Question 2.1:

(i) Calculate the number of electrons which will together weigh one gram.

(ii) Calculate the mass and charge of one mole of electrons.

Answer

(i) Mass of one electron = 9.10939×10^{-31} kg

∴ Number of electrons that weigh 9.10939×10^{-31} kg = 1

Number of electrons that will weigh 1 g = (1×10^{-3} kg)

$$= \frac{1}{9.10939 \times 10^{-31}\ \text{kg}} \times \left(1 \times 10^{-3}\ \text{kg}\right)$$

$= 0.1098 \times 10^{-3 + 31}$

$= 0.1098 \times 10^{28}$

$= 1.098 \times 10^{27}$

(ii) Mass of one electron = 9.10939×10^{-31} kg

Mass of one mole of electron = $(6.022 \times 10^{23}) \times (9.10939 \times 10^{-31}$ kg)

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

Charge on one electron = 1.6022×10^{-19} coulomb

Charge on one mole of electron = (1.6022×10^{-19} C) (6.022×10^{23})

$= 9.65 \times 10^{4}$ C

Question 2.2:

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

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

(Assume that mass of a neutron = 1.675×10^{-27} kg).

(iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3 at STP.

Will the answer change if the temperature and pressure are changed?

Answer

(i) Number of electrons present in 1 molecule of methane (CH_4)

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

Number of electrons present in 1 mole i.e., 6.023×10^{23} molecules of methane

$= 6.022 \times 10^{23} \times 10 = 6.022 \times 10^{24}$

(ii) (a) Number of atoms of ^{14}C in 1 mole= 6.023×10^{23}

Since 1 atom of ^{14}C contains (14 – 6) i.e., 8 neutrons, the number of neutrons in 14 g of ^{14}C is $(6.023 \times 10^{23}) \times 8$. Or, 14 g of ^{14}C contains $(6.022 \times 10^{23} \times 8)$ neutrons.

Number of neutrons in 7 mg

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

$= 2.4092 \times 10^{21}$

(b) Mass of one neutron = 1.67493×10^{-27} kg

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

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

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

(iii) (a) 1 mole of NH_3 = {1(14) + 3(1)} g of NH_3

= 17 g of NH_3

$= 6.022\times 10^{23}$ molecules of NH_3

Total number of protons present in 1 molecule of NH_3

= {1(7) + 3(1)}

= 10

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

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

$= 6.023 \times 10^{24}$

⇒ 17 g of NH_3 contains (6.023×10^{24}) protons.

Number of protons in 34 mg of NH^3

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

$= 1.2046 \times 10^{22}$

(b) Mass of one proton = 1.67493×10^{-27} kg

Total mass of protons in 34 mg of NH_3

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

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

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

Question 2.3:

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

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

Answer

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

Atomic mass = 13

Atomic number = Number of protons = 6

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

= 13 – 6 = 7

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

Atomic mass = 16

Atomic number = 8

Number of protons = 8

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

= 16 – 8 = 8

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

Atomic mass = 24

Atomic number = Number of protons = 12

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

= 24 – 12 = 12

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

Atomic mass = 56

Atomic number = Number of protons = 26

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

= 56 – 26 = 30

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

Atomic mass = 88

Atomic number = Number of protons = 38

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

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= 88 – 38 = 50

Question 2.4:

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

(i) Z = 17, A = 35

(ii) Z = 92, A = 233

(iii) Z = 4, A = 9

Answer

(i) ${}^{35}_{17}\text{Cl}$ (ii) ${}^{233}_{92}\text{U}$ (iii) ${}^{9}_{4}\text{Be}$

Question 2.5:

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

Answer

From the expression,

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

We get,

$$\nu = \frac{c}{\lambda}$$ ……. (i)

Where,

ν = frequency of yellow light

c = velocity of light in vacuum = 3×10^8 m/s

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

Substituting the values in expression (i):

$$\nu = \frac{3\times10^8}{580\times10^{-9}} = 5.17\times10^{14}\ \text{s}^{-1}$$

Thus, frequency of yellow light emitted from the sodium lamp

= 5.17×10^{14} s^{-1}

Wave number of yellow light, $\bar{\nu} = \frac{1}{\lambda}$

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

Question 2.6:

Find energy of each of the photons which

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

(ii) have wavelength of 0.50 Å.

Answer

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

$E = h\nu$

Where,

h = Planck's constant = 6.626×10^{-34} Js

ν = frequency of light = 3×10^{15} Hz

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

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

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

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

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

h = Planck's constant = 6.626×10^{-34} Js

c = velocity of light in vacuum = 3×10^{8} m/s

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

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

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

Question 2.7:

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

Answer

Frequency (ν) of light $= \dfrac{1}{\text{Period}}$

$$= \frac{1}{2.0\times10^{-10}\ \text{s}} = 5.0\times10^{9}\ \text{s}^{-1}$$

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

Where,

c = velocity of light in vacuum = 3×10^{8} m/s

Substituting the value in the given expression of λ:

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

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

Question 2.8:

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

Answer

Energy (E) of a photon = hν

Energy (E_n) of 'n' photons = nhν

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

Where,

λ = wavelength of light = 4000 pm = 4000 $\times10^{-12}$ m

c = velocity of light in vacuum = 3 × 10^{8} m/s

h = Planck's constant = 6.626 × 10^{-34} Js

Substituting the values in the given expression of n:

$$n = \frac{(1)\times\left(4000\times10^{-12}\right)}{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)} = 2.012\times10^{16}$$

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

Question 2.9:

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

Answer

(i) Energy (E) of a photon = hν $= \frac{hc}{\lambda}$

Where,

h = Planck's constant = 6.626 × 10^{-34} Js

c = velocity of light in vacuum = 3 × 10^{8} m/s

λ = wavelength of photon = 4 × 10^{-7} m

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

$$E = \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{4\times10^{-7}} = 4.9695\times10^{-19}\ \text{J}$$

Hence, the energy of the photon is 4.97 × 10^{-19} J.

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

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

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

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

= (3.1020 – 2.13) eV

= 0.9720 eV

Hence, the kinetic energy of emission is 0.97 eV.

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

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

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

Where, $(h\nu - h\nu_0)$ is the kinetic energy of emission in Joules and '*m*' is the mass of the photoelectron. Substituting the values in the given expression of v:

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

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

v = 5.84 × 10^5 ms^{-1}

Hence, the velocity of the photoelectron is 5.84 × 10^5 ms^{-1}.

Question 2.10:

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

Answer

Energy of sodium (E) $= \dfrac{N_A hc}{\lambda}$

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

= 4.947 × 10^5 J mol^{-1}

= 494.7 × 10^3 J mol^{-1}

= 494 kJ mol^{-1}

Question 2.11:

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

Answer

Power of bulb, P = 25 Watt = 25 Js^{-1}

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

Substituting the values in the given expression of E:

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

E = 34.87 × 10^{-20} J

Rate of emission of quanta per second

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

Question 2.12:

Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (ν_0) and work function (W_0) of the metal.

