NCERT Solutions for Class-XI Chemistry

Chapter-2
NCERT Chemistry Class 11

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

1. (i) Mass of one electron = 9.10939×10^{-31} kg
$\therefore$ Number of electrons that weigh 9.10939×10^{-31} kg = 1
$\therefore$ 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

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?

2. No the answer will not vary with the change in temperature and pressure because the number of subatomic particles like protons, neutrons and electrons is fixed for each and every element and it does not vary with temperature and pressure.

3. How many neutrons and protons are there in the following nuclei?
$^{13}_{6}C$, $^{16}_{8}O$, $^{24}_{12}Mg$, $^{56}_{26}Fe$, $^{88}_{38}Sr$

3. 136C: 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)
= 88 – 38 = 50

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

4. Suppose for element X with following representation
$_{Z}X^{A}$
Where A = Mass Number of Element and Z = Atomic Number of the element.
(i) Z = 17, A = 35.
The representation with given atomic number and mass number is as follows:
Element with Atomic Number 17 is Chlorine.
$_{17}Cl^{35}$
(ii) Z = 92, A = 233.
The representation with given atomic number and mass number is as follows:
Element with Atomic Number 92 is Uranium.
$_{92}U^{233}$
(iii) Z = 4, A = 9
Ans: The representation with given atomic number and mass number is as follows:
Element with Atomic Number 4 is Beryllium.
$_{4}Be^{9}$

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

5. From the expression,

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

We get,

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

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

Thus, frequency of yellow light emitted from the sodium lamp

= 5.17×10^{14} s^{-1}

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

$$= \frac{1}{580\times10^{-9}} = 1.72\times10^6\,m^{-1}$$

6. Find energy of each of the photons which

(i) correspond to light of frequency 3× 1015 Hz.

(ii) have wavelength of 0.50 Å.

6. By Planck's Quantum Theory we have the following relation:

Energy, E = h×v

Where

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

v = frequency

Using the Planck's relation we will solve the numerical.

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

Frequency, v = 3×10^{15} Hz

By Planck's relation we have,

Energy, E = h×v

= $[6.626\times10^{-34}] \times [3\times10^{15}]$

= 1.9878×10^{-18} J

Therefore the energy of the photon corresponding to light of frequency 3×10^{15} Hz is 1.988×10^{-18} J.

(ii) Have wavelength of 0.50 Å.

We know the following basic relation,

Speed of Light = [Frequency] × [Wavelength]

We know speed of light = 3×10^8 m/s

Frequency, v = $[3\times10^8] / [0.5 \times 10^{-10}]$

v = 6×10^{18} Hz

By Planck's relation we have,

Energy, E = h×v

= $[6.626\times10^{-34}] \times [6\times10^{18}]$

= 3.9756×10^{-15} J

Therefore the energy of the photon corresponding to light of frequency 3×10^{15} Hz is 3.98×10^{-15} J.

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

7. 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 $= \dfrac{c}{v}$

Where,

c = velocity of light in vacuum = 3×108 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 of $(\bar{v})$ light $= \dfrac{1}{\lambda} = \dfrac{1}{6.0\times10^{-2}} = 1.6\times10^{1}\,\text{m}^{-1} = 16.66\,\text{m}$

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

8. By Planck's relation we have,

Energy, E = h×v

But we know v = [c] / [λ]

Where

c = Speed of Light

v= Frequency

λ = Wavelength

So E = hc /λ

$= [[6.626\times10^{-34}] \times [3\times10^{8}]] / [4000\times10^{-12}]$

$= [1.9878\times10^{-25}] / [4000\times10^{-12}]$

$= 4.9695\times10^{-17}$ J

4.9695×10^{-17} J is the Energy of 1 photon

So 1 J is the energy of X number of photons

$X = [1] / [4.9695\times10^{-17}]$

$X = 2.012\times10^{16}$ photons

So the number of photons which provide 1 J of energy is 2.012×10^{16} photons.

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

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

9. (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:

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

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

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

$= h\nu - h\nu_0$

$= (E - W)eV$

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

$= (3.1020 - 2.13)$ eV

$= 0.9720$ eV

Hence, the kinetic energy of emission is 0.97 eV.

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

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

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

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:

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

$$= \sqrt{0.3418\times10^{12}\,m^2s^{-2}}$$

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

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

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}.

