NCERT Solutions for Class 11 Chemistry Chapter 2 Structure of Atom

Question 2.1 (i) Calculate the number of electrons which will together weigh one gram.

Answer :

As the mass of one electron we know is $9.11 \times 10^{-31} kg$.

Therefore,

$$1g = 10^{-3} kg = \left(\frac{1}{9.11 \times 10^{-31}}\right) \times 10^{-3} \ electrons$$

$$\Rightarrow 1.098 \times 10^{27} \ \ electrons.$$.

Question 2.1 (ii) Calculate the mass and charge of one mole of electrons.

Answer :

As the mass of one electron is equal to $9.11 \times 10^{-31} kg$

Therefore, Mass of 1 mole or 6.022×10^{23} electrons $= (9.11 \times 10^{-31}) \times (6.022 \times 10^{23}) = 5.48 \times 10^{-7} kg$.

Charge on one electron is $1.602 \times 10^{-19} \ coulomb$.

Therefore, the charge on 1 mole of electrons will be:

$$\Rightarrow (1.602 \times 10^{-19}) \times (6.022 \times 10^{23}) = 9.65 \times 10^{4} \ coulombs$$.

Question 2.2 (i) Calculate the total number of electrons present in one mole of methane.

Answer :

1 molecule of methane CH_4 contains $6+4=10$ electrons.

Therefore, 1 mole of methane will contain:

$6.022\times10^{23}\times10=6.022\times10^{24}$ **electrons.**

Question 2.2 (ii) Find (a) the total number of neutrons in 7 mg of ^{14}C. (Assume that mass of a neutron $=1.675\times10^{-27}\ kg$).

Answer :

As 1 atom of ^{14}C contains $=14-6=8\ neutrons$.

and the number of atoms in ^{14}C in 1 mole is 6.022×10^{23} atoms.

Therefore, the number of neutrons in 14g of ^{14}C in 1 mole $=8\times6.022\times10^{23}\ neutrons$.

The number of neutrons in $7mg$:

$$=\left(\frac{8\times6.022\times10^{23}\times7}{14000}\right)$$

$$=2.4088\times10^{21}\ neutrons$$.

Question 2.2 (ii) Find (b) the total mass of neutrons in 7 mg of ^{14}C. (Assume that mass of a neutron = 1.675 × 10 kg).

Answer :

As the mass of one neutron is $1.674 \times 10^{-27} kg$.

Then the mass of total neutrons in 7grams of ^{14}C:

$= (2.4088 \times 10^{21})(1.675 \times 10^{-27} kg) = 4.035 \times 10^{-6} kg$.

Question 2.2 (iii) Find (a) the total number of protons in 34 mg of NH_3 at STP.
Will the answer change if the temperature and pressure are changed ?

Answer :

1 mole of ammonia $NH_3 = 17g\ NH_3 = 6.022 \times 10^{23}\ molecules\ of\ NH_3$.

and 1 atom of NH_3 contains $= 7 + 3 = 10\ protons$.

Therefore, the number of protons in 1 mole of $NH_3 = 6.022 \times 10^{24}\ protons$.

Number of protons in 3mg of NH_3 : $= \frac{(6.022 \times 10^{24} \times 34)}{17} \times 1000 = 1.2044 \times 10^{22}\ protons$

No, there will be no effect of temperature and pressure.

Question 2.2 (iii) Find (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 :

As the mass of one proton is $1.6726 \times 10^{-27} kg$

Therefore, the mass of $1.2044 \times 10^{22}\ protons$ will be:

$= (1.6726 \times 10^{-27}) \times (1.2044 \times 10^{22}) kg = 2.0145 \times 10^{-5} kg.$

No, there will be no effect of temperature and pressure.

Question 2.3(i) How many neutrons and protons are there in the following nuclei?

$^{13}_{6}\mathrm{C}$

Answer :

Given the nucleus of carbon: $^{13}_{6}\mathrm{C}$

Atomic number **(Z)** = 6

Mass number **(A)** = 13

Number of protons **(Z)** = 6

Number of neutrons **(A-Z)** = 13-6 = 7

Question 2.3(ii) How many neutrons and protons are there in the following nuclei?

$^{16}_{8}O$

Answer :

Given the nucleus of oxygen: $^{16}_{8}O$

Atomic number **(Z)** = 8

Mass number **(A)** = 16

Number of protons **(Z)** = 8

Number of neutrons **(A-Z)** = 16-8 = 8

2.3How many neutrons and protons are there in the following nuclei?

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Question 2.3(iii) How many neutrons and protons are there in the following nuclei?

$^{24}_{12}Mg$

Answer :

Given the nucleus of Magesium: $^{24}_{12}Mg$

Atomic number **(Z)** = 12

Mass number **(A)** = 24

Number of protons **(Z)** = 12

Number of neutrons **(A-Z)** = 24-12 = 12

Question 2.3(iv) How many neutrons and protons are there in the following nuclei?

$^{56}_{26}\mathrm{Fe}$

Answer :

Given the nucleus of Iron: $^{56}_{26}\mathrm{Fe}$

Atomic number **(Z)** = 26

Mass number **(A)** = 56

Number of protons **(Z)** = 26

Number of neutrons **(A-Z)** = 56-26 = 30

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Question 2.3(v) How many neutrons and protons are there in the following nuclei?

$^{88}_{38}\mathrm{Sr}$

Answer :

Given the nucleus of Strontium: $^{88}_{38}\mathrm{Sr}$

Atomic number **(Z)** = 38

Mass number **(A)** = 88

Number of protons **(Z)** = 38

Number of neutrons **(A-Z)** = 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.

Answer :

For the given atomic number **Z=17** and mass number **A=35;**

Atom is $^{35}Cl_{17}$.

Question 2.4 Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)

(ii) Z = 92 , A = 233.

Answer :

For the given atomic number **Z=92** and mass number **A=233;**

Atom is $^{233}U_{92}$.

Question 2.4 Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)

(iii) Z = 4 , A = 9.

Answer :

For the given atomic number **Z=4** and mass number **A=9;**

Atom is $^{9}Be_{4}$.

Question 2.5 Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency (v) and wavenumber (ν) of the yellow light.

Answer :

Given the wavelength of the yellow light emitted from a sodium, lamp is $\lambda = 580nm = 580 \times 10^{-9}m$.

And the frequency will be:

$$\nu = \frac{c}{\lambda} = \frac{3.0 \times 10^{8}m/s}{580 \times 10^{-9}m} = 5.17 \times 10^{14}s^{-1}$$

Therefore the wavenumber,

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

Question 2.6 Find energy of each of the photons which

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

Answer :

If a photon has a frequency of $\vartheta = 3 \times 10^{15} Hz$.

Then, the energy of each of the photons will be:

$$E = h\vartheta = (6.626 \times 10^{-34} J.s) \times (3 \times 10^{15} s^{-1})$$

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

Question 2.6 Find energy of each of the photons which

(ii) have wavelength of 0.50 Å.