Answer

Threshold wavelength of radian $(\lambda_0) = 6800\text{Å}$ = 6800 × 10^{-10} m

Threshold frequency (ν_0) of the metal

$$= \frac{c}{\lambda_0} = \frac{3\times10^8 \text{ ms}^{-1}}{6.8\times10^{-7}\text{m}}$$ = 4.41 × 10^{14} s^{-1}

Thus, the threshold frequency (ν_0) of the metal is 4.41 × 10^{14} s^{-1}.

Hence, work function (W_0) of the metal = $h\nu_0$

= (6.626 × 10^{-34} Js) (4.41 × 10^{14} s^{-1})

= 2.922 × 10^{-19} J

Question 2.13:

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

Answer

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

$$E = 2.18\times10^{-18}\left[\frac{1}{n_i^2} - \frac{1}{n_f^2}\right]$$

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

$$E = 2.18\times10^{-18}\left[\frac{1}{4^2} - \frac{1}{2^2}\right]$$

$$= 2.18\times10^{-18}\left[\frac{1-4}{16}\right]$$

$$= 2.18\times10^{-18}\times\left(-\frac{3}{16}\right)$$

E = – (4.0875 × 10^{-19} J)

The negative sign indicates the energy of emission.

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

$\left(\text{since } E = \dfrac{hc}{\lambda}\right)$

Substituting the values in the given expression of λ:

$$\lambda = \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{4.0875\times10^{-19}}$$

$$\lambda = 4.8631\times10^{-7}\ \text{m}$$
$$= 486.3\times10^{-9}\,\text{m}$$
$$= 486\ \text{nm}$$

Question 2.14:

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

Answer

The expression of energy is given by,

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

Where,

Z = atomic number of the atom

n = principal quantum number

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

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

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

$$= \left(2.18\times10^{-18}\ \text{J}\right)\left(\frac{1}{(5)^2}\right) \qquad \left(\text{Since } \frac{1}{\infty} = 0\right)$$

$$= 0.0872\times10^{-18}\ \text{J}$$

$$\Delta E = 8.72\times10^{-20}\ \text{J}$$

Hence, the energy required for ionization from n = 5 to n = ∞ is 8.72 × 10^{-20} J.

Energy required for n_1 = 1 to n = ∞,

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

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

$$= \left(2.18\times10^{-18}\right)[1-0]$$

$$= 2.18\times10^{-18}\ \text{J}$$

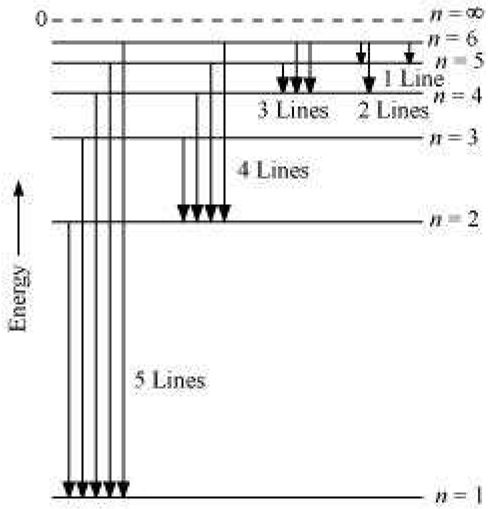
Hence, less energy is required to ionize an electron in the 5th orbital of hydrogen atom as compared to that in the ground state.

Question 2.15:

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

Answer

When the excited electron of an H atom in n = 6 drops to the ground state, the following transitions are possible:



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

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

Given,

$n = 6$

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

Question 2.16:

(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 hydrogen atom.

Answer

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

$$E_5 = \frac{-\left(2.18\times10^{-18}\right)}{(5)^2} = \frac{-2.18\times10^{-18}}{25}$$

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

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

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

For,

$n = 5$

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

$r_5 = 1.3225 \text{ nm}$

Question 2.17:

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

Answer

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

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

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

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

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

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

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

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

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

Question 2.18:

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.

Answer

Energy (E) of the nth Bohr orbit of an atom is given by,

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

Where,

Z = atomic number of the atom

Ground state energy = – 2.18 × 10^{-11} ergs

= –2.18 × 10^{-11} × 10^{-7} J

= – 2.18 × 10^{-18} J

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

$\Delta E = E_5 - E_1$

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

$$= (2.18\times10^{-18})\left[1-\frac{1}{25}\right]$$

$$= (2.18\times10^{-18})\left(\frac{24}{25}\right) = 2.0928\times10^{-18}\text{ J}$$

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

$$= \frac{(6.626\times10^{-34})(3\times10^8)}{(2.0928\times10^{-18})}$$

$$= 9.498\times10^{-8}\text{ m}$$

Question 2.19:

The electron energy in hydrogen atom is given by E_n = (–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?

Answer

Given,

$$E_n = -\frac{2.18\times10^{-18}}{n^2}\text{J}$$

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

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

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

$$= \left[\frac{2.18\times10^{-18}}{4} - 0\right] \text{J}$$

= 0.545×10^{-18} J

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

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

Here, λ is the longest wavelength causing the transition.

$$\lambda = \frac{(6.626\times10^{-34})(3\times10^{8})}{5.45\times10^{-19}} = 3.647\times10^{-7}\,\text{m}$$

= 3647×10^{-10} m

= 3647 Å

Question 2.20:

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

Answer

According to de Broglie's equation,

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

Where,

λ = wavelength of moving particle

m = mass of particle

v = velocity of particle

h = Planck's constant

Substituting the values in the expression of λ:

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

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

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

Question 2.21:

The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

Answer

From de Broglie's equation,

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

Given,

Kinetic energy (K.E) of the electron = 3.0×10^{-25} J

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

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

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

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

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

Substituting the value in the expression of *λ:*

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

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

Hence, the wavelength of the electron is 8.9625×10^{-7} m.

Question 2.21:

The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

Answer

From de Broglie's equation,

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

Given,

Kinetic energy (K.E) of the electron = 3.0×10^{-25} J

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

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

$$= \sqrt{\frac{2(3.0\times10^{-25}\text{ J})}{9.10939\times10^{-31}\text{ kg}}}$$

$$= \sqrt{6.5866\times10^{4}}$$

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

Substituting the value in the expression of λ:

$$\lambda = \frac{6.626\times10^{-34}\text{ Js}}{(9.10939\times10^{-31}\text{ kg})(811.579\text{ ms}^{-1})}$$

$$\lambda = 8.9625\times10^{-7}\text{ m}$$

Hence, the wavelength of the electron is 8.9625×10^{-7} m.

