

$= (3 \times 10^8 ms^{-1})(30s)$

$= 9 \times 10^9 m$

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

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

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

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

(d) No. quanta in 2 J of energy

$\frac{2J}{32.27 \times 10^{-20} J}$

$= 6.19 \times 10^{18}$

$= 6.2 \times 10^{18}$

Q.48. In astronomical observations, signals observed from the distant stars are generally weak. If the photon detector receives a total of $3.15 \times 10^{-18} J$ 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 for E:

$$E = \frac{(6.626\times10^{-34}Js)(3\times10^{8}ms^{-1})}{(600\times10^{-9})} = 3.313\times\ 10^{-19}\ J$$

Energy held by one photon = $3.313\times\ 10^{-19}\ J$

No. photons received with $3.15\times\ 10^{-18}\ J$ energy

$$= \frac{3.15\times10^{-18}J}{3.313\times10^{-19}J}$$

$$= 9.5$$

$$\approx 10$$

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 $2.5\times\ 10^{15}\ J$, calculate the energy of the source.

Ans.

Frequency of radiation (ν),

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

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

Energy (E) of source = Nhv

Where,

N is the no. photons emitted

h is Planck's constant

v denotes the frequency of the radiation

Substituting these values in the expression for (E):

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

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

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

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 for (λ_0):

$$\lambda_0 = \frac{(6.626 \times 10^{-34} Js)(3 \times 10^{8} ms^{-1})}{(1.9 \times 1.602 \times 10^{-19} J)} = 6.53 \times 10^{-7} m$$

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 for ν_0:

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

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

ν_0 = $4.593 \times 10^{14} s^{-1}$

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

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

Kinetic energy = h (ν – ν_0)

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

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

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

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

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

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

Since K.E= $\frac{1}{2}mv^2 = 9.3149 \times 10^{-20}J$ $\nu = \sqrt{\frac{2(9.3149\times10^{-20}J)}{9.10939\times10^{-31}}}$

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

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

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

Ans.

(a) If the threshold wavelength is $\lambda_0\, nm(= \lambda_0 \times 10^{-9}m)$, the K.E. of the radiation would be:

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

Three equations can be formed by these values:

$h(\frac{1}{\lambda} - \frac{1}{\lambda_0}) = \frac{1}{2}m\nu^2$ $hc(\frac{1}{500\times10^9} - \frac{1}{\lambda_0\times10^{-9}m}) = \frac{1}{2}m(2.55 \times 10^{+5} \times 10^{-2}ms^{-1})$

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

Similarly,

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

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

Dividing equation (3) by equation (1):

$\frac{[\frac{\lambda_0-400}{400\lambda_0}]}{[\frac{\lambda_0-500}{500\lambda_0}]} = \frac{(5.35\times10^{+3}ms^{-1})^2}{(2.55\times10^{+3}ms^{-1})^2}$ $\frac{5\lambda_0-2000}{4\lambda_0-2000} = \frac{(5.35)^2}{(2.55)^2} = \frac{28.6225}{6.5025}$

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

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

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

$\lambda_0 = 539.8 nm$

$\lambda_0 = 540 nm$

Therefore, the threshold wavelength (λ_0) is 540 nm

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

⇒ 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 for E:

$E = \frac{(6.626\times10^{-34} Js)(3\times10^{8} ms^{-1})}{(256.7\times10^{-9} m)} = 7.744 \times 10^{-19} J$

$= \frac{7.744\times10^{-19}}{1.602\times10^{-19}} eV$

$E = 4.83\, eV$

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 10^7 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)

$$\frac{1}{2}m_e\nu^2$$

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

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

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

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

= E – K.E

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

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

$= 7.6 \times 10^{3} eV$

Q. 55. Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as v = 3.29 x 10^{15} (Hz) [1/3^2 – 1/n^2] Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

Ans.

Wavelength of the transition = 1285 nm

= $1285 \times 10^{-9} m$ (Given)

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

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

$= \frac{3 \times 10^8 ms^{-1}}{1285 \times 10^{-9} m}$

Now, v = $2.33 \times 10^{14} s^{-1}$

Substituting the value of v, the equation becomes:

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

$\frac{1}{9} - \frac{1}{n^2} = \frac{2.33 \times 10^{14}}{3.29 \times 10^{15}}$

$\frac{1}{9} - 0.7082 \times 10^{-1} = \frac{1}{n^2}$

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

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

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

n=4.98 n=5

Therefore, in order to observe this transition at 1285 nm, n = 5. The spectrum must lie the infra-red region.

