Chapter – 2 (Structure of Atom)

Exercise Questions:

Question: 1 i) Calculate the number of electrons which will together weight one gram.

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

Answer:

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

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

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

1 (1×10^{-3}kg) / 9.10939×10^{-31} kg

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

= 0.1098×10^{28}

= 1.098×10^{27}

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

Mass of one mole of electron = (6.022×10^{23})(9.10939×10^{-31})

= 5.48×10^{-7}kg

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

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

= 9.65×10^{4}c

Question: 2 i) Calculate the total number of electrons present in one mole of Methane.

ii) Find the total number and the total mass of neutrons in 7 mg of 14C.

iii) Find the total number and the total mass of protons in 34 mg of NH3 at STP.

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

Answer:

(i) Molecule of CH_4 (methane) contains electron = 10

Therefore 1 mole (6.022×10^{23} atoms) contains electron = 6.022×10^{24}

(ii) a) 1g atom of 14C = 14g = 6.022×10^{23} atoms = $6.022 \times 10^{24} \times 8$ neutrons

Thus 14g or 14000 mg have $6.022 \times 10^{24} \times 8$ neutrons

Therefore 7 mg will have neutrons = $6.022 \times 10^{24} \times 8 / 14000 \times 7 = 2.4088 \times 10^{22}$

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

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

(iii) a) 1 mol of NH3 = 17g NH3 = 6.022×10^{23} molecules of NH3 = $(6.022\times10^{23})(7 + 3)$ proton = 6.022×10^{24} protons

Therefore 34 mg i.e 0.034 g NH3 = $6.022 \times 10^{24} \times 0.034/1 = 1.2044 \times 10^{22}$ protons

b) mass of 1 proton = 1.6726×10^{-27} kg

Therefore mass of 1.2044×10^{22} protons = $(1.6726 \times 10^{-27})(1.2044 \times 10^{22})$ kg = 2.0145×10^{-5} kg

No, the answer will not change with change in temperature & pressure

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

Answer:

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

Atomic mass = 13

Atomic number = Number of protons = 6

Number of neutrons = (Atomic mass) – (Atomic number) = 13 – 6 = 7

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

Atomic mass = 16

Atomic number = 8

Number of protons = 8

Number of neutrons = (Atomic mass) – (Atomic number) = 16 – 8 = 8

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

Atomic mass = 24

Atomic number = Number of protons = 12

Number of neutrons = (Atomic mass) – (Atomic number) = 24 – 12 = 12

${}_{26}{}^{56}$ 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

Question: 4 Write the complete symbol for the atom with the given atomic number

(Z) and atomic mass (A)

I) Z = 17 , A = 35

II) Z = 92 , A = 233

III) Z = 4 , A = 9

Answer:

1) The element with atomic number(Z) 17 & mass number (A) 35 is chlorine = $^{35}_{17}Cl$

2) The element with atomic number(Z) 92 & mass number (A)233 is uranium = $^{233}_{92}U$

3) The element with atomic number(Z) 4 & mass number (A) 9 is berellium = $_4{}^9Be$

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

Answer:

From the expression, $\lambda = c/ v$

We get,

$V = c / \lambda$ ....(i)

Where,

V = frequency of yellow light

C = velocity of the light in vacuum = $3 \times 10^8 m / s$

λ = wavelength of yellow light = 580 nm = 580×10^{-9}m substituting the values of in expression (i)

$V = 3 \times 10^8 / 580 \times {}^{10-9} = 5.17 \times 10^{14} / S$

Thus, frequency of yellow light emitted from the sodium lamp

$= 5.17 \times 10^{14} / s$

Wave number of yellow light = $1 / \lambda$

$= 1/ 580 \times 10^{-9} = 1.72 \times 10^8 / m$

Question: 6 Find the energy of each of the photons which

I) Correspond to light of frequency 3 × 10 ^ 15 Hz

II) Have wavelength of 0.50 A°.