Answer :

For the wavelength $\lambda = 0.50 \times 10^{-10} m$.

The energy of each of the photons will be:

Question 2.7 Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0×10^{-10} s .

Answer :

Given frequency, wavelength, and the wave number of a light wave:

$$Frequency(\nu) = \frac{1}{Period} = \frac{1}{2.0 \times 10^{-10} s} = 5 \times 10^9 s^{-1}$$

$$Wavelength(\lambda) = \frac{c}{\nu} = \frac{3.0 \times 10^8 ms^{-1}}{5 \times 10^9 s^{-1}} = 6.0 \times 10^2 m$$

$$Wave\ number\ (\vartheta) = \frac{1}{\lambda} = \frac{1}{6.0 \times 10^2} m = 16.66\ m^{-1}$$

Question 2.8 What is the number of photons of light with a wavelength of 4000 pm that provide 1J of energy?

Answer :

Given the wavelength of light $\lambda = 4000\ pm = 4000 \times 10^{-12} m = 4 \times 10^{-9} m$.

and Energy is 1J of energy:

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

Therefore, the number of photons of light with a wavelength of 4000 pm that provides 1J of energy is:

$$= 2.012 \times 10^{16}\ photons.$$

Question 2.9(i) A photon of wavelength $4 \times 10^{-7} m$ strikes on metal surface, the work function of the metal being 2.13 eV. Calculate (i) the energy of the photon (eV)

Answer :

The photon is having a wavelength of $4 \times 10^{-7} m$ strikes on a metal surface, where the work function of the metal being is $2.13\ eV$.

So, Energy of the photon:

$$E = h\nu = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} J.s) \times (3.0 \times 10^{8} ms^{-1})}{(4 \times 10^{-7} m)}$$

$$= 4.97 \times 10^{-19} J$$

$$= \frac{4.97 \times 10^{-19} J}{1.602 \times 10^{-19}} eV$$

$$= 3.10\ eV$$

Question 2.9(ii) A photon of wavelength 4 × 10^-7 m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate (ii) the kinetic energy of the emission

Answer :

The photon is having a wavelength of $4 \times 10^{-7} m$ strikes on a metal surface, where the work function of the metal being is $2.13\ eV$.

The kinetic energy of the emission will be:

$$K.E. = h\nu - h\nu_o = 3.10 - 2.13\ eV = 0.97\ eV$$

Question 2.9 A photon of wavelength 4 × 10 m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate (iii) the velocity of the photoelectron ($1 \text{ eV} = 1.6020 \times 10^{-19} \text{J}$) .

Answer :

The photon is having a wavelength of $4 \times 10^{-7} m$ strikes on a metal surface, where the work function of the metal being is $2.13 \ eV$.

From the previous part, we have the Kinetic Energy (K.E.):

$$K.E. = \frac{1}{2} m v^2 = 0.97 \ eV = 0.97 \times 1.602 \times 10^{-19} J$$

$$\Rightarrow \frac{1}{2} \left(9.11 \times 10^{-31} kg\right) \times v^2 = 0.97 \times 1.602 \times 10^{-19} J$$

$$\because \left(mass \ of \ an \ electron = 9.11 \times 10^{-31} kg\right)$$

$$\Rightarrow v^2 = 0.341 \times 10^{12} = 34.1 \times 10^{10} m/s$$

$$\Rightarrow v = 5.84 \times 10^5 m/s$$

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 :

Given the wavelength of the electromagnetic radiation is $242 \ nm$ which is just sufficient to ionize the sodium atom.

So, the ionization energy required will be:

$$E = Nh\nu = N\frac{hc}{\lambda}$$

$$= 4.945 \times 10^5 Jmol^{-1} = 494.5kJmol^{-1}$$

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 :

Given that the light is monochromatic yellow of wavelength $\lambda = 0.57\mu m$.

Hence the energy emitted by the bulb will be:

$$E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} Js \times 3.0 \times 10^8 m/s}{0.57 \times 10^{-6} m} = 3.48 \times 10^{-19} J$$

Therefore, the number of photons emitted per second:

$$\frac{25Js^{-1}}{3.48 \times 10^{-19} J} = 7.18 \times 10^{19}$$

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

Answer :

Given the wavelength of radiation is $6800\ \mathring{A}$.

$$Energy\ given = Work\ function + Kinetic\ energy.$$

But the electrons are emitted with zero velocity from a metal surface when it is exposed to radiation. That means **the kinetic energy will be zero** .

So, the **Threshold frequency** ν_o will be:

$$\nu_o = \frac{c}{\lambda_o} = \frac{3.0 \times 10^8 m/s}{6800 \times 10^{-10} m} = 4.14 \times 10^{14} s^{-1}$$

and the **Work function** will be:

$$W_o = h\nu_o = 6.626 \times 10^{-34} Js \times 4.14 \times 10^{14} s^{-1} = 2.92 \times 10^{-19} J$$

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 :

When an electron in a hydrogen atom undergoes a transition from an energy level with $n = 4$ to an energy level $n = 2$, there will be an emission of energy whose wavelength can be found by:

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

$$= 109677\left(\frac{1}{2^2} - \frac{1}{4^2}\right) cm^{-1} = 20564.4 cm^{-1}$$

and wavelength will be equal to:

$$\lambda = \frac{1}{\vartheta} = \frac{1}{20564.4} = 486 \times 10^{-7} cm = 486 \times 10^{-9} m = 486 nm$$

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 energy which is required to ionize an H atom if the electron occupies n=5 orbit is:

$$E_n = \frac{-21.8 \times 10^{-19}}{n^2} J \ atom^{-1}$$

For ionization from 5th orbit, $n_1 = 5 \ and \ n_2 = \infty$

Therefore, $= 21.8 \times 10^{-19} \times \left(\frac{1}{5}^2 - \frac{1}{\infty}\right) = 8.72 \times 10^{-20} J$

For ionization from 1st orbit, $n_1 = 1 \ and \ n_2 = \infty$

Therefore,

$$\triangle E' = 21.8 \times 10^{-19} \times \left(\frac{1}{1}^2 - \frac{1}{\infty}\right) = 21.8 \times 10^{-19} J$$

$$\frac{\triangle E'}{\triangle E} = \frac{21.8 \times 10^{-19}}{8.72 \times 10^{-20}} = 25$$

Hence, 25 times less energy is required to ionize an electron in the 5th orbital of the 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 a H atom in n = 6 drops to the ground state?

Answer :

The number of lines produced when an electron from n^{th} shell drops to the ground state:

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

According to the question, the maximum number of emission lines when the excited electron of an H atom in $n = 6$ drops to the ground state will be:

$$\frac{6(6-1)}{2} = 15$$

These are produced because of the following transitions:

$6 \rightarrow 5$ $6 \rightarrow 4$ $6 \rightarrow 3$ $6 \rightarrow 2$ $6 \rightarrow 1$

$5 \rightarrow 4$ $5 \rightarrow 3$ $5 \rightarrow 2$ $5 \rightarrow 1$

$4 \rightarrow 3$ $4 \rightarrow 2$ $4 \rightarrow 1$

$3 \rightarrow 2$ $3 \rightarrow 1$

$2 \rightarrow 1$

Question 2.16 (i) The energy associated with the first orbit in the hydrogen atom is $-2.18 \times 10^{-18}\ \text{J atom}^{-1}$. What is the energy associated with the fifth orbit?