Question 2.23:

(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] $4s^2$ $3d^1$.

Answer

(i) (a) H^- ion

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

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

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

(b) Na^+ ion

The electronic configuration of Na atom is $1s^2\ 2s^2\ 2p^6\ 3s^1$.

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

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

(c) O^{2-} ion

The electronic configuration of 0 atom is $1s^2\ 2s^2\ 2p^4$.

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

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

(d) F^- ion

The electronic configuration of F atom is $1s^2\ 2s^2\ 2p^5$.

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

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

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

Completing the electron configuration of the element as

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

∴ Number of electrons present in the atom of the element

= 2 + 2 + 6 + 1 = 11

∴ Atomic number of the element = 11

(b) $2p^3$

Completing the electron configuration of the element as

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

∴ Number of electrons present in the atom of the element = 2 + 2 + 3 = 7

∴ Atomic number of the element = 7

(c) $3p^5$

Completing the electron configuration of the element as

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

∴ Number of electrons present in the atom of the element = 2 + 2 + 5 = 9

∴ Atomic number of the element = 9

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

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

∴ Atomic number of the element = 3

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

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

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

∴ Atomic number of the element = 15

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

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

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

∴ Atomic number of the element = 21

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

Question 2.24:

What is the lowest value of *n* that allows g orbitals to exist?

Answer

For g-orbitals, l = 4.

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

∴ For l = 4, minimum value of n = 5

Question 2.25:

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

Answer

For the 3*d* orbital:

Principal quantum number (n) = 3

Azimuthal quantum number (l) = 2

Magnetic quantum number (m_l) = – 2, – 1, 0, 1, 2

Question 2.26:

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

Answer

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

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

(ii) The electronic configuration of the atom is

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

Question 2.27:

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

Answer

H_2^+:

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

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

H_2:

Number of electrons in H_2 = 1 + 1 = 2

O_2^+:

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

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

Question 2.28:

(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 3*d* orbital.

(iii) Which of the following orbitals are possible?

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

Answer

(i) n = 3 (Given)

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

∴ For n = 3

l = 0, 1, 2

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

For l = 0, m = 0

l = 1, m = – 1, 0, 1

l = 2, m = – 2, – 1, 0, 1, 2

∴ For n = 3

l = 0, 1, 2

m_0 = 0

m_1 = – 1, 0, 1

m_2 = – 2, – 1, 0, 1, 2

(ii) For 3*d* orbital, *l* = 2.

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

∴ For *l* = 2

m_2 = – 2, – 1, 0, 1, 2

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

For *p*-orbital, *l* = 1.

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

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

Similarly,

For *f*-orbital, *l* = 4.

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

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

Question 2.29:

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

(a) *n* = 1, *l* = 0; **(b)** *n* = 3; *l* =1 **(c)** *n* = 4; *l* = 2; **(d)** *n* = 4; *l* =3.

Answer

(a) *n* = 1, *l* = 0 (Given)

The orbital is 1*s.*

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

The orbital is 3*p.*

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

The orbital is 4*d.*

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

The orbital is 4*f.*

Question 2.30:

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

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

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

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

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

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

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

Answer

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

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

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

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

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

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

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

For $n = 3$,

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

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

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

Question 2.31:

How many electrons in an atom may have the following quantum numbers?**(a)** $n = 4$, $m_s = -\frac{1}{2}$ **(b)** $n = 3$, $l = 0$

Answer

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

∴ For $n = 4$,

Total number of electrons = $2 (4)^2$

= 32

The given element has a fully filled orbital as

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

Hence, all the electrons are paired.

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

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

Question 2.32:

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

Answer

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

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

Where,

n = 1, 2, 3, …

According to de Broglie's equation:

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

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

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

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

$$\text{or } 2\pi r = n\lambda...........(3)$$

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

Question 2.33:

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

Answer

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

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

Where,

$n_1 = 2$

$n_2 = 4$

Z = atomic number of helium

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

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

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

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

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

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

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

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

n_1 = 1and n_2 = 2.

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

Question 2.34:

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 2.18×10^{-18} J $atom^{-1}$

Answer

Energy associated with hydrogen-like species is given by,

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

For ground state of hydrogen atom,

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

$$= 0 - \left[-2.18\times10^{-18}\left\{\frac{(1)^2}{(1)^2}\right\}\right]\text{J}$$

$$\Delta E = 2.18\times10^{-18}\ \text{J}$$

For the given process,

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

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

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

$$= 0 - \left[-2.18\times10^{-18}\left\{\frac{(2)^2}{(1)^2}\right\}\right]$$

$$\Delta E = 8.72\times10^{-18}\ \text{J}$$

∴ The energy required for the process is 8.72×10^{-18} J.

Question 2.35:

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

Answer

1 m = 100 cm

1 cm = 10^{-2} m

Length of the scale = 20 cm

= 20×10^{-2} m

Diameter of a carbon atom = 0.15 nm

= 0.15×10^{-9} m

One carbon atom occupies 0.15×10^{-9} m.

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

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

$$= 133.33 \times 10^{7}$$

$$= 1.33 \times 10^{9}$$

Question 2.36:

2×10^{8} atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

Answer

Length of the given arrangement = 2.4 cm

Number of carbon atoms present = 2×10^{8}

∴ Diameter of carbon atom

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

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

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

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

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

Question 2.37:

The diameter of zinc atom is $2.6\,\mathring{A}$.Calculate (a) radius of zinc atom in pm and (b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

Answer

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

$$= \frac{2.6\,\mathring{A}}{2}$$
$$= 1.3 \times 10^{-10}\ \text{m}$$
$$= 130 \times 10^{-12}\ \text{m} = 130\ \text{pm}$$

(b) Length of the arrangement = 1.6 cm

= 1.6×10^{-2} m

Diameter of zinc atom = 2.6×10^{-10} m

∴ Number of zinc atoms present in the arrangement

$$= \frac{1.6 \times 10^{-2}\ \text{m}}{2.6 \times 10^{-10}\ \text{m}}$$
$$= 0.6153 \times 10^{8}\ \text{m}$$
$$= 6.153 \times 10^{7}$$

Question 2.38:

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

Answer

Charge on one electron = 1.6022×10^{-19} C

⇒ 1.6022×10^{-19}C charge is carried by 1 electron.