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

Ans.

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

$r = \frac{0.529n^2}{Z}$ Å

$r = \frac{5.29n^2}{Z} pm$

For radius (r_1) = 1.3225 nm

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

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

$= 1322.25\, pm \;\; n_1^2 = \frac{r_1 Z}{52.9}$

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

Similarly,

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

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

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

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

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

⇒ n1 = 5 and n2 = 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:

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

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

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

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

Ans.

As per de Broglie's equation,

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

$$= \frac{(6.626 \times 10^{-34})}{9.103939 \times 10^{-31} kg(1.6 \times 10^{6} ms^{-1})}$$

$$= 4.55 \times 10^{-10} m \lambda = 455 pm$$

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),

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

$$= 494 \, ms^{-1}$$

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 10^{6} ms^{-1}, calculate the de Broglie wavelength associated with it.

Ans.

As per de Broglie's equation,

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

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 for λ:

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

$$= \frac{(6.626\times10^{-34})}{9.103939\times10^{-31}\,kg(2.19\times10^{6}\,ms^{-1})}$$

$$= 3.32 \times\ 10^{-10} m \lambda =\ 332 pm$$

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

Ans.

As per de Broglie's expression,

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

$$= \frac{(6.626\times10^{-34})}{0.1kg(4.37\times10^{5}\,ms^{-1})}$$

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

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, $\Delta x.\Delta p >= h/4\pi$

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 :

$= 2.637 \times 10^{-23}$ Jsm^{-1}

$\Delta p = 2.637 \times 10^{-23}$ kg.m.s^{-1} (1 J = 1 kgm^{2}s^{-2})

Uncertainty in the momentum of the electron = 2.637×10^{-23} kg.m.s^{-1} = 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 greater magnitude of effective nuclear charge.

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.

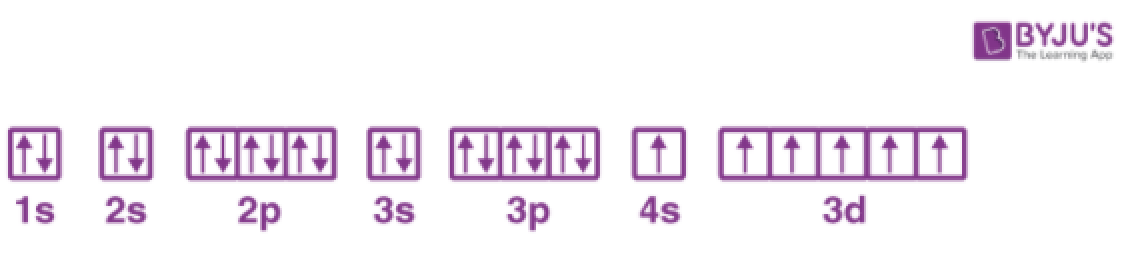
(c) Chromium (Cr):

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.

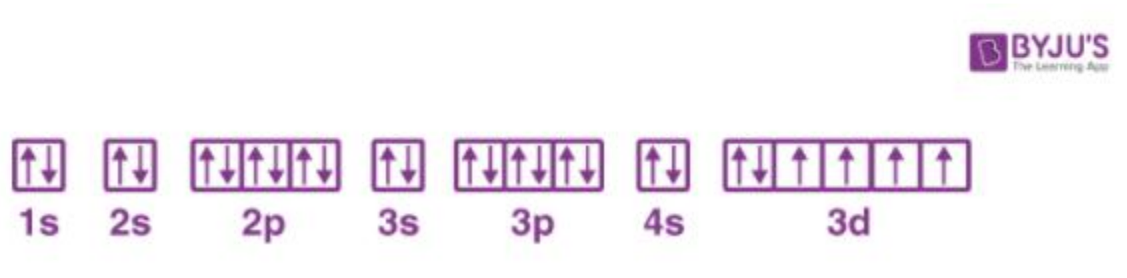
(d) Iron (Fe):

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 some 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, a total of 4 subshells are possible when n=4: the s, p, d and f subshells.

(b) No. orbitals in the n^{th} shell = n^2

For n = 4

Therefore, the total no. orbitals when n=4 is 16

If each orbital fully occupied, each orbital will have 1 electron with m_s value of -1/2.

Therefore, total no. electrons with an m_s value of (-1/2) is 16.