10. Finding Energy of Photon:

By Planck's relation we have,

Energy, E = h×v

But we know v = [c] / [λ]

Where

c = Speed of Light

v= Frequency

λ = Wavelength

So E = hc /λ

$= [[6.626\times10^{-34}] \times [3\times10^{8}]] / [242\times10^{-9}]$

$= [1.9878\times10^{-25}] / [242\times10^{-9}]$

$= 8.214\times10^{-19}$ J/atom

To convert the energy from J/atom to kJ mol^{-1} we carry out the following conversion process:

$E = [8.214\times10^{-19}\times 6.023\times10^{23}] / 1000$

$= [494.729\times 10^{3}] / 1000$

$= 494.729$ kJ/mol

Therefore the energy of the photon 494.73 kJ/mol

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

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

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

Substituting the values in the given expression of E:

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

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

Rate of emission of quanta per second

$$= \frac{25}{34.87\times10^{-20}} = 7.169\times10^{19}\,s^{-1}$$

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.

12. To find threshold frequency, v_o:

We know the following basic relation,

Speed of Light = [Frequency] × [Wavelength]

We know speed of light = 3×10^{8} m/s

Frequency, $v_o = [3\times10^{8}] / [6800 \times 10^{-10}]$

$v_o = 4.412\times10^{14}$ Hz

Work Function, $W_0 = [6.626\times10^{-34}] \times [4.412\times10^{14}]$

$= 2.923\times10^{-19}$ J

Therefore the threshold frequency is 4.41×10^{14} Hz and the work function is 2.92×10^{-19} J.

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?

13. 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:

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

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

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

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

The negative sign indicates the energy of emission.

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

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

Substituting the values in the given expression of λ:

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

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

$= 4.863 \times 10^{-9}$ m

= 486 nm

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).

14. The expression of energy is given by,

$E_n = -[2.18\times 10^{-18}]\ Z^2/n^2$

Where,

Z = atomic number of the atom

N = principal quantum number

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

Therefore $\Delta E = E2-E1 = -21.8\times 10^{-19}\times [1/n_2^2-1/n_1^2]$

$= 21.8\times 10^{-19} \times [1/n_2^2-1/n_1^2]$

$=21.8\times 10^{-19}\times [1/5^2-1/\infty]$

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

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

Therefore $\Delta E' = 21.8\times 10^{-19}\times [1/1^2 - 1/\infty]$

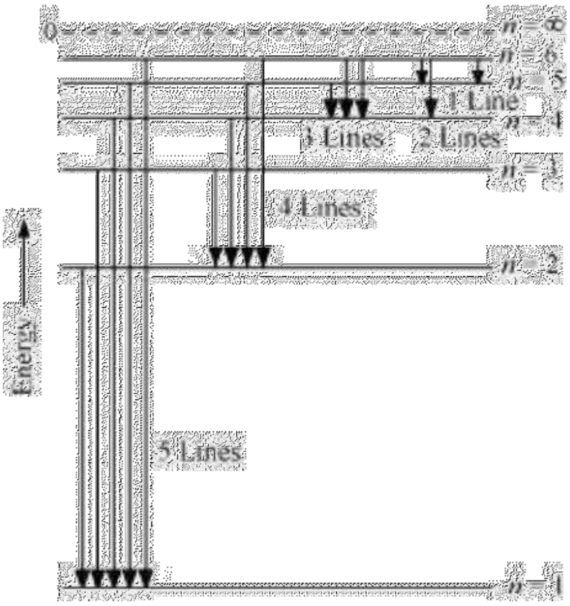
$= 21.8\times 10^{-19}J$

Now $\Delta E'/\Delta E = 21.8\times 10^{-19}/8.72\times 10^{-20} = 25$

Thus the energy required to remove electron from 1st orbit is 25 times than the required to electron from 5th orbit.

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?

15. 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 nth level drops down to the ground state is given by $\frac{n(n-1)}{2}$.

Given, n = 6

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

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.

16. Energy of an electrons = $-(2.18 \times 10^{-18})/n^2$

Where n = principal quantum number

Now the Energy associated with the fifth orbit of hydrogen atom is

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

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

(ii) Radius of Bohr's nth for hydrogen atom is given by,

R_n = (0.0529nm) / n2

For,

N = 5

$R_5 = (0.0529 \text{ nm}) (5)^2$

R_5 = 1.3225 nm

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

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

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

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

For $(\bar{\nu})$ to be minimum, n_f should be minimum.