∴ Number of electrons carrying a charge of 2.5×10^{-16} C

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

Question 2.39:

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

Answer

Charge on the oil drop = 1.282×10^{-18}C

Charge on one electron = 1.6022×10^{-19}C

∴Number of electrons present on the oil drop

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

$$= 0.8001 \times 10^{1}$$

$$= 8.0$$

Question 2.40:

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

Answer

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

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

Question 2.41:

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

Answer

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

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

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

Question 2.42:

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

Answer

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

∴ Number of neutrons in the element

= *x* + 31.7% of *x*

= *x* + 0.317 *x*

= 1.317 *x*

According to the question,

Mass number of the element = 81

∴ (Number of protons + number of neutrons) = 81

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

$$2.317x = 81$$

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

$$= 34.95$$

$$\therefore x \simeq 35$$

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

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

∴ The atomic symbol of the element is $^{81}_{35}Br$.

Question 2.43:

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

Answer

Let the number of electrons in the ion carrying a negative charge be *x.*

Then,

Number of neutrons present

= *x* + 11.1% of *x*

= *x* + 0.111 *x*

= 1.111 *x*

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

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

∴ Number of protons in the neutral atom = *x* – 1

Given,

Mass number of the ion = 37

∴ (*x* – 1) + 1.111*x* = 37

2.111*x* = 38

x = 18

∴The symbol of the ion is $^{37}_{17}Cl^{-}$.

Question 2.44:

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

Answer

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

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

Since the ion is tripositive,

⇒ Number of electrons in neutral atom = *x* + 3

∴ Number of protons in neutral atom = *x* + 3

Given,

Mass number of the ion = 56

$$\therefore (x+3)+(1.304x)=56$$
$$2.304x=53$$
$$x=\frac{53}{2.304}$$
$$x=23$$

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

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

Question 2.45:

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

Answer

The increasing order of frequency is as follows:

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

The increasing order of wavelength is as follows:

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

Question 2.46:

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

Answer

Power of laser = Energy with which it emits photons

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):

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

= 0.3302×10^{7} J

= 3.33×10^{6} J

Hence, the power of the laser is 3.33×10^{6} J.

Question 2.47:

Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate (a) the frequency of emission, (b) distance traveled by this radiation in 30 s (c) energy of quantum and (d) number of quanta present if it produces 2 J of energy.

Answer

Wavelength of radiation emitted = 616 nm = 616×10^{-9} m (Given)

(a) Frequency of emission (ν)

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

Where,

c = velocity of radiation

λ = wavelength of radiation

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

$$\nu = \frac{3.0\times10^{8}\ \text{m/s}}{616\times10^{-9}\ \text{m}}$$

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

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

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

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

Distance travelled by this radiation in 30 s

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

= 9.0×10^{9} m

(c) Energy of quantum (E) = hν

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

Energy of quantum (E) = 32.27×10^{-20} J

(d) Energy of one photon (quantum) = 32.27×10^{-20} J

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

Number of quanta in 2 J of energy

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

= 6.19×10^{18}

= 6.2×10^{18}

Question 2.48:

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

Answer

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

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

Where,

λ = wavelength of radiation

h = Planck's constant

c = velocity of radiation

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

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

E = 3.313×10^{-19} J

Energy of one photon = 3.313×10^{-19} J

Number of photons received with 3.15×10^{-18} J energy

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

= 9.5

≈ 10

Question 2.49:

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

Answer

Frequency of radiation (v),

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

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

Energy (E) of source = Nhv

Where,

N = number of photons emitted

h = Planck's constant

v = frequency of radiation

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

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

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

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

Question 2.49:

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

Answer

Frequency of radiation (v),

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

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

Energy (E) of source = Nhv

Where,

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N = number of photons emitted

h = Planck's constant

v = frequency of radiation

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

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

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

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

Question 2.51:

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

Answer

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

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

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

Where,

λ_0 = threshold wavelength

h = Planck's constant

c = velocity of radiation

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

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

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

Hence, the threshold wavelength λ_0 is 653 nm.

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

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

Where,

ν_0 = threshold frequency

h = Planck's constant

Substituting the values in the given expression of ν_0:

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

(1 eV = 1.602 × 10^{-19} J)

ν_0 = 4.593 × 10^{14} s^{-1}

Hence, the threshold frequency of radiation (ν_0) is 4.593 × 10^{14} s^{-1}.

(c) According to the question:

Wavelength used in irradiation (λ) = 500 nm

Kinetic energy = h (ν – ν_0)

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

$$= \left(6.626\times10^{-34}\ Js\right)\left(3.0\times10^{8}\ ms^{-1}\right)\left(\frac{\lambda_0-\lambda}{\lambda\lambda_0}\right)$$

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

$$= \frac{\left(1.9878\times10^{-26}\right)\left(153\times10^{9}\right)}{(653)(500)}J$$

= 9.3149 × 10^{-20} J

Kinetic energy of the ejected photoelectron = 9.3149 × 10^{-20}J

Since K.E $= \frac{1}{2}mv^2 = 9.3149\times10^{-20}\ J$

$$v = \sqrt{\frac{2\left(9.3149\times10^{-20}\ J\right)}{9.10939\times10^{-31}\ kg}}$$

$$= \sqrt{2.0451\times10^{11}\ m^2s^{-2}}$$

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

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

Question 2.52:

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

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

Answer

(a) Assuming the threshold wavelength to be λ_0 nm $\left(=\lambda_0 \times 10^{-9}\,\text{m}\right)$, the kinetic energy of the radiation is given as:

$$\text{h}\left(\nu-\nu_0\right)=\frac{1}{2}mv^2$$

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

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

$$\text{h}c\left(\frac{1}{500\times10^{9}}-\frac{1}{\lambda_0\times10^{-9}\,\text{m}}\right)=\frac{1}{2}m\left(2.55\times10^{+5}\times10^{-2}\,\text{ms}^{-1}\right)$$

$$\frac{\text{h}c}{10^{-9}\ \text{m}}\left[\frac{1}{500}-\frac{1}{\lambda_0}\right]=\frac{1}{2}m\left(2.55\times10^{+3}\ \text{ms}^{-1}\right)^2 \qquad (1)$$

Similarly,

$$\frac{\text{h}c}{10^{-9}\text{m}}\left[\frac{1}{450}-\frac{1}{\lambda_0}\right]=\frac{1}{2}m\left(3.45\times10^{+3}\ \text{ms}^{-1}\right)^2 \qquad (2)$$

$$\frac{\text{h}c}{10^{-9}\text{m}}\left[\frac{1}{400}-\frac{1}{\lambda_0}\right]=\frac{1}{2}m\left(5.35\times10^{+3}\ \text{ms}^{-1}\right)^2 \qquad (3)$$

Dividing equation (3) by equation (1):

$$\frac{\left[\dfrac{\lambda_0 - 400}{400\lambda_0}\right]}{\left[\dfrac{\lambda_0 - 500}{500\lambda_0}\right]} = \frac{\left(5.35\times10^{+3}\ \text{ms}^{-1}\right)^2}{\left(2.55\times10^{+3}\ \text{ms}^{-1}\right)^2}$$

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

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

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

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

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

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

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

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

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

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

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

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

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

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

Question 2.53:

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

Answer

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

$E = W_0 + K.E$

$\Rightarrow W_0 = E - K.E$

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

Where,

c = velocity of radiation

h = Planck's constant

λ = wavelength of radiation

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

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

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

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

E = 4.83 eV

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

K.E = 0.35 V

K.E = 0.35 eV

∴ Work function, $W_0 = E - K.E$

= 4.83 eV – 0.35 eV

= 4.48 eV

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

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

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

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

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

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

Question 2.54:

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 1.5×10^7 ms^{-1}, calculate the energy with which it is bound to the nucleus.