Answer :

The energy associated with the first orbit in the hydrogen atom is $-2.18 \times 10^{-18}\ \text{J atom}^{-1}$ $\qquad (Given)$

The energy of an electron in n^{th} shell is given by:

$$E = \frac{-2.18 \times 10^{-18}}{n^2} J\ atom^{-1}$$

So, the energy associated with the **fifth orbit** would be:

$$E = \frac{-2.18 \times 10^{-18}}{5^2} J\ atom^{-1} = \frac{-2.18 \times 10^{-18}}{25} J\ atom^{-1}$$

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

Question 2.16 (ii) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

Answer :

The radius of Bohr's n^{th} orbit for the hydrogen atom is given by,

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

So, for $n = 5$ we have

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

$$r_5 = 1.3225\ nm$$

Question 2.17 Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

Answer :

Balmer formula:

$$\vartheta = \frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

As we can note from the formula that the wavenumber is inversely proportional to the wavelength.

Hence, for the longest wavelength transition in the Balmer series of atomic hydrogen wavenumber has to be the smallest or n_2 should be minimum i.e., $n_2 = 3$.

For the Balmer series, $n_1 = 2$

Thus, the expression of wavenumber is given by,

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

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

$$\vartheta = (1.097 \times 10^7\ m^{-1}) \left[\frac{9-4}{36}\right]$$

$$\vartheta = (1.097 \times 10^7\ m^{-1}) \left[\frac{5}{36}\right] = 1.5236 \times 10^6\ m^{-1}$$

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 :

The ground state energy:

$$E_1 = -2.18 \times 10^{-11}\ ergs$$

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

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

The energy required to shift the electron from the **first** Bohr orbit to the **fifth** Bohr orbit is:

$$\triangle E = E_5 - E_1$$

And the expression for the energy of an electron is given by:

$$E_n = -\frac{2n^2 me^4 Z^2}{n^2 h^2}$$

where **m** is mass of an electron, **Z** is the atomic mass of an atom, **e** is a charge of an electron, and **h** is the Planck's constant.

Now, substituting the values in the equation, we get

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

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

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

Hence, the wavelength of the emitted light will be:

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} Js)(3 \times 10^{8} m/s)}{(2.0928 \times 10^{-18} J)}$$

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

Question 2.19 The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18})/n^2\ \mathrm{J}$. 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 :

The expression for the energy of an electron in hydrogen is:

$$E_n = -\frac{2n^2 m e^4 Z^2}{n^2 h^2}$$

Where **m** is mass of electrons, **Z** is the atomic mass of an atom, **e** is the charge of an electron, and **h** is the Planck's constant.

and electron energy in the hydrogen atom is given by,

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

The electron energy in $n = 2$ orbit is:

$$E_n = -\frac{(2.18 \times 10^{-18})}{2^2} J = 0.5465 \times 10^{-18} J$$

Therefore, the energy required for the ionization from $n = 2$ is $5.45 \times 10^{-19} J$

Now, the longest wavelength of light that can be used to cause this transition will be:

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

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} Js)(3 \times 10^8)m/s}{5.45 \times 10^{-19} J}$$

$$= 3.674 \times 10^{-7} m = 3.674 \times 10^{-5} cm$$

Question 2.20 Calculate the wavelength of an electron moving with a velocity of $2.05 \times 10^7 \text{ ms}^{-1}$.

Answer :

The wavelength of an electron is given by the **de Broglie's** equation:

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

Where,

λ is the wavelength of moving particle,

m is the mass of the particle, i.e., $9.11 \times 10^{-31} kg$

v is the velocity of the particle,i.e., $2.05 \times 10^7 \text{ ms}^{-1}$ (**Given)**

and **h** is the Planck's constant value, i.e., $(6.626 \times 10^{-34} Js)$

Now, substituting the values in the equation, we get

$$\lambda = \frac{(6.626 \times 10^{-34} Js)}{(9.11 \times 10^{-31} kg)(2.05 \times 10^{7} m/s)} = 3.548 \times 10^{-11} m$$

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

Question 2.21 The mass of an electron is $9.1 \times 10^{-31}\ \mathrm{kg}$ If its K.E. is $3.0 \times 10^{-25}\ \mathrm{J}$, calculate its wavelength.

Answer :

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

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

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

Hence we get,

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

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

Hence the wavelength is given by,

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

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

Question 2.22 Which of the following are isoelectronic species i.e., those having the same number of electrons?

$\mathrm{Na^+},\ \mathrm{K^+},\ \mathrm{Mg^{2+}},\ \mathrm{Ca^{2+}},\ \mathrm{S^{2-}},\ \mathrm{Ar}$

Answer :

Calculating the number of electrons for each species.

Na has 11 electrons then, Na^+ will have $(11 - 1) = 10$ electrons.

K has 19 electrons then, K^+ will have $(19 - 1) = 18$ electrons.

Mg has 12 electrons then, Mg^{2+} will have $(12 - 2) = 10$ electrons.

Ca has 20 electrons then, Ca^{2+} will have $(20 - 2) = 18$ electrons.

S has 16 electrons then, S^{2-} will have $(16 + 2) = 18$ electrons.

Ar has 18 electrons.

Hence, the following are isoelectronic species:

$Na^+\ and\ Mg^{2+}$ having 10 electrons each.

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

Question : **2.23** (i) Write the electronic configurations of the following ion:

(a) H^{-}

Answer :

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

Now, the electronic configuration of H^{-} will be $1s^2$.

Question 2.23 (i) Write the electronic configurations of the following ion :

(b) Na^{+}

Answer :

The electronic configuration of Na having $Z = 11$ is : $1s^2 2s^2 2p^6 3s^1$

Now, the electronic configuration of Na^{+} will be $1s^2 2s^2 2p^6$.

Question 2.23 (i) Write the electronic configurations of the following ion :

(c) O^{2-}

Answer :

The electronic configuration of O having $Z = 8$ is : $1s^2 2s^2 2p^4$

Now, the electronic configuration of O^{2-} will be $1s^2 2s^2 2p^6$.

Question 2.23 (ii) What are the atomic numbers of elements whose outermost electrons are represented by

(a) $3\mathrm{s}^1$

Answer :

With given outermost electrons $3\mathrm{s}^1$,

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

Hence the number of electrons present in the atom of the element is:

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

Therefore, the atomic number of the element is 11 which is **Sodium (Na)** .