For the Balmer series, a transition from $n_i = 2$ to $n_f = 3$ is allowed.

Hence, taking $n_f = 3$, we get:

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

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

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

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

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

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.

18. Finding Energy:

To convert energy from ergs to joules

1 ergs is equal to 10^{-7} Joules

So -2.18×10^{-11} ergs $= -2.18 \times 10^{-11} \times 10^{-7}$

So Ground State Electron energy = $= -2.18 \times 10^{-18}$ J

Energy to shift the electron from n = 1 to n= 5 state is given by following relation:

$Eh = E_5 - E_1$

The energy of hydrogen atom is given by the following equation:

$En = -2n^2 me^4 Z^2 / n^2 h^2$

Where

m = mass of electrons

Z = atomic mass of atom

E = charge of electron

h = Planck's constant

So the energy required is $= \dfrac{-2.18 \times 10^{-18}}{5^2} - \dfrac{-2.18 \times 10^{-18}}{1^2}$

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

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

Therefore the energy to shift the electron from n = 1 to n = 5 state is 2.093×10^{-18} J.

Finding Wavelength:

By Planck's relation we have,

Energy, E = h×v

But we know v = [c] / [λ]

Where

c = Speed of Light

v= Frequency

λ = Wavelength

So E = hc / λ

λ = hc / E

= $[[6.626\times10^{-34}] \times [3\times10^{8}]] / [2.0928 \times 10^{-18}]$

= $[1.9878\times10^{-25}] / [2.0928 \times 10^{-18}]$

= 9.498×10^{-8} m

Therefore the wavelength is 9.5×10^{-8} m

19. 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. What is the longest wavelength of light in cm that can be used to cause this transition?

19. Given,

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

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

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

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

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

= 0.545×10^{-18} J

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

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

Here, λ is the longest wavelength causing the transition.

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

= 3647 × 10–10 m

= 3647 Å

20. Calculate the wavelength of an electron moving with a velocity of 2.05 × 107 ms^{-1}.

20. According to the de Broglie's Equation:

λ = h/mv

Where,

λ = wavelength of moving particle
m = mass of electron
v= velocity of particle
h= Planck's constant [6.62×10^{-34}]
$\lambda = \{[6.62\times 10^{-34}] / [9.1\times 10^{-31}] \times [2.05 \times 10^{7}]\}$
$= \{[6.62\times 10^{-34}] / [1.8655 \times 10^{-23}]$
$= 3.552\times 10^{-11}$ m
Therefore the wavelength is 3.55×10^{-11} m

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.

21. From de Broglie's equation,

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

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

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

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

$$= \sqrt{\frac{2\left(3.0\times1-0^{-25}J\right)}{9.10939\times10^{-31}kg}}$$

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

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

Substituting the value in the expression of λ:

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

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

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

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

22. Isoelectronic species are defined as those species which belong to different atoms or ions which possess same number of electrons but different magnitude of nuclear charge. A positive charged ion denotes the loss of an electron & A negative charged ion represents the gain of an electron by a species.
1] Number of electrons in sodium [Na] = 11
Therefore, Number of electrons in sodium ion [Na^{+}] = 10
2] Number of electrons in potassium ion [K^{+}] = 18
3] Number of electrons in magnesium ion [Mg^{2+}] = 10
4] Number of electrons in calcium ion [Ca^{2+}] = 18
5] Number of electrons in sulphur [S] = 16

∴ Number of electrons in sulphur ion $[S^{2-}]$ = 18

6] Number of electrons in argon [Ar] = 18

Hence, the following ions are isoelectronic species:

1] Na^+ and Mg^{2+} [10 electrons each]

2] K^+, Ca^{2+}, S^{2-} and Ar [18 electrons each]

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] $4s^2\ 3d^1$

23. (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 $1s^2\ 2^{s2}\ 2p^6$

(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).

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

24. Quantum number refers to the set of four special numbers which provide the entire information about electrons of a particular atom.

Thus the letter 'n' represents the principle quantum number and the Azimuthal quantum number 'l' can have any value between to n-1.