Answer:

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

E = hv

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

ν = frequency of light = 3×10^{15} Hz

Substituting the values in the given expression of E:
E = (6.626 × 10^{-34}) (3 × 10^{15}) E = 1.988 × 10^{-18} J

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

$E = hc/\lambda$
h = Planck's constant = 6.626 × 10^{-34} Js
c = velocity of light in vacuum = 3 × 10^{8} m/s
Substituting the values in the given expression of E:
E=(6.626 x 10^{-34})(3 x 10^{8})/0.50 x 10^{-10} = 3.976 x 10^{-15} J E = 3.98 x 10^{-15} J

Question:7 Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0 × 10^-10 s.

Answer:
Frequency (v) of light = 1 / period
= 1 / 2.0 x 10^{-10} = 5.0 x 10^{9} / s
Wavelength of light = c / v
Where,
C = velocity of light in vacuum
Substituting the value in the given equation of wavelength,
λ = 3 x 10^{8} / 5.0 x 10^{9} = 6.0 x 10^{-2} m
Wave number of light = 1 / 6.0 x 10^{-2} = 1.66 x 10^{1} / m = 16.66 m

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

Answer:
Energy of a photon = hv
Energy on n photon = nhv
N = $E_n \lambda$ / hc
Where,
λ = wavelength of light = 4000 x 10^{-12} m
c = velocity of light in vacuum
h = Plank's constant
Substituting the values in expression of n:
n = 1 x (4000 x 10^{-12}) / (6.626 x 10^{-34}) (3 x 10^{8}) = 2.012 x 10^{16}
Hence, the number of photons with a wavelength of 4000 and energy of 1J are 2.012 x 10^{16}

Question: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

III) the velocity of the photoelectron.

Answer:

Wavelength of the photon, $\lambda = 4 \times 10^{-7}$ m Work function of the metal, W = 2.13 eV = 2.13 x 1.6020 x 10^{-19} J = 3.14 x 10^{-19} J

(i) Energy of photon = ch/λ = {3 x 10^{8} ms-1 x 6.626 x 10^{-34} Js}/{4 x 10^{-7} m} = 4.97 x 10^{-19} J = {4.97 x 10^{-19} J}/{1.6020 x 10^{-19} J/eV} = 3.10 eV

(ii) Kinetic energy of the emission = Ephoton - W = 4.97 x 10^{-19} J - 3.41 x 10^{-19} J = 1.56 x 10^{-19} J = 3.10 eV - 2.13 eV = 0.97 eV

(iii) Kinetic energy of the emitted photoelectron = 1/2 m x v2

Or,

$v = \sqrt{2x \text{K.E.} / m}$

$= \sqrt{2x\ 1.56 \times 10^{-19} / 9.1 \times 10^{-31}\text{kg}}$

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

Question :10 Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionization energy of sodium in kJ mol^-1.

Answer:

Energy of sodium = $N_A hc / \lambda$

= (6.023 x 10^{23} / mol)(6.626 x 10^{-34}Js)(3 x 10^{8}m / s) / 242 x 10^{-9}m

= 4.947 x 10^{5}J / mol

= 494 kJ / mol

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

Answer:

Power of bulb, P = 25 Watt= 25 J/ s

Energy of one photon, E = hv = hc/ λ

Substituting the values in expression of E:

E = (6.626 x 10^{-34})(3 x 10^{8}) / (0.57 x 10^{-6}) = 34.87 x 10^{-20}J

E = 34.87 x 10^{-20} J

Rate of emission of quanta per second

$= 25 / 34.87 \times 10^{-20}$ J
$= 7.169 \times 10^{19}$ / s

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

Answer:

Threshold wavelength of radian (λ) = $6800A^0 = 6800 \times 10^{-10}$m

Threshold frequency of the metal = $c / \lambda_0 = 3\times 10^8 / 6.8 \times 10^{-7}$m $= 4.41 \times 10^{14}$ / s

Thus, the threshold frequency of the metal is 4.41×10^{14} / s

Hence, the work function of the metal = hv_0

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

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

Question :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:

According to Formula :

Wave number=$\nu=R[1/n_1^2 -1/n_2^2]$

where

R=109678 cm^{-1}n1=2n2=4

$\nu=109678[1/2^2 - 1/4^2]$

=109678[(4-1)/16]

=109678x3/16

As we know that wave number=1/Wavelength

$\Rightarrow \nu=1/\lambda$

$\lambda=1/\nu$

= 1/[109678x3/16]

= 16/109678x3

$= 486\times10^{-7}$ cm

$= 486\times 10^{-9}$m

= 486 nm

∴wavelength of light emitted is 486 nm

Question :14 How much energy is required to ionise a H atom if the electron occupies n = 5 orbit? Campare your answer with the ionization enthalpy of H atom.