Question 2.23 (ii) What are the atomic numbers of elements whose outermost electrons are represented by

(b) $2p^3$

Answer :

With given outermost electrons $2p^3$,

The complete electronic configuration is $1s^2 2s^2 2p^3$

Hence the number of electrons present in the atom of the element is:

$= 2 + 2 + 3 = 7$.

Therefore, the atomic number of the element is 7 which is **NItrogen (N)** .

Question 2.23 (ii) What are the atomic numbers of elements whose outermost electrons are represented by

(c) $3p^5$

Answer :

With given outermost electrons $3p^5$,

The complete electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$

Hence the number of electrons present in the atom of the element is:

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

Therefore, the atomic number of the element is 17 which is **Chlorine (Cl)** .

Question 2.23 (iii) Which atoms are indicated by the following configurations?

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

Answer :

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

Therefore, the atomic number of the element is **3** which is **Lithium** , a p-block element.

Question 2.23 (iii) Which atoms are indicated by the following configurations?

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

Answer :

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

Therefore, the atomic number of the element is **15** which is **Phosphorus** , a p-block element.

Question 2.23 (iii) Which atoms are indicated by the following configurations?

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

Answer :

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

Therefore, the atomic number of the element is **21** which is **Scandium** , a d-block element.

Question 2.24 What is the lowest value of *n* that allows g orbitals to exist?

Answer :

For g-orbital, the value of Azimuthal quantum number (l) will be 4.

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

Therefore, for l =4, the minimum value of n should be 5.

Question 2.25 An electron is in one of the 3 *d* orbitals. Give the possible values of *n* , *l* and m_l for this electron.

Answer :

For d-orbital, the value of Azimuthal quantum number (l) = 2.

When l =2, the values of m are: $-2, -1, 0, +1, +2$

Now, for the 3d orbital:

The value of Principal quantum number, $n = 3$

Azimuthal quantum number, $l = 2$

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

Question 2.26(i) An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons

Answer :

Given an atom of an element contains 29 electrons and 35 neutrons.

Now, for an atom to be neutral, the number of protons is equal to the number of electrons.

Therefore, the **number of protons** in the atom of the given element will be **29** .

Question 2.26(ii) An atom of an element contains 29 electrons and 35 neutrons. Deduce (ii) the electronic configuration of the element.

Answer :

Given an atom of an element contains 29 electrons and 35 neutrons.

The electronic configuration of the atom will be:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ which is the electronic configuration of **copper** .

Question 2.27 Give the number of electrons in the species H_2^+ , H_2 and O_2^+

Answer :

The number of electrons in H_2 molecule is $(1+1)=2$.

The number of electrons in H_2^+ molecule will be one less than the number of electrons in H_2molecule. i.e, $(2-1)=1$.

The number of electrons present in O_2^+ molecule will be one less than the number of electrons present in O_2 molecule. i.e., $(16-1)=15$

Question 2.28 (i) An atomic orbital has *n* = *3* . What are the possible values of *l* and *m* $_l$?

Answer :

For a given value of the principal quantum number (n) , the azimuthal quantum number (l) can have values from 0 to $(n-1)$.

Therefore, for given atomic orbital $n=3$,

The value of l can take values from 0 to $(3-1)=2$, i.e., $l=0,1,2$.

And for a given value of l , the Magnetic quantum number m_l can have $(2l+1)$ values.

When value of $l=0$ then, $m=0$,

or $l=1$ then, $m=-1,0,+1$

or $l=2$ then, $m=-2,-1,0,+1,+2$

or $l=3$ then, $m=-3,-2,-1,0,+1,+2,+3$

Question 2.28 (ii) List the quantum numbers (*m* $_l$ and *l*) of electrons for 3d orbital.

Answer :

For 3d-orbital, the values of Principal quantum number is $(n) = 3$ and Azimuthal quantum number $(l) = 2$.

Therefore, for $(l) = 2$,

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

i.e., $m_l = -2, -1, 0, +1, +2$

Question 2.28 (iii) Which of the following orbitals are possible?

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

Answer :

$1p$ is **NOT** possible because for $n = 1$, the value of l is zero. $(for\ p, l = 1)$

$2s$ is possible because, when $n = 2, l = 0, 1$. $(for\ s, l = 0)$.

$2p$ is possible because when $n = 2, l = 0, 1$. $(for\ p, l = 1)$.

$3f$ is **NOT** possible because for $n = 3$, the value of $l = 0, 1, 2$. $(for\ f, l = 3)$

Question 2.29 Using s, p, d notations, describe the orbital with the following quantum numbers.

(a) $n = 1, l = 0$

Answer :

Here, n is principal quantum number and l is azimuthal quantum number.

Then the orbital with given quantum numbers $n = 1, l = 0$ is $1s$ which can have a maximum of 2 electrons.

Question 2.29 Using s, p, d notations, describe the orbital with the following quantum numbers.

(b) $n = 3, l = 1$

Answer :

Here, n is principal quantum number and l is azimuthal quantum number.

Then the orbital with given quantum numbers $n = 3, l = 1$ is $3p$ which can have a maximum of 6 electrons.

Question 2.29 Using s, p, d notations, describe the orbital with the following quantum numbers.

(c) $n = 4, l = 2$

Answer :

Here, n is principal quantum number and l is azimuthal quantum number.

Then the orbital with given quantum numbers $n = 4, l = 2$ is $4d$ which can have a maximum of 10 electrons.

Question 2.29 Using s, p, d notations, describe the orbital with the following quantum numbers.

(d) $n = 4, l = 3$

Answer :

Here, n is principal quantum number and l is azimuthal quantum number.

Then the orbital with given quantum numbers $n = 4, l = 3$ is $4f$ which can have a maximum of 14 electrons.

Question 2.30 (a) Explain, giving reasons, which of the following sets of quantum numbers are not possible.

$n = 0, l = 0 \quad m_1 = 0, m_s = +1/2$

Answer :

Given quantum numbers : $n = 0, l = 0 \quad m_1 = 0, m_s = +1/2$

NOT possible, because n cannot be equal to zero.

Question 2.30 (b) Explain, giving reasons, which of the following sets of quantum numbers are not possible.

$n = 1, l = 0 \quad m_1 = 0, m_s = -1/2$

Answer :

Given quantum numbers : $n = 1, l = 0 \quad m_1 = 0, m_s = -1/2$

It is possible and it is **1s orbital.**

Question 2.30 (c) Explain, giving reasons, which of the following sets of quantum numbers are not possible.

$n = 1, l = 1 \quad m_1 = 0, m_s = +1/2$

Answer :

Given quantum numbers : $n = 1, l = 1 \quad m_1 = 0, m_s = +1/2$

It is **NOT** possible because when $n = 1,\ l \neq 1$.

Question 2.30 (d) Explain, giving reasons, which of the following sets of quantum numbers are not possible.

$n = 2, l = 1 \quad m_1 = 0, m_s = -1/2$

Answer :

Given quantum numbers :

$n = 2, l = 1 \quad m_1 = 0, m_s = -1/2$

It is possible and it is **2p orbital.**

Question 2.30 (e) Explain, giving reasons, which of the following sets of quantum numbers are not possible.