Answer

Energy of incident photon (*E*) is given by,

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

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

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

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

Energy of the electron ejected (K.E)

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

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

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

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

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

= *E* – K.E

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

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

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

$$\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} = \left(\frac{5.35}{2.55}\right)^2 = \frac{28.6225}{6.5025}$$
$$\frac{5\lambda_0 - 2000}{4\lambda_0 - 2000} = 4.40177$$
$$17.6070\lambda_0 - 5\lambda_0 = 8803.537 - 2000$$
$$\lambda_0 = \frac{6805.537}{12.607}$$
$$\lambda_0 = 539.8 \text{ nm}$$
$$\lambda_0 \simeq 540 \text{ nm}$$

Question 2.55:

Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as $v = 3.29 \times 10^{15}$ (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.

Answer

Wavelength of transition = 1285 nm

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

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

Since $v = \dfrac{c}{\lambda}$

$$= \frac{3.0 \times 10^{8} \text{ ms}^{-1}}{1285 \times 10^{-9} \text{ m}}$$

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

Substituting the value of v in the given expression,

$$3.29\times10^{15}\left(\frac{1}{9}-\frac{1}{n^2}\right)=2.33\times10^{14}$$

$$\frac{1}{9}-\frac{1}{n^2}=\frac{2.33\times10^{14}}{3.29\times10^{15}}$$

$$\frac{1}{9}-0.7082\times10^{-1}=\frac{1}{n^2}$$

$$\Rightarrow\frac{1}{n^2}=1.1\times10^{-1}-0.7082\times10^{-1}$$

$$\frac{1}{n^2}=4.029\times10^{-2}$$

$$n=\sqrt{\frac{1}{4.029\times10^{-2}}}$$

n = 4.98

$n \approx 5$

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

The spectrum lies in the infra-red region.

Question 2.56:

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

Answer

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

$$r=\frac{0.529\,n^2}{Z}\,\text{Å}$$

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

For radius (r_1) = 1.3225 nm

= 1.32225×10^{-9} m

= 1322.25×10^{-12} m

= 1322.25 pm

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

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

Similarly,

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

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

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

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

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

⇒ n_1 = 5 and n_2 = 2

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

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

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

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

= 2.303 × 10^6 m^{-1}

∴ Wavelength (λ) associated with the emission transition is given by,

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

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

= 0.434 ×10^{-6} m

λ = 434 nm

Question 2.57:

Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is 1.6×10^6 ms^{-1}, calculate de Broglie wavelength associated with this electron.

Answer

From de Broglie's equation,

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

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

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

λ = 455 pm

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

Question 2.58:

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

Answer

From de Broglie's equation,

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

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

Where,

v = velocity of particle (neutron)

h = Planck's constant

m = mass of particle (neutron)

λ = wavelength

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

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

$= 4.94 \times 10^2$ ms^{-1}

v = 494 ms^{-1}

$\therefore$ Velocity associated with the neutron = 494 ms^{-1}

Question 2.59:

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

Answer

According to de Broglie's equation,

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

Where,

λ = wavelength associated with the electron

h = Planck's constant

m = mass of electron

v = velocity of electron

Substituting the values in the expression of *λ:*

$$\lambda = \frac{6.626 \times 10^{-34}\ \text{Js}}{\left(9.10939 \times 10^{-31}\ \text{kg}\right)\left(2.19 \times 10^{6}\ \text{ms}^{-1}\right)}$$

$$= 3.32 \times 10^{-10}\,\text{m} = 3.32 \times 10^{-10}\ \text{m} \times \frac{100}{100}$$

$$= 332 \times 10^{-12}\ \text{m}$$

λ = 332 pm

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

Question 2.60:

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

Answer

According to de Broglie's expression,

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

Substituting the values in the expression,

$$\lambda = \frac{6.626\times10^{-34}\ \text{Js}}{(0.1\ \text{kg})\left(4.37\times10^{5}\ \text{ms}^{-1}\right)}$$

$$\lambda = 1.516\times10^{-38}\ \text{m}$$

Question 2.61:

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.

Answer

From Heisenberg's uncertainty principle,

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

Where,

Δx = uncertainty in position of the electron

Δp = uncertainty in momentum of the electron

Substituting the values in the expression of Δp:

$$\Delta p = \frac{1}{0.002\,\text{nm}}\times\frac{6.626\times10^{-34}\ \text{Js}}{4\times(3.14)}$$

$$= \frac{1}{2\times10^{-12}\ \text{m}}\times\frac{6.626\times10^{-34}\ \text{Js}}{4\times3.14}$$

= 2.637×10^{-23} Jsm^{-1}

Δp = 2.637×10^{-23} kgms^{-1} (1 J = 1 kgms^{2}s^{-1})

∴ Uncertainty in the momentum of the electron = 2.637×10^{-23} kgms^{-1}.

Actual momentum $= \dfrac{h}{4\pi_m \times 0.05\ \text{nm}}$

$$= \frac{6.626\times10^{-34}\ \text{Js}}{4\times3.14\times5.0\times10^{-11}\ \text{m}}$$

= 1.055×10^{-24} kgms^{-1}

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

Question 2.62:

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

1. n = 4, l = 2, m_l = −2 , m_s = −1/2
2. n = 3, l = 2, m_l= 1 , m_s = +1/2
3. n = 4, l = 1, m_l = 0 , m_s = +1/2
4. n = 3, l = 2, m_l = −2 , m_s = −1/2
5. n = 3, l = 1, m_l = −1 , m_s= +1/2
6. n = 4, l = 1, m_l = 0 , m_s = +1/2

Answer

For n = 4 and l = 2, the orbital occupied is 4*d.*

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

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

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

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

Question 2.63:

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

Answer

Nuclear charge experienced by an electron (present in a multi-electron atom) is dependant upon the distance between the nucleus and the orbital, in which the electron is present. As the distance increases, the effective nuclear charge also decreases.

Among *p*-orbitals, 4*p* orbitals are farthest from the nucleus of bromine atom with (+35) charge. Hence, the electrons in the 4*p* orbital will experience the lowest effective nuclear charge. These electrons are shielded by electrons present in the 2*p* and 3*p* orbitals along with the *s*-orbitals. Therefore, they will experience the lowest nuclear charge.