For n = 1 [K shell] has L = 0 [one subshell]

For n = 2 [L shell] has L = 0, 1 [2 subshell]

For n = 3 [M shell] has L = 0, 1, 2 [3 subshell]

For n = 4 [N shell] has L = 0, 1, 2, 3, [4 subshell]

For n = 5 [O shell] has L = 0, 1, 2, 3, 4, [5 subshell]

∴ For l = 4, minimum value of n = 5

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

25. For the 3d orbital:

Principal quantum number (n) = 3

Azimuthal quantum number (l) = 2

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

26. (i) In an atom the number of electrons is always equal to the number of protons to balance the charge in the atom and thus an atom is electrically neutral.
So number of protons = number of electrons = 29
(ii) The electronic configuration of the atom is given as below
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$
This electronic configuration refers to the electronic configuration of copper.

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

27. 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

28. (i) An atomic orbital has n = 3. What are the possible values of l and m_l?
(ii) List the quantum numbers (m_l and l) of electrons for 3d orbital.
(iii)Which of the following orbitals are possible?
1p, 2s, 2p and 3f

28. (i) 'n' refers to the principle quantum number
'l' refers to the azimuthal quantum number
'm' refers to the magnetic quantum number
l can have any value from to n – 1.
So for n= 3, the permissible value of l = 0, 1, 2
M can have any value from –l, -l+1 ………..0, 1 ……… l
For l = 0
m = 0
For l = 1
m = +1, 0, -1
For l = 2
m = +2, +1, 0, -1, -2
(ii) List the quantum numbers (*ml* and *l)* of electrons for 3*d* orbital.
When l = 2 the value of m = -2, -1, 0, +1, +2
Now, for the 3d orbital:
Principal quantum number [n] = 3
Azimuthal quantum number [L] = 2
Magnetic quantum number [mL] = -2, -1, 0, 1, 2

(iii) Which of the following orbitals are possible?
1*p*, 2*s*, 2*p* and 3*f*
For a given value of n the value of l can range from 0 to n-1
So for n =1, l = 0
Thus 1p is not possible
For n = 2, l = 0 and 1
Therefore, 2s and 2p are possible orbitals.
For n = 3, l = 0, 1 and 2
So 3p is not possible.

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.

29. (a) n = 1, l = 0 (Given)
The orbital is 1s.
(b) For n = 3 and l = 1
The orbital is 3p.
(c) For n = 4 and l = 2
The orbital is 4d.
(d) For n = 4 and l = 3
The orbital is 4f.

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

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

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

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

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

(d) n = 3 l = 3 m_l = – 3 $m_s = +\frac{1}{2}$

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

30. (a) $n = 0$, $l = 0$, $m_l = 0$, $m_s = + ½$
The given set of quantum numbers is not possible because the value of the principal quantum number that is [n] cannot be zero.
(b) $n = 1$, $l = 0$, $m_l = 0$, $m_s = – ½$
Ans: This set of quantum numbers is possible.
(c) $n = 1$, $l = 1$, $m_l = 0$, $m_s = + ½$

Ans: This set is not possible because for principle quantum number n = 1, the azimuthal quantum number ‘l’ cannot be one.
(d) $n = 2, l = 1, m_l = 0, m_s = -½$
Ans: This set of quantum numbers is possible.
(e) $n = 3, l = 3, m_l = -3, m_s = +½$
Ans: This set is not possible because for principle quantum number n = 3, the azimuthal quantum number ‘l’ cannot be three.
(f) $n = 3, l = 1, m_l = 0, m_s = +½$
Ans: This set of quantum numbers is possible.

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

31. (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 3s orbital. Therefore, the number of electrons having n = 3 and l = 0 is 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.

32. As we know that the angular momentum of an electron is given by mvr = [nh]/ 2π …………………. [1]
Where
m = mass of the electron
v = velocity of the electron
r = radius of the orbit
h = Planck’s constant
According to the de Broglie’s Equation:
λ = h/mv
Where,
λ = wavelength of moving particle
m = mass of electron
v= velocity of particle
h= Planck’s constant [6.62×10^{-34}]
mv = h / λ
Substituting the above value in [1]

$[hr] / \lambda = [nh]/ 2\pi$

$2\pi r = n\lambda$

Since we know that $2\pi r$ is the circumference of a circle, so it is proved in above equation that circumference of the Bohr's orbit is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.

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

33. 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

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

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

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

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

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

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

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

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.