$n = 3, l = 3 \quad m_1 = -3, m_s = +1/2$

Answer :

Given quantum numbers :

$n = 3, l = 3 \quad m_1 = -3, m_s = +1/2$

It is **NOT** possible because when $n = 3$, $l \neq 3$.

Question 2.30 (f) Explain, giving reasons, which of the following sets of quantum numbers are not possible.

$n = 3, l = 1 \quad m_l = 0, m_s = +1/2$

Answer :

Given quantum numbers :

$n = 3, l = 1 \quad m_l = 0, m_s = +1/2$

It is possible and it is **3p** orbital.

Question 2.31 How many electrons in an atom may have the following quantum numbers?

(a) $n = 4, m_s = -{}^1/_2$

Answer :

The total number of electrons in an atom for a value of n is given by: $n = 2n^2$

Therefore, the total no. of electrons when $n = 4$,

$\Rightarrow n = 2 \times 4^2 = 32$ and half of them i.e. 16 will have $m_s = -\frac{1}{2}$.

Question 2.31 How many electrons in an atom may have the following quantum numbers?

(b) $n = 3,\ l = 1$

Answer :

When $n = 3\ and\ l = 1$ then it is $3s$ orbital which can have **2 electrons** .

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 :

According to the Bohr's postulate of angular momentum,

$$mvr = \frac{nh}{2\pi}$$

which can be written as: $2\pi r = \frac{nh}{mv}$ **......................................(1)**

Then according to de Broglie's equation for wavelength,

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

Now, substituting the values of equation (2) in equation (1) we get,

$$2\pi r = n\lambda$$

Thus, the circumference of the Bohr's orbit for 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 the transition of H-like particles,

$$\vartheta = \frac{1}{\lambda} = R_H Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

For He^+ transition spectrum,

$Z = 2$, $n_2 = 4$, and $n_1 = 2$

Therefore,

$$\vartheta = \frac{1}{\lambda} = R_H 2^2 \left[\frac{1}{2^2} - \frac{1}{4^2}\right] = \frac{3R_H}{4}$$

Then for the hydrogen spectrum,

$\vartheta = \frac{3R_H}{4}$ and $Z = 1$

Therefore,

$$\vartheta = \frac{1}{\lambda} = R_H \times 1 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\Rightarrow R_H \times 1 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = \frac{3R_H}{4}$$

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

The values of $n_1 \ and \ n_2$ can be found by the hit and trial method in the above equation.

So, we get $n_1 = 1$ and $n_2 = 2$, i.e., the transition is from $n = 2$ to $n = 1$.

Question 2.34 Calculate the energy required for the process

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

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

Answer :

For the hydrogen-like particles,

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

For H-atom, Ionization energy:

$$I.E. = E - E_1 = 0 - (-\frac{2\pi^2me^4}{1^2h^2}) = \frac{2\pi^2me^4}{h^2}$$

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

For the given process, the energy required will be:

$$E_n - E_1$$

$$= 0 - (-\frac{2\pi^2me^4}{1^2h^2})$$

$$= 4 \times \frac{2\pi^2me^4}{h^2}$$

$$= 4 \times 2.18 \times 10^{-18} J$$

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

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 :

Given the diameter of a carbon atom which is $0.15\ nm = 0.15 \times 10^{-9} m = 1.5 \times 10^{-10} m$

Then 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 will be:

$$= \frac{2 \times 10^{-1} m}{1.5 \times 10^{-10} m} = 1.33 \times 10^{9}$$

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 :

The arrangement length is given which is $2.4\ cm = 2.4 \times 10^{-2} m$.

and the number of atoms of carbon which are arranged in this length is given 2×10^{8} .

Let the radius of carbon atom be $'r'$ then,

$$(2r) \times (2 \times 10^{8}) = 2.4 \times 10^{-2} m$$

$$\Rightarrow r = \frac{2.4 \times 10^{-2} m}{(2 \times 10^{8}) \times 2} = 0.060 \times 10^{-9} m$$ or $r = 0.060\ nm$

Hence the radius of carbon atom is $r = 0.060\ nm$.

Question 2.37 The diameter of zinc atom is 2.6 Å. Calculate (a) radius of zinc atom in pm

Answer :

If the diameter of zinc atom is $2.6\ \mathring{A}$ then, its **radius** would be:

$$= \frac{2.6\ \mathring{A}}{2} = 1.3\ \mathring{A} = 1.3 \times 10^{-10} m\ or\ 130\ nm$$

Question 2.37 The diameter of zinc atom is 2.6 Å. Calculate (b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

Answer :

The number of atoms present in a length of $1.6\ cm$ will be:

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

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

Question 2.38 A certain particle carries $2.5 \times 10^{-16}\ \mathrm{C}$ of static electric charge. Calculate the number of electrons present in it.

Answer :

As the charge carried by one electron is $1.602 \times 10^{-19} C$

Therefore, the number of electrons present in particle carrying $2.5 \times 10^{-16}\ \mathrm{C}$ charge will be:

$$= \frac{2.5 \times 10^{-16} C}{1.6022 \times 10^{-19} C} = 1.560 \times 10^{3}$$

$$= 1560\ electrons$$

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 $-1.282 \times 10^{-18}\ \mathrm{C}$, calculate the number of electrons present on it.

Answer :

Given charge on the oil drop is $-1.282 \times 10^{-18} C$

and the charge carried by one electron is $-1.6022 \times 10^{-19} C$

Therefore, the number of electrons present on the oil drop carrying $-1.282 \times 10^{-18} C$ charge is:

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

$$= 0.8001 \times 10 = 8.0$$

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 :

The thin foil of heavy atoms, like gold, platinum, etc. have a nucleus carrying a large amount of positive charge. Therefore, some α particles will easily get deflected back.

These α particles also deflect through small angles because of the large number of a positive charge.

Hence if we use light atoms, their nuclei will have a small positive charge, hence the number of αparticles getting deflected even through small angles will be negligible.

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 way to represent an element along with its atomic mass (A) and atomic number (Z) is $^{A}_{Z}\text{X}$. Here the atomic number of an element is fixed. However, its mass number is not fixed as it depends upon the isotope taken.

Hence, it is essential to indicate the mass number.

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 of an atom be $'p'$.

Then the number of neutrons will be,

$$p + \frac{31.7}{100}p = 1.317p$$

and the mass number is 81 (Given).

Mass number = number of neutrons + number of protons.

Therefore,

$$81 = 1.317p + p$$

$$\Rightarrow p = \frac{81}{2.317} = 35$$

Thus, there are 35 numbers of protons which is also its atomic number.

Hence, the symbol for the element is $^{81}_{35}\mathrm{Br}$.

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 :

Given an ion has mass number 37 and possesses one unit of negative charge.

Let the number of electrons be $'e'$ then,

the number of neutrons will be: $e + \frac{11.1e}{100} = 1.111e$

The number of electrons in the neutral atom $= e - 1$ (ion possesses one unit of negative charge).