Question 2.64:

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

Answer

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

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

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

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

Question 2.65:

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

Answer

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

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

Question 2.66:

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

Answer

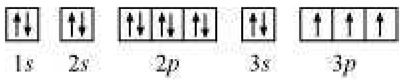
(a) Phosphorus (P):

Atomic number = 15

The electronic configuration of P is:

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

The orbital picture of P can be represented as:



From the orbital picture, phosphorus has **three** unpaired electrons.

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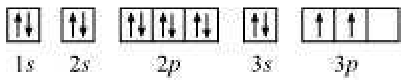
(b) Silicon (Si):

Atomic number = 14

The electronic configuration of Si is:

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

The orbital picture of Si can be represented as:



From the orbital picture, silicon has **two** unpaired electrons.

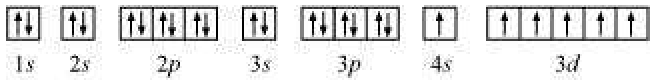
(c) Chromium (Cr):

Atomic number = 24

The electronic configuration of Cr is:

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

The orbital picture of chromium is:



From the orbital picture, chromium has **six** unpaired electrons.

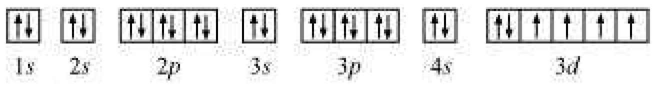
(d) Iron (Fe):

Atomic number = 26

The electronic configuration is:

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

The orbital picture of chromium is:



From the orbital picture, iron has **four** unpaired electrons.

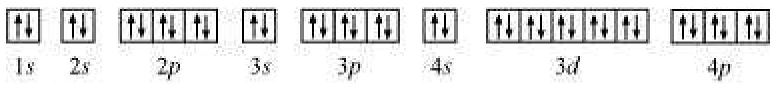
(e) Krypton (Kr):

Atomic number = 36

The electronic configuration is:

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

The orbital picture of krypton is:



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

Question 2.67:

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

Answer

(a) n = 4 (Given)

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

∴ l = 0, 1, 2, 3

Thus, four sub-shells are associated with n = 4, which are *s*, *p*, *d* and *f*.

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

For n = 4

Number of orbitals = 16

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

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

UNIT-2

STRUCTURE OF ATOM

- Atom is the smallest indivisible particle of the matter. Atom is made of electron, proton and neutrons.

	ELECTRON	PROTON	NEUTRON
Discovery	Sir. J. J. Thomson (1869)	Goldstein (1886)	Chadwick (1932)
Nature of charge	Negative	Positive	Neutral
Amount of charge	1.6×10^{-19} Coloumb	1.6×10^{-19} Coloumb	–
Mass	9.11×10^{-31} kg	1.672614×10^{-27} kg	1.67492×10^{-27} kg

- Nucleus was discovered by Rutherford in 1911.
- Atomic number (Z) : the number of protons present in the nucleus (Moseley 1913).
- **Mass Number (A) :** Sum of the number of protons and neutrons present in the nucleus.
- Wavelength, frequency and wave velocity are related to each other by

$c = v\lambda$ where c = velocity in light = 3.0×10^{-8} m/s

v = frequency of s^{-1} or Hz

λ = wavelength in metres

- Wave number $\left(\bar{v}\right)$ is the reciprocal of wavelength $\left(\bar{v} = \frac{1}{\lambda}\right)$.
- According to Planck's quantum theory, the energy is emitted or absorbed not continuously but discontinuously in the form of energy packets called quanta. A quantum of light is called photon. The energy of a quantum is E = hv, where h = Planck's constant, v = frequency of radiation.
- The line spectrum of hydrogen consists of Lyman Series (in UV region), Balmer series (visible region), Paschen, Brackett and Pfund series (IR region).

The wave number of lines can be calculated by the following relation :

$$\bar{v} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

R = Rydberg's constant = 109677 cm^{-1}

For, Lyman series : $n_1 = 1$, $n_2 = 2, 3, 4,$ ....

Balmer series : $n_1 = 2$, $n_2 = 3, 4, 5,$ ....

Paschen series : $n_1 = 3$ and $n_2 = 4, 5, 6$ ....

Brackett series : $n_1 = 4$ and $n_2 = 5, 6, 7,$ ....

Pfund series : $n_1 = 5$ and $n_2 = 6, 7, 8,$ ....

- The energy of electron in hydrogen atom is given by :

$$E_n = -\frac{2\pi^2 mZ^2e^4}{n^2h^2}$$

M = mass of electron, e = charge on electron, Z = atomic number of element

- For hydrogen atom, energy of electron in n^{th} orbit is :

$$E_n = \frac{-1.312\times10^6 Z^2}{n^2}\,\text{Jmol}^{-1} = \frac{-2.178\times10^{-18} Z^2}{n^2}\ \text{Jatom}^{-1} = \frac{-13.60Z^2 eV}{n^2}\ \text{atom}^{-1}$$

where Z = atomic number of H or hydrogen like ions.

- The lowest energy state of an electron in atom is called ground state ($n = 1$), when an electron absorb energy, it jumps to higher energy level called excited state, (first excited state $n = 2$ for H).
- The energy absorbed or emitted during electronic transition is given by the difference of the energies of two levels, *i.e.,*

$$E_2 - E_1 = -2.18 \times 10^{-18}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) \text{ J/atom such that } n_2 > n_1.$$

- The radius of the n^{th} orbit is given by $r_n = \frac{(0.529\text{Å})n^2}{z}$.
- **Photoelectric effect :** When radiation with certain minimum frequency (v_0), called threshold frequency, strike the surface of a metal, electrons (called photoelectrons) are ejected from the surface. With this frequency, the kinetic energy of the photoelectrons ejected is zero. However, if the incident radiation having frequency $v > v_0$, the difference of energy ($hv - hv_0$) is converted into kinetic energy of the photoelectrons *i.e.,* $\frac{1}{2}mv^2 = hv - hv_0$.

  The minimum energy hv_0 required for emission of photoelectrons is called **threshold energy** or **work function.** No photoelectric effect is shown if incident frequency is less than v_0 even if intensity of a radiation is increased. However, number of photoelectrons ejected is proportional to the intensity of incident radiation.

- According to **de Broglie concept,** all material particles (microscopic as well as macroscopic) possess wave character as well as particle character. The wave associated with a material particle is called de Broglie wave or matter wave.

  The relationship between the wavelength (λ) of the wave and the mass (m)

of the material particle moving with a velocity ν is called **de Broglie equation**. It is given by

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

where h is Planck's constant and p is momentum of the particle.
The wave nature of electron has been confirmed by **Davisson and Germer's experiment** whereas the particle nature is confirmed by **scintillation method** as well as by the **photoelectric effect**.