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 $atom^{-1}$

34. The energy of hydrogen atom is given by the following equation:

$En = -2n^2me^4Z^2/n^2h^2$

Where

m = mass of electrons

Z = atomic mass of atom

E = charge of electron

h = Planck's constant
For the process mentioned in the question, energy required = $E_n - E_1$
$= 0 - [-2n^2me^4 / 1^2h^2]$
$= 4 \times 2n^2me^4/h^2$
But we know $2n^2me^4/h^2 = 2.18 \times 10^{-18}$J
$= 4 \times 2.18 \times 10^{-18}$
$= 8.72 \times 10^{-18}$J
Therefore the energy required for the process is 8.72×10^{-18}J

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.

35. 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
$$= \frac{20 \times 10^{-2}\,\text{m}}{0.15 \times 10^{-9}\,\text{m}}$$
$= 133.33 \times 10^7$
$= 1.33 \times 10^9$

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.

36. Number of carbon atoms which can be place along the line is given by n = [l] / [diameter, d]
D= [l] / [n]
D = $[2.4\times 10^{-2}] / [2 \times 10^8]$
D = 1.2×10^{-10} m
Therefore radius = $[1.2\times 10^{-10}] / 2$
= 0.6×10^{-10} m
Therefore the radius of carbon atom is 0.6 Å.

37. The diameter of zinc atom is 2.6 Å. 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.

37. (a) Radius of zinc atom $= \frac{\text{Diameter}}{2}$

$$= \frac{2.6\,\text{Å}}{2}$$

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

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

(b) Length of the arrangement = 1.6 cm

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

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

∴ Number of zinc atoms present in the arrangement

$$= \frac{1.6 \times 10^{-2}\,\text{m}}{2.6 \times 10^{-10}\,\text{m}}$$

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

$= 6.153 \times 10^{7}$

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

38. Charge on electron, = ne

Where

n = number of electrons

e = 1.6×10^{-19} J

n = $[2.5 \times 10^{-16}] / [1.6 \times 10^{-19}]$

n = 1562 electrons

Therefore the number of electrons present is 1562 electrons.

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.

39. 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

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

$= 0.8001 \times 10^{1}$

= 8.0

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?

40. In 1911, Rutherford performed alpha rays scattering experiment to demonstrate and to find the structure of atom.

Heavy atoms have a heavy nucleus carrying a great quantity of positive charge. Hence, some alpha particles are easily deflected back on hitting the nucleus. Besides a bit of alpha particles are deflected through small angles because of large positive charge on the nucleus. So with the aid of heavy nuclei atoms it was possible to understand or detect the minute deflections of alpha particles.

If light atoms are used, their nuclei will be light & moreover, they will hold a small positive charge on the nucleus. Hence, the number of particles deflected back & those deflected through some angle will be trifling. Hence deflections of alpha particles from light atoms remain unnoticeable.

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

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

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

$^{79}_{35}Br$ can be written but $^{35}_{79}Br$ 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.

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

42. Let number of protons = p
Let number of neutrons= n
$p + n = 81$
$p + \{p + [31.7\times p] / 100\} = 81$
$p + 1.317p = 81$
$2.317p = 81$
$p = 81 / 2.317$
$p = 94.95899$
$p = 35$
Atomic Number = 35
The element is bromine.
The atomic symbol is $^{81}{}_{35}Br$.

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.

43. 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)
$\therefore$ Number of protons in the neutral atom = $x - 1$
Given,
Mass number of the ion = 37

$\therefore$ (x – 1) + 1.111x = 37

2.111x = 38 x

= 18

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

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.

44. Let number of protons = p

Let number of neutrons= n

Since the ion has three units of positive charge. So the ion is tripositive.

p + 3 + n = 56

p + 3 + {p + [30.4× p] / 100} = 56

p + 3 + 1.304p = 56

2.304p = 53

p = 53 / 2.304

p = 23

Therefore the number of protons = 23 + 3

= 26

Atomic Number = 26

The element is chlorine.

The atomic symbol is $^{56}{}_{26}Cl$.

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.

45. 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

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.

46. Energy, E = nhv = nhc/λ

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]:

$= \{[5.6\times 10^{24}] \times [6.626\times 10^{-34}] \times [3\times 10^{8}\ ms^{-1}]\} / [337.1\times 10^{-9}m]$

$= [1.113168] / [337.1\times 10^{-9}m]$
$= 3.302\times 10^{6}$ J
Hence, the power of the laser is 3.302×10^{6} J.

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.