Therefore, the number of protons will be $= e - 1$.

Mass number = number of protons + number of neutrons

therefore,

$1.111e + e - 1 = 37$

$\Rightarrow 2.111e = 38$

$\Rightarrow e = 18$

Therefore, the number of protons is equal to the atomic number.

$\Rightarrow e - 1 = 18 - 1 = 17$

Hence, the symbol for an ion will be:

${}^{37}_{17}\mathrm{Cl}^{-1}$

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 :

Given an ion has mass number 56 and possesses three units of negative charge.

Let the number of electrons be $'e'$ then,

the number of neutrons will be: $e + \frac{30.4e}{100} = 1.304e$

The number of electrons in the neutral atom $= e + 3$ (ion possesses three units of positive charge).

Therefore, the number of protons will be $= e + 3$.

Mass number = number of protons + number of neutrons

therefore,

$1.304e + e + 3 = 56$

$\Rightarrow 2.304e = 53$

$\Rightarrow e = 23$

Therefore, the number of protons is equal to the atomic number.

$\Rightarrow e + 3 = 23 + 3 = 26$

Hence, the symbol for an ion will be:

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

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 of radiations will be:

Radiation from FM radio **<** amber light from traffic signal **<** radiation from microwave oven **<** X-rays **<**cosmic rays from outer space.

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 :

The energy emitted by the nitrogen laser is:

$$= 3.3 \times 10^{6} J$$

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,

Answer:

The wavelength of neon gas is $616\ nm$ or $616 \times 10^{-9} m$

Hence the frequency of this radiation will be:

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

Question 2.47 Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate (b) distance traveled by this radiation in 30 s

Answer :

The velocity of neon gas radiation is $3.0 \times 10^8 m/s$.

Therefore, the distance travelled in $30s$ will be:

$$30s \times (3.0 \times 10^8 m/s) = 9.0 \times 10^9 m$$

Question 2.47 Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate (c) energy of quantum

Answer :

The energy of quantum will be:

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

$$= \frac{(6.626 \times 10^{-34} Js) \times (3.0 \times 10^8 m/s)}{616 \times 10^{-9} m} = 32.27 \times 10^{-20} J$$

Question 2.47 Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate (d) number of quanta present if it produces 2 J of energy.

Answer :

If it produces 2J of energy then, the number of quanta present in it will be $'N'$.

Therefore,

$E = Nh\nu$ or $N = \frac{E}{h\nu}$,

Where $E = 2J$ and $h\nu = 32.27 \times 10^{-20} J$ from the previous part.

$$N = \frac{2J}{32.27 \times 10^{-20} J} = 6.2 \times 10^{18}$$

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 :

Let the number of photons received by the detector be $'N'$.

Then, the total energy it receives from the radiation of $600 nm$ will be:

$E = Nh\nu = N\frac{hc}{\lambda}$ Or $N = \frac{\lambda E}{hc}$

Where,

$\lambda = 600\ nm = 600 \times 10^{-9} m$,

$E = 3.15 \times 10^{-18} J$,

$h = 6.626 \times 10^{-34} Js$ and $c = 3.0 \times 10^{8} m/s$

Substituting the values in the equation above, we get

$$N = \frac{(3.15 \times 10^{-18} J) \times (600 \times 10^{-9} m)}{(6.626 \times 10^{-34} Js) \times (3.0 \times 10^{8} m/s)}$$

$= 9.51 \approx 10$ **number of photons is received by the detector.**

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 :

Given the duration of a radiation source $= 2\ ns = 2 \times 10^{-9} s$ and the number of pulse source is $= 2.5 \times 10^{15}$, then its frequency will be:

$$\nu = \frac{1}{2 \times 10^{-9} s} = 0.5 \times 10^{9} s^{-1}$$

and the energy of the source for the given frequency will be:

$$E = Nh\nu = (2.5 \times 10^{15})(6.626 \times 10^{-34} Js) \times (0.5 \times 10^{9} s^{-1})$$

$$= 8.28 \times 10^{-10} J$$

Question 2.50 The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

Answer :

We have two wavelengths of $\lambda_1 = 589\ nm = 589 \times 10^{-9} m$ and $\lambda_2 = 589.6\ nm = 589.6 \times 10^{-9} m$.

Calculating the frequency for each:

$$\nu_1 = \frac{c}{\lambda_1} = \frac{3.0 \times 10^8 m/s}{589 \times 10^{-9} m} = 5.093 \times 10^{14} s^{-1}$$

$$\nu_2 = \frac{c}{\lambda_2} = \frac{3.0 \times 10^8 m/s}{589.6 \times 10^{-9} m} = 5.088 \times 10^{14} s^{-1}$$

Therefore, the energy difference between two excited states will be:

$$\triangle E = E_2 - E_1 = h(\nu_2 - \nu_1)$$

$$= (6.626 \times 10^{-34} Js)(5.093 - 5.088) \times 10^{14} s^{-1}$$

$$= 3.31 \times 10^{-22} J$$

Question 2.51 The work function for caesium atom is 1.9 eV. Calculate (a) the threshold wavelength

Answer :

Given the work function for the Caesium atom is $1.9\ eV$.

i.e., $W_o = h\nu_o = \frac{hc}{\lambda_o}$ Or $\lambda_o = \frac{hc}{W_o}$

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

$$\therefore 1.9\ eV = 1.9 \times 1.602 \times 10^{-19} J = 3.0438 \times 10^{-19} J$$

$$\lambda_o = \frac{(6.626 \times 10^{-34} Js)(3.0 \times 10^{8}) J}{3.0438 \times 10^{-19} J} = 6.5306 \times 10^{-7} m$$

Therefore, the threshold wavelength is 6.53× 10 $^{-7}$ m.

Question 2.51 The work function for caesium atom is 1.9 eV. Calculate (b) the threshold frequency of the radiation.

Answer :

To find threshold frequency:

$$W_o = h\nu_o$$

Where

h = Planck's constant

ν_o = threshold frequency

$$\nu_o = \frac{W_o}{h}$$

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

$$= \frac{3.0438 \times 10^{-19} J}{6.626 \times 10^{-34} Js}$$

$$= 4.5937 \times 10^{14} Hz$$

Question 2.51 The work function for caesium atom is 1.9 eV. If the caesium element is irradiated with a wavelength 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

Answer :

Finding the kinetic energy of the ejected electrons:

K.E of the ejected photoelectron:

$$= h[\nu - \nu_o] = hc\left[\frac{1}{\lambda} - \frac{1}{\lambda_o}\right]$$

$$= (6.626 \times 10^{-34} Js)(3.0 \times 10^{8} m/s)\left[\frac{1}{500 \times 10^{-9} m} - \frac{1}{654 \times 10^{-9} m}\right]$$

$$= 19.878 \times 10^{-26} Jm\left[\frac{154}{327000} \times 10^{9} m^{-1}\right]$$

$$= 9.361 \times 10^{-20} J$$

Finding the Velocity of the ejected electrons:

$$KE = \frac{1}{2}mv^2$$

Where,

m = mass of electron

v = velocity of electron

Therefore, the velocity is given by,

$$v = \sqrt{\frac{2KE}{m}}$$

$$v = \sqrt{\frac{2 \times 9.361 \times 10^{-20} J}{9.1 \times 10^{-31} kg}} = 4.52 \times 10^{5}\ m/s$$