- **Heisenberg's uncertainty Principle** states that "It is impossible to measure simultaneously the position and momentum of a microscopic particle with absolute accuracy. If one of them is measured with greater accuracy, the other becomes less accurate. The product of their uncertainties is always equal to or greater than $\frac{h}{4\pi}$."

  Mathematically $\Delta x \times \Delta p \geq \frac{h}{4\pi}$

  where Δx = uncertainty in position,
  Δp = uncertainty in momentum

  de Broglie concept as well as uncertainty principle have no significance in everyday life because they have significance only for microscopic particles but we come across macroscopic bodies in everyday life.

- **Quantum numbers**

  The four quantum numbers provide the following informations :

  (1) Principal quantum number (n)

  $n = 1, 2, 3, 4, ...\infty$

  It identifies shell, determines sizes and energy of orbitals and number of orbitals in the n^{th} shell which is equal to n^2.

  (2) Azimuthal quantum number (l)

  For a given value of n, it can have n values ranging from 0 to $n - 1$. It identifies subshell, determines the shape of orbitals, energy of orbitals in multi-electron atoms along with principal quantum number and orbital angular momentum, *i.e.,* $\sqrt{l(l+1)}\frac{h}{2\pi}$. The number of orbitals in a subshell = $2l + 1$.

Subshell notation	s	p	d	f	g
Value of 'l'	0	1	2	3	4
Number of orbitals	1	3	5	7	9

(3) **Magnetic orbital quantum number (m_l)**
For a given value of 'l', m_l has a total of $(2l + 1)$ values ranging from $-l$ to $+l$ including '0'. It determines the orientation of orbital.

(4) **Magnetic spin quantum number (m_s)**
It can take the values of +½ or −½ and determines the orientation of spin.

- **Pauli's Exclusion Principle :** "No two electrons in an atom can have the same set of four quantum numbers." Two electrons can have same values for *n, l* and m_l provided their spins are opposite (m_s is different). Therefore an orbital can have at the most two electrons if they have opposite spins.
- **Hund's Rule of maximum Multiplicity :** "The electrons start pairing only when all the degenerate orbitals of a subshell are singly occupied with parallel spins." *e.g.,* N : $1s^2, 2s^2, 2p_x^{\,1}, 2p_y^{\,1}, 2p_z^{\,1}$.
- **Aufbau Principle :** "Orbitals are filled up in increasing order of their energy with the help of pauli principle and Hund's rule."
1. Orbitals are filled up in the increasing order of their $(n + 1)$ values.
2. If two orbitals have same $(n + 1)$ values, then the one which has lower n value, will be filled up first.
Increasing order of energy :
$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$
Exception of Aufbau principle : Extra stability is associated with the exactly half-filled and fully-filled orbitals. Thus the $p^3, p^6, d^5, d^{10}, f^7, f^{14}$ etc. have extra stability, *i.e.,* lower energy and therefore, more stable.

1 - MARK QUESTIONS

1. Indicate the number of electrons, protons and neutrons in the element $^{238}_{92}U$.
[**Ans.** $e = 92, p = 92, n = 146$]

2. Name the experiment used in determination of the charge of an electron.

3. Arrange the electron (e), protons (p) and alpha particle (α) in the increasing order for the values of e/m (charge/mass). [**Ans.** $\alpha < p < e$]

4. Calculate the mass of one mole of electron. [Given : $m_e = 9.11 \times 10^{-31}$ kg]
[**Ans.** 0.55 mg]

5. Write the dimensions of Planck's constant. Mention some other physical quantity, which has the same dimension.

6. Name the element which was discovered in the sun by spectroscopic method.
[**Ans.** Helium (He)]

7. Which of the following will not show deflection from the path on passing through an electric field, Proton, Cathode rays, Anode rays, Electron, Neutron.

[**Hint :** Neutron (n) will not show deflection since it is electrically neutral]

8. Out of electron and proton, which one will have a higher velocity to produce matter waves of the same wavelength ?

[**Hint :** $\lambda = \frac{h}{mv}$, For same wavelength, m_e is very small as compared to m_p. Therefore v_e will be higher as compared to v_p.]

9. An anion A^{3-} has 18 electrons. Write the atomic number of A. [**Ans.** 15]

10. What is the value of orbital angular momentum of 6*s* orbital ?

11. What physical meaning is attributed to the square of the absolute value of wave function $|\Psi^2|$?

12. Name two physical quantities which can be estimated by principal quantum number (*n*).

13. Which shell would be the first to have *g*-subshell ?

14. How many electrons in an atom can have $n + l = 6$? [**Ans.** 18]

15. Name three quantum numbers which arise as a result of the solution of Schrodinger wave equation.

16. Write electronic configuration of the Cr^{3+} ion. [Atomic number of Cr = 24]

17. The ion of an element has configuration $[Ar]3d^4$ in +3 oxidation state. Write the electronic configuration of its atom.

18. State Pauli's exclusion principle.

19. How many nodes are there in 3*s* orbital ? [**Ans.** 2 nodes]

20. Why 1*p*, 2*d* and 3*f* subshells are not possible ?

21. How many unpaired electrons are present in Fe^{3+} ion ?

22. State Hunds rule of maximum multiplicity.

23. Using *s, p, d* notations, describe the orbital with the following quantum numbers :

(a) $n = 4, l = 2$ (b) $n = 1, l = 0$ [**Ans.** (a) 4*d* (b) 1*s*]

24. Which quantum number determines the orientation of atomic orbital ?

25. Which orbital is non-directional ?

26. Write the correct set of four quantum numbers for the valence electron (outermost electron) of potassium (Z = 19).

27. Which principle is not obeyed in writing of electronic configuration :

2*s*	2*p*		
$\uparrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$

2- MARK QUESTIONS

1. Give examples of each of the following :

(a) Isotope of ${}^{35}_{17}Cl$ (b) Isobar of ${}^{40}_{18}Ar$

(c) Isotone of ${}^{15}_{7}N$ (d) Isoelectronic species of S^{2-}

[**Ans.** (a) ${}^{37}_{17}Cl$ (b) ${}^{40}_{20}Ca$ (c) ${}^{16}_{8}O$ (d) K^+, Ca^{2+}, Ar, Cl^-, S^{2-}]

2. Describe the cathode ray experiment. How will you detect the spot where the rays strike ?

3. Outline Rutherford's contribution to understand the nucleus of an atom.

4. Calculate the percentage of higher isotope of neon which has average atomic mass 20.2 and the isotopes have the mass numbers 20 and 22.

[**Ans.** 10% ${}^{22}_{10}Ne$]

5. Account for the following :

(a) Cathode rays are produced only when the pressure of the gas inside the discharge tube is very very low.