47. 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):

$$\nu = \frac{3.0\times 10^{8}\, m/s}{616\times 10^{-9}\, m}$$

$= 4.87 \times 10^{8} \times 10^{9} \times 10^{-3}\ s^{-1}\ \nu$
$= 4.87 \times 10^{14}\ s^{-1}$
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
$= (3.0 \times 10^{8}\ ms^{-1})\ (30\ s)$
$= 9.0 \times 10^{9}$ m
(c) Energy of quantum (E) = hν
$(6.626 \times 10^{-34}\ Js)\ (4.87 \times 10^{14}\ s^{-1})$
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

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

$= 6.19 \times 10^{18}$
$= 6.2 \times 10^{18}$

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.

48. By Planck's relation we have,
Energy, E = h×v
But we know v = [c] / [λ]
Where
c = Speed of Light
v = Frequency
λ = Wavelength

So $E = hc/\lambda$

$= [[6.626\times10^{-34}] \times [3\times10^{8}]] / [600\times10^{-9}]$

$= [1.9878\times10^{-25}] / [600\times10^{-9}]$

$= 3.313\times10^{-19}$ J

Therefore number of photons received $= [3.15 \times 10^{-18}] / [3.313\times10^{-19}]$

$= 9.5$

Therefore the number of photons received cannot be in fraction

So number of photons received is 9.

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.

49. Frequency of radiation (ν),

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

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

Energy (E) of source = $Nh\nu$

Where,

N = number of photons emitted

h = Planck's constant

ν = frequency of radiation

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

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

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

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

50. 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.

50. To find frequency:

Speed of Light = [Frequency] × [Wavelength]

We know speed of light = 3×10^{8} m/s

Frequency, $v_1 = [3\times10^{8}] / [589 \times 10^{-9}]$

$v_1 = 5.093\times10^{14}$ Hz

Frequency, $v_2 = [3\times10^{8}] / [589.6 \times 10^{-9}]$

$V_2 = 5.088\times10^{14}$ Hz

Change in Energy, $\Delta E = E_2 - E_1 = h[v_2 - v_1]$

$= [6.626\times 10^{-34}] \{[5.093 - 5.088] \times 10^{14}\}$

$= [6.626\times 10^{-34}] \{[5\times 10^{-3}] \times 10^{14}\}$

$= 3.313\times 10^{-22}$J

Therefore the energy difference between two excited states is 3.31×10^{-22}J.

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.

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

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

$$\lambda_0 = \frac{\left(6.626\times10^{-34}\,\text{Js}\right)\left(3.0\times10^{8}\,\text{ms}^{-1}\right)}{1.9\times1.602\times10^{-19}\,\text{J}}$$

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

$$\nu_0 = \frac{1.9\times1.602\times10^{-19}\,\text{J}}{6.626\times10^{-34}\,\text{Js}}$$

(1 eV = 1.602×10^{-19} J) ν_0
= $4.593 \times 10^{14}\ \text{s}^{-1}$

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

(c) According to the question:
Wavelength used in irradiation (λ) = 500 nm
Kinetic energy = h ($\nu - \nu_0$)

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

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

$$= \left(1.9878\times10^{-26}\,\text{Jm}\right)\left[\frac{(653-500)10^{-9}\,\text{m}}{(653)(500)10^{-8}\,\text{m}^2}\right]$$

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

= 9.3149×10^{-20} J

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

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

$$v = \sqrt{\frac{2\left(9.3149 \times 10^{-20}\,J\right)}{9.10939 \times 10^{-31}\,kg}}$$

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

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

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

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

52. Let us suppose threshold wavelength to be λ nm the kinetic energy of the radiation is given as:

$h(v - v_o) = \frac{1}{2}mv^2$

Or

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

$$hc\left[\frac{1}{500 \times 10^{9}} - \frac{1}{\lambda \times 10^{-9}}\right] = \frac{1}{2}m\left[2.55 \times 10^5 \times 10^{-2}\right]^2$$

$$\left[\frac{hc}{10^{-9}}\right]\left[\frac{1}{500} - \frac{1}{\lambda}\right] = \frac{1}{2}m\left[2.55 \times 10^3\right]^2 \text{ .... [1]}$$