Question 2.52 Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength

Answer :

Let us assume the threshold wavelength to be $\lambda\ nm$ and the kinetic energy of the radiation is given as:

$$h(\nu - \nu_o) = \frac{1}{2}mv^2$$

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

...............................(1)

Similarly, we can also write,

.................................(2)

.................................(3)

Now, dividing the equations (3) with (1),

$$\frac{5\lambda - 2000}{4\lambda - 2000} = \frac{(5.35)^2}{(2.55)^2}$$

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

$$\Rightarrow 17.6070\lambda - 5\lambda = 8803.537 - 2000$$

$$\Rightarrow \lambda = \frac{6805.537}{12.607} = 539.8\ nm$$

Therefore, the wavelength is $540\ nm$.

Question 2.52 Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (b) Planck's constant.

Answer :

We have the threshold wavelength. $\lambda_o = 540\ nm$

Then substituting this value in any of the equation (look in the previous part), we get

$$\left(\frac{hc}{10^{-9}}\right)\left(\frac{1}{400} - \frac{1}{540}\right) = \frac{1}{2}m(5.35 \times 10^3)^2$$

Taking the mass of an electron to be $9.11 \times 10^{-31} kg$.

$$h = \frac{(9.11 \times 10^{-31})(5.35 \times 10^3)^2 \times 10^{-9} \times (400 \times 540)}{2 \times (3.0 \times 10^8)(140)}$$

$= 6.705 \times 10^{-34} Js$ approximately.

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:

Given work function of the metal, W and the Wavelength, $\lambda = 256.7\ nm$

From the Law of conservation of energy, the energy of an incident photon E is equal to the sum of the work function W of radiation and its kinetic energy K.E i.e.,

The energy of incident radiation

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} Js)(3.0 \times 10^{8})m/s}{256.7 \times 10^{-9} m}$$

$= 7.74 \times 10^{-19} J$ Or

$$= \frac{7.74 \times 10^{-19} J}{1.602 \times 10^{-19}} eV = 4.83\ eV$$

Since the potential applied gives the kinetic energy to the radiation, therefore K.E of the electron $= 0.35\ eV$

Therefore, Work Function $W = 4.83 - 0.35 = 4.48\ eV$

Question 2.54: If the photon of the wavelength 150 pm strikes an atom and one of it's inner bound electrons is ejected out with a velocity of $1.5 \times 10^{7}\ \mathrm{m\ s^{-1}}$, calculate the energy with which it is bound to the nucleus.

Answer :

Given the wavelength of a photon which strikes an atom is $150\ pm = 150 \times 10^{-12} m$.

Then the energy associated with this photon will be:

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} Js)(3.0 \times 10^{8}) m/s}{150 \times 10^{-12} m}$$

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

Given the velocity of ejected inner bounded electron: $v = 1.5 \times 10^{7} m/s$.

Then, the energy associated with this electron will be, Kinetic energy.

Hence finding

$$KE = \frac{1}{2} m v^2$$

Where

m = mass of electron, v = velocity of electron

$$KE = \frac{1}{2}(9.11 \times 10^{-31} kg)(1.5 \times 10^{7})^2$$

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

Hence the energy with which the electrons are bounded to the nucleus is:

$$= 13.25 \times 10^{-16} J - 1.02 \times 10^{16} J$$

$$= 12.23 \times 10^{-16} J$$

$$= \frac{12.23 \times 10^{-16} J}{1.602 \times 10^{-19}}$$

$$= 7.63 \times 10^{3}\ eV$$

Question 2.55 Emission transitions in the Paschen series end at orbit *n = 3* and start from orbit n and can be represented as $v = 3.29 \times 10^{15}\ (\text{Hz}) \left[\frac{1}{3^2} - \frac{1}{n^2}\right]$

Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

Answer :

Given transition in the Paschen series end at orbit $n = 3$ and starts from orbit n: $v = 3.29 \times 10^{15}\ (\text{Hz}) \left[\frac{1}{3^2} - \frac{1}{n^2}\right]$ **........................(1)**

$\nu = \frac{c}{\lambda} = \frac{3.0 \times 10^{8} m/s}{1285 \times 10^{-9} m}$ **........................(2)**

Equating both (1) and (2) equations: we get

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

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

$$\left[\frac{1}{3^2} - \frac{1}{n^2}\right] s^{-1} = 0.07096 s^{-1}$$

$$\frac{1}{9} - 0.07096 = \frac{1}{n^2}$$

$$\frac{1}{n^2} = 0.0401 \Rightarrow n^2 = 25 \Rightarrow n = 5$$

Therefore, the radiation corresponding to 1285 nm lies in the **infrared 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 n^{th} orbit of H-like particles is given by:

$= 0.529 \times \frac{n^2}{Z} \mathring{A}$ Or $= 52.9 \times \frac{n^2}{Z}\ pm$

Here, starting radius, $r_1 = 1.3225\ nm = 1322.5\ pm = 52.9 n_1^2$

Ending radius, $r_2 = 211.6\ pm = 52.9 \left(\frac{n_2^2}{Z}\right)$

Therefore, $\frac{r_1}{r_2} = \frac{1322.5\ pm}{211.6\ pm} = \frac{n_1^2}{n_2^2}$

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

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

If $n_2 = 2$ and $n_1 = 5$, then the transition is from 5^{th} orbit to 2^{nd} orbit.

Therefore, it belongs to the Balmer Series .

Frequency ν is given by:

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

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

Wavelength :

$$\lambda = \frac{1}{\nu} = \frac{100}{1.097 \times 21 \times 10^7} m = 434 \times 10^{-9} m = 434\ nm$$.

Therefore, it lies in the visible range.

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

Answer :

According to de- Broglie's equation for the wavelength.

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

Given the velocity of electron $v = 1.6 \times 10^6 m/s$

and mass of electron $m = 9.11 \times 10^{-31} kg$

So, the wavelength will be:

$$\lambda = \frac{(6.626 \times 10^{-34} Js)}{(9.11 \times 10^{-31} kg)(1.6 \times 10^{6} m/s)}$$

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

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 :

Given the wavelength of neutron: $\lambda = 800\ pm$

and the mass of neutron $m = 1.675 \times 10^{-27} kg$

So, According to the de-Broglie's equation,

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

Substituting the values in above equation:

$$800 \times 10^{-12} m = \frac{6.626 \times 10^{-34} Js}{(1.675 \times 10^{-27} kg) \times (v)}$$

$$\Rightarrow v = \frac{6.626 \times 10^{-34} Js}{(1.675 \times 10^{-27} kg) \times (800 \times 10^{-12} m)}$$

$$v = 4.94 \times 10^{4} m/s$$

Question 2.59 If the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6 \text{ ms}^{-1}$, calculate the de Broglie wavelength associated with it.