(b) Can a thin foil of aluminium be used in place of gold (Au) in Ruther ford experiment ? Give suitable explanation.

[**Hint :** Lighter nuclei cannot exhibit proper deflection of α-particles.]

6. Distinguish between an atomic emission spectrum and an atomic absorption spectrum.

7. The energies of electrons are said to be quantized. Explain.

8. A laser used to read compact disc (CD) emits red light of wavelength 700 nm. How many photons does it emit each second if its power is 1 W ?

[**Ans.** 3.5×10^{18} s^{-1}]

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

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

11. Explain, giving reasons, which of the following sets of quantum numbers are not possible :

(a) $n = 0$ $l = 0$ $m_l = 0$ $m_s = + ½$

(b) $n = 1$ $l = 0$ $m_l = 0$ $m_s = - ½$

(c) $n = 1$ $l = 1$ $m_l = 0$ $m_s = + ½$

(d) $n = 3$ $l = 1$ $m_l = 0$ $m_s = 0$

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

The ionisation energy for the H atom in the ground state is 2.18×10^{-18} J atom^{-1}.

[**Ans.** 8.72×10^{-18} J]

13. (a) An atomic orbital has $n = 3$. Write the possible values of l and m_l ?

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

14. Draw the boundary surface diagrams of d_{xy} and $d_{x^2-y^2}$ orbitals.

15. What is meant by degenerate orbitals ? Illustrate with the help of one example.

16. How does a 1s orbital differ from a 2s orbital ? Mention two points in support of your answer.

17. Calculate the wave number for the shortest wavelength transition in the Balmer series of atomic hydrogen. [**Ans.** 27419.25 cm^{-1}]

18. Calculate (a) wave number and (b) frequency of yellow radiation having wavelength 5800Å.

$$\begin{pmatrix} \bar{\nu} = 1.724 \times 10^{4} \text{ cm}^{-1} \\ \nu = 5.172 \times 10^{14} 5 \text{ s}^{-1} \end{pmatrix}$$

19. Calculate the energy associated with the first orbit of He^{+}. What is the radius of this orbit. [**Ans. :** $E_1 - 8.72 \times 10^{-18}$ J, r = 0.2645 nm]

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

[**Ans.** (3.3×10^{6}]

21. 2×10^{8} atoms of carbon at arranged side by side. Calculate the radius of carbon atom if the length of this arrangements is 2.4 cm. [**Ans.** 0.06 nm]

22. In milikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the statis electric charge on the oil drop is -1.282×10^{-18} C, calculate the number of electrons present on it [**Ans.** 8]

23. An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutions than electron. Assign the symbol of this ion

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

3 - MARK QUESTIONS

1. Differentiate between :
(a) Photon and quantum
(b) Orbit and orbital
(c) de Broglie waves and electromagnetic waves

2. (a) State Heisenberg uncertainty principle.
(b) "Electron cannot exist within the atomic nucleus." Justify the statement on the basis of Heisenberg uncertainty principle.
[**Hint :** Radius of nucleus = 10^{-15}m]

3. Calculate the wavelength of an electron that has been accelerated in aparticle accelerator through a potential difference of 1keV. [1 eV = 1.6×10^{-19} J]
[**Ans.** 3.87×10^{-7} m]

4. The kinetic energy of a subatomic particle is 5.86×10^{-25} J. Calculate the frequency of the particle wave. [**Ans.** $1.76 \times 109\ s^{-1}$]

5. Calculate the energy required to excite the electron in the atom from $n = 1$ to $n = 2$. The ionization enthalpy of hydrogen atom is 1.312×10^6 J mol^{-1}.

6. Calculate and compare the products of uncertainty in position and uncertainty in velocity for a milligram-sized object and that of an electron ($m_e = 9.11 \times 10^{-31}$ kg). What conclusion would you draw from result ?

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

8. How many series are found in the spectrum of atomic hydrogen ? Mention their names and the regions in which they appear.

9. (a) Write the electronic configurations of the followings :
(i) H^- (ii) Na+ (iii) O^{2-}
(b) Symbols $^{79}_{35}Br$ and ^{79}Br can be written, whereas symbols $^{79}_{35}Br$ and ^{35}Br are not acceptable. Explain why ?
(c) An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons. Assign the symbol of this ion.

5 - MARK QUESTIONS

1. (a) Define photoelectric effect ? Mention its one practical application in daily life.

(b) 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.** 4.41×10^{14} s^{-1}, 2.91×10^{-19} J]

2. (a) State de Broglie relation. Why it is not meaningful for a moving cricket ball ?

(b) Which out of (i) electron (*e*), (ii) proton (*p*) (iii) α particle, has maximum de Broglie's wavelength ? [Assume that all are moving with the same velocity.]

(c) The wavelength associated with particle 'A' is 5×10^{-8} m. Calculate the wavelength associated with particle B, if its momentum is half of A.

[**Ans.** 1×10^{-7} m]

3. (a) What is the significance of the statement "Product of uncertainty in position and momentum is always constant." ?

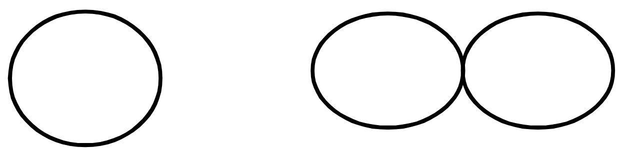
(b) Why is uncertainty principle not applicable to macroscopic and semimicro particles ?

(c) An electron has a speed of 40 ms^{-1} accurate upto 99.99%. What is the uncertainty in locating its position ? ($m_e = 9.11 \times 10^{-31}$ kg)

[**Ans.** 1.45×10^{-2} m]

4. (a) State Aufbau principle.

(b) What is the physical significance of lines in the following depiction of atomic orbital ?



s-orbital p-orbital

(c) Explain the following with suitable reason :

(i) In potassium, the 19th electron enters in 4*s* subshell instead of 3*d* subshell.

(ii) Chromium has configuration $3d^5\ 4s^1$ and not $3d^4\ 4s^2$.

(iii) The three electrons present in 2*p* subshell of nitrogen (N) remain unpaired and have the parallel spins.

5. (a) The work function for caesium atom is 1.9 ev. Calculate (i) the threshold wave length and (ii) the threshold frequency of the radiation. If the caesium element is irradiated with a radiation of wave length 500 nm, calculatethe kinetic energy and the velocity of the ejected photo electron

(i) 652 nm

(ii) $4.598 \times 10^{14}\ S^{-1}$

(iii) 3.97×10^{-13} J, 9.33×10^{8} m S^{-1}

(b) Among the following pairs of orbitals, which orbital will experience the larger effective nuclear charge

(i) 2s and 3s

(ii) 3d and 3p