Similarly we can also write,

$$\left[\frac{hc}{10^{-9}}\right]\left[\frac{1}{450} - \frac{1}{\lambda}\right] = \frac{1}{2}m\left[3.45 \times 10^3\right]^2 \text{ .... [2]}$$

$$\left[\frac{hc}{10^{-9}}\right]\left[\frac{1}{400} - \frac{1}{\lambda}\right] = \frac{1}{2}m\left[5.35 \times 10^3\right]^2 \text{ .... [3]}$$

Dividing equation [3] and [1]

$$\frac{\left[\frac{\lambda - 400}{400\lambda}\right]}{\left[\frac{\lambda - 500}{500\lambda}\right]} = \frac{\left[5.35 \times 10^3\right]^2}{\left[2.55 \times 10^3\right]^2}$$

$$= \frac{5\lambda - 2000}{4\lambda - 2000} = \frac{[5.35]^2}{[2.55]^2}$$

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

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

$\lambda = [6805.537] / [12.607]$

$\lambda = 539.8$ nm

The wavelength is 540 nm.

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.

53. 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.,

$E = W_0 + K.E$

$\Rightarrow W0 = E - K.E$

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:

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

$= 7.744 \times 10^{\square 19}$ J

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

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

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

= 4.83 eV – 0.35 eV

= 4.48 eV

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

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

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

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

$\lambda_0 = 539.8$ nm

$\lambda_0 = 540$ nm

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.

54. Energy, $E = h\nu = hc/\lambda$

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]:
$= \{6.626 \times 10^{-34}] \times [3\times 10^{8} ms^{-1}]\} / [150\times 10^{-12}m]$
$= [1.9878 \times 10^{-28}] / [337.1\times 10^{-9}m]$
$= 1.3252 \times 10^{-15}$ J
We know the formula for kinetic energy which is given as follows:
$\frac{1}{2} mv^2$ = kinetic Energy
Where
m = mass of electron
v = velocity of electron
K.E. $= \frac{1}{2} \times \{[9.1\times 10^{-31}] \times [1.5 \times 10^{7}]^2$
$= 1.02375\times 10^{-16}$ J
Energy which bounded the electron to nucleus is:
$= 13.25\times 10^{-16}J-1.025\times 10^{-16}J$
$=12.227\times 10^{-16}$ J
$= [12.227\times 10^{-16}]/ [1.602\times 10^{-19}]$
$=7.63\times 10^{3}eV$
Therefore the energy which bounded the electron to nucleus is $7.63\times 10^{3}eV$.

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.

55. Wavelength of transition = 1285 nm
$= 1285 \times 10^{-9}$ m (Given)

$$\nu = 3.29\times10^{15}\left(\frac{1}{3^2}-\frac{1}{n^2}\right)$$

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

$$= \frac{3.0\times10^{8}ms^{-1}}{1285\times10^{-9}m}$$

Now,
$\nu = 2.33 \times 10^{14} s^{-1}$
Substituting the value of ν in the given expression,

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

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

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

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

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

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

n = 4.98 n

≈ 5

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

The spectrum lies in the infra-red region.

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.

56. The radius of the n^{th} orbit of hydrogen – like particles = $0.529n^2/Z$ Å

Now r_1 = 1.3225 nm or 1322.5 pm = $52.9n_1^2 / Z$

And

R_2 = 211.6pm = $52.9n_2^2/Z$

Taking the ratio of r1 and r_2

So r_1/r_2=1322.5 / 211.6 = n_1^2/n_2^2

$n_1^2/n_2^2 = 6.25$

$n_1/n_2 = 2.5$

Therefore $n_2 = 2$, $n_1 = 5$

By referring to Balmer Formula:

We know that

Wave Number $R_h\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]R_h$

Here $n_1 = 2$ and $n_2 = 5$ and the value of R_h = 109678

Wave Number $= 109678\left[\frac{1}{2^2} - \frac{1}{5^2}\right]$

= [109678×21] / 100

= 2303238 / 100

= 23032.38 cm^{-1}

We know that Wave Number = [1] / [Wavelength, λ]

Therefore, Wavelength = [1 / 23032.38]

$\lambda = 4.3417 \times 10^{-5}$ cm

Therefore the wavelength of the light is 4.34×10^{-5} cm

The transition belongs to visible region.

57. Dual behavior 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.

57. From de Broglie's equation,

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

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

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

∴ de Broglie's wavelength associated with the electron is 455 pm.

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.