Answer :

Given the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6 \text{ ms}^{-1}$.

And we know the mass of electron which is $9.11 \times 10^{-31} kg$

Hence the de-Broglie's wavelength associated with the electron will be:

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} Js}{(9.11 \times 10^{-31} kg) \times (2.19 \times 10^6 m/s)}$$

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

Question 2.60 The velocity associated with a proton moving in a potential difference of 1000 V is $4.37 \times 10^5 \text{ ms}^{-1}$. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

Answer :

Given a proton is moving with velocity $4.37 \times 10^5 \text{ ms}^{-1}$.

And if the hockey ball of mass $0.1 kg$ is also moving with the same velocity, then

According to de-Broglie's equation we have,

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

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

$$= 1.516 \times 10^{-28} m$$

Question 2.61 If the position of the electron is measured within an accuracy of $\pm 0.002 \text{ nm}$, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is $h/4\pi_m \times 0.05 \text{ nm}$, is there any problem in defining this value.

Answer :

We have given the uncertainty in position, i.e., $\triangle x = \pm 0.002\ nm = 2 \times 10^{-12} m$.

According to Heisenberg's Uncertainty Principle:

$$\triangle x \times \triangle p = \frac{h}{4\pi}$$

Where,

$\triangle x$ is uncertainty in the position of the electron.

$\triangle p$ is uncertainty in the momentum of the electron.

Then, $\triangle p = \frac{h}{4\pi \times \triangle x}$

$$\triangle p = \frac{6.626 \times 10^{-34} Js}{4\pi \times (2 \times 10^{-12} m)} = 2.636 \times 10^{-23} Jsm^{-1}$$

Or $2.636 \times 10^{-23} kgms^{-1} \left(1J - 1kgms^2 s^{-1}\right)$

The actual momentum of the electron:

$$\frac{h}{4\pi_m \times 0.05\ nm} = \frac{6.626 \times 10^{-34} Js}{4\pi \times 0.05 \times 10^{-9} m}$$

$$\Rightarrow p = 1.055 \times 10^{-24} kg\ m/s$$

Therefore, it cannot be defined because the actual magnitude of the momentum is smaller than the uncertainty.

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:

Answer :

Quantum number provides the entire information about an electron of a particular atom.

Principal quantum number $'n'$

Azimuthal quantum number $'l'$

Magnetic quantum number $'m_l'$

Spin quantum number $'m_s'$.

The orbitals occupied by the electrons are:

1. 4d-orbital

2. 3d-orbital

3. 4p-orbital

4. 3d-orbital

5. 3p-orbital

6. 4p-orbital

For the same orbitals, electrons will have the same energy and higher the value of $(n+l)$ value higher is the energy.

Therefore, the increasing order of energies:

Question 2.63 The bromine atom possesses 35 electrons. It contains 6 electrons in *2p* orbital, 6 electrons in *3p* orbital and 5 electron in *4p* orbital. Which of these electron experiences the lowest effective nuclear charge?

Answer :

As the **p-orbital** is farthest from the nucleus hence the electrons in **(4p)subshell** experiences the lowest effective nuclear charge.

Question 2.64 Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? (i) *2s* and *3s*

Answer :

Nuclear charge is defined as the net positive charge experienced by an electron in the orbital of an atom exerted by the nucleus of the atom.

Closer orbitals experience more nuclear charge than outer orbitals.

Therefore, **(i) 2s and 3s**

The 2s orbital is more closer to the nucleus than 3s orbital hence 2s will experience a larger effective nuclear charge compared to 3s.

Question 2.64 Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? (ii) *4d* and *4f*

Answer :

Nuclear charge is defined as the net positive charge experienced by an electron in the orbital of an atom exerted by the nucleus of the atom.

Closer orbitals experience more nuclear charge than outer orbitals.

Therefore, **(ii) 4d and 4f**

The 4d orbital is more closer to the nucleus than 4f orbital hence 4d will experience a larger effective nuclear charge compared to 4f.

Question 2.64 Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? (iii) *3d* and *3p*

Answer :

Nuclear charge is defined as the net positive charge experienced by an electron in the orbital of an atom exerted by the nucleus of the atom.

Closer orbitals experience more nuclear charge than outer orbitals.

Therefore, **(ii) 3d and 3p**

The 3p orbital is more closer to the nucleus than 3d orbital hence 3p will experience a larger effective nuclear charge compared to 3d.

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 the orbital of an atom exerted by the nucleus of the atom.

Silicon has a greater nuclear charge $(+14)$ than aluminium $(+13)$.

Hence, the effective nuclear charge exerted on the unpaired 3p electron of silicon would be greater as compared to that of aluminium.

Question 2.66 Indicate the number of unpaired electrons in:

(a) P

Answer :

The electronic configuration of P: $(1s^2)(2s^2)(2p^6)(3s^2)(3p_x^1 p_y^1 p_z^1)$.

Hence, the number of unpaired electrons are **3** in 3p orbital.

Question 2.66 Indicate the number of unpaired electrons in :

(b) Si

Answer :

The electronic configuration of Si : $(1s^2)(2s^2)(2p^6)(3s^2)(3p_x^1 p_y^1)$.

Hence, the number of unpaired electrons are **2** in $3p(p_x\ and\ p_y)$ orbital.

Question 2.66 Indicate the number of unpaired electrons in :

(c) Cr

Answer :

The electronic configuration of Cr : $(1s^2)(2s^2)(2p^6)(3s^2)(3p^6)(3d^5)(4s^1)$.

Hence, the number of unpaired electrons are **6** (1 in 4s and 5 in 3d) **.**

Question 2.66 Indicate the number of unpaired electrons in :

(d) Fe

Answer :

The electronic configuration of Fe : $(1s^2)(2s^2)(2p^6)(3s^2)(3p^6)(3d^6)(4s^2)$.

Hence, the number of unpaired electrons are **4** (in 3d) **.**

Question 2.66 Indicate the number of unpaired electrons in :

(e) Kr

Answer :

As Krypton (Kr) is a noble gas whose atomic number is 36 and have all orbitals filled.

Hence, there are no unpaired electrons in Kr element.

Question 2.67 (a) How many subshells are associated with n = 4?

Answer :

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

Therefore, for $n=4$,

l can have values from 0 to 3 .

i.e., $l=0,1,2,3$.

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

Question 2.67 (b) How many electrons will be present in the subshells having *m* *s* value of *–1/2* for *n* = *4* ?

Answer :

The number of orbitals in the n^{th} shell is equal to n^2 .

So, for $n = 4$, there are $4^2 = 16$ orbitals present.

As each orbital has one electron with spin $m_s = -\frac{1}{2}$.

Hence, there will be 16 electrons with $m_s = -\frac{1}{2}$.