58. According to the de Broglie's Equation:

$\lambda = h/mv$

Where,

λ = wavelength of moving particle

m = mass of electron

v= velocity of particle

h= Planck's constant $[6.62\times 10^{-34}]$

$v = \{[6.62\times 10^{-34}] / [1.675\times 10^{-27}] \times [800 \times 10^{-12}]\}$

$= \{[6.62\times 10^{-34}] / [1.34\times 10^{-36}]$

$= 494.4776$ m / s

Therefore the velocity is 494.48 m / s.

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.

59. According to de Broglie's equation, $\lambda = \frac{h}{m\nu}$

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:

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

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

$= 332 \times 10^{\square 12}$ m

$\lambda = 332$ pm

∴ Wavelength associated with the electron = 332 pm

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.

60. According to the de Broglie's Equation:

$\lambda = h/mv$

Where,

λ = wavelength of moving particle

m = mass of electron

v= velocity of particle

h= Planck's constant $[6.62\times 10^{-34}]$

$\lambda = \{[6.62\times 10^{-34}] / [0.1] \times [4.37 \times 10^{5}]\}$

$= \{[6.62\times 10^{-34}] / [43.7 \times 10^{3}]$

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

Therefore the wavelength is 1.516×10^{-38} m.

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 $h/4\pi m \times 0.05$ nm, is there any problem in defining this value.

61. 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:

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

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

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

$\Delta p = 2.637 \times 10^{-23}$ $kgms^{-1}$ (1 J = 1 $kgms^{2}s^{-1}$)

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

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

$$= \frac{6.626\times 10^{-34}\text{ Js}}{4\times 3.14\times 5.0\times 10^{-11}\text{ m}}$$

$= 1.055 \times 10^{-24}$ $kgms^{-1}$

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

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, ms = –1/2

(2) n = 3, l = 2, m_l = 1, ms = +1/2

(3) n = 4, l = 1, m_l = 0, ms = +1/2

(4) n = 3, l = 2, m_l = –2, ms = –1/2

(5) $n = 3, l = 1, m_l = -1, ms = +1/2$
(6) $n = 4, l = 1, m_l = 0, ms = +1/2$

62. Quantum number refers to the set of four special numbers which provide the entire information about electrons of a particular atom.
'n' refers to the principle quantum number
'l' refers to the azimuthal quantum number
'm' refers to the magnetic quantum number
For n = 4 and L = 2, the orbital occupied is 4d.
For n = 3 and L = 2, the orbital occupied is 3d.
For n = 4 and L = 1, the orbital occupied is 4p.
Hence, the six electrons i.e., 1, 2,3,4,5, and 6 are present in the 4d, 3d, 4p, 3d, 3p, and 4p orbital respectively.
Therefore, the increasing order of energies of the electrons is 5[3p] < 2[3d] = 4[3d] <3[4p] = 6[4p] < 1[4d].

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

63. 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, 4p orbitals are farthest from the nucleus of bromine atom with (+35) charge. Hence, the electrons in the 4p orbital will experience the lowest effective nuclear charge. These electrons are shielded by electrons present in the 2p and 3p orbitals along with the s-orbitals. Therefore, they will experience the lowest nuclear charge.

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

64. 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. The closer the orbital, the greater is the nuclear charge experienced by the electron [s] in it and the nuclear charge is inversely proportional to the distance of the electron from the nucleus.
(i) 2*s* and 3*s*
3s is farther away from the nucleus as compared to 2s. Hence 2s will experience larger effective nuclear charge as compared to 3s.
(ii) 4*d* and 4*f*
In this case 4d will experience more nuclear charge as compared to 4f as 4d is more near to the nucleus..
(iii) 3*d* and 3*p*

3f is farther away from the nucleus as compared to 3p. Hence 3p will experience larger effective nuclear charge as compared to 3f.

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

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

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

66. Indicate the number of unpaired electrons in:

(a) P,

(b) Si,

(c) Fe and

(d) Kr.

66. (a) P – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Therefore number of unpaired electron = 3

(b) Si – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Therefore number of unpaired electron = 2 [as p orbital can have a maximum of 6 electrons]

(c) Cr – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

Therefore number of unpaired electron = 6

(d) Fe – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

Therefore number of unpaired electron = 4

(e) Kr – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Krypton has no unpaired electron as it is noble gas.

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?

67. (a) n = 4 (Given)

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

$\therefore$ 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 ms value of $-\frac{1}{2}$.

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