Answer:

The expression of energy is given by,

Where,

Z = atomic number of the atom

n = principal quantum number

For ionization from n1 = 5 to ,

Therefore ΔE = E2- E1 = - 21.8 X10^{-19} (1/n22-1/n12)

= 21.8 X10^{-19} (1/n22-1/n12)

= 21.8 X10^{-19} (1/52-1/∞)

= 8.72 x 10^{-20} J

For ionization from 1st orbit, n1= 1,

Therefore ΔE' = 21.8x10^{-19} (1/12-1/∞)

= 21.8x10^{-19} J

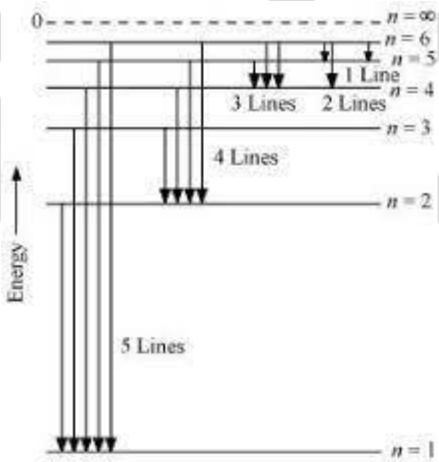
Now ΔE'/ ΔE = 21.8x10^{-19} / 8.72x10^{-20} = 25

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

Question :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:

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



Hence, a total number of (5 + 4 + 3 + 2 + 1) 15 lines will be obtained in the emission spectrum. The number of spectral lines produced when an electron in the n^{th} level drops down to the ground state is given by, Given = n(n – 1) / 2

n = 6

Number of spectral lines = 6(6 – 1) / 2 = 15

Question :16 I) The energy associated with the first orbit in the hydrogen atom is -2. 18 × 10^-18 J atom^-1. What is the energy associated with the fifth orbit?

II) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

Answer:

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

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

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

ii.) Radius of Bohr's nth orbit for hydrogen atom is given by:

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

For,

n = 5

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

$r_5 = 1.3225$ nm.

Question :17 Calculate the wavelength for the longest wavelength transition in the Balmer series of atomic hydrogen.

Answer:

According to Balmer formula

$\bar{\nu} = 1/\lambda = RH[1/n1^2 - 1/n2^2]$

For the Balmer series, ni = 2.

Thus, the expression of wavenumber($\bar{\nu}$) is given by,

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

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, nf should be minimum. For the Balmer series, a transition from ni = 2 to nf = 3 is allowed. Hence, taking nf = 3, we get:

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

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

$= (1.097 \times 10^7)[5 / 36]$

$\bar{\nu} = 1.5236 \times 106\ m^{-1}$

Question :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:

Here we have to use formula ,

$\Delta E = E_5 - E_1$

Where E_5 denotes energy in 5th orbit.

E_1 denotes energy in 1st orbit.

$\Delta E = E_5 - E_1 = 2.18 \times 10^{-11} [1/n_1^2 - 1/n_5^2]$ erg

Because we know,

$E_f - E_i = 2.18 \times 10^{-11} (1/n_i^2 - 1/n_f^2)$ erg]

$\Delta E = 2.18 \times 10^{-11} [1/1^2 - 1/5^2]$ erg

$= 2.18 \times 10^{-11} [1 - 1/25]$ erg

$= 2.18 \times 10^{-11} \times 24/25$ erg

$= 2.0928 \times 10^{-11}$ erg or, 2.0928×10^{-18} J

Now,

$\Delta E = hc/\lambda$

$\lambda = hc/\Delta E$

$= 6.626 \times 10^{-34}$ Js $\times 3 \times 10^{8}$ m/s/2.0928×10^{-18} J

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

$= 949.8 \times 10^{-10}$ m

= 949.8 A°

Hence, wavelength = 949.8 A°

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

Answer:

Energy required to shift an electron from n = 2 to n = ∞.

$\Delta E = E_\infty - E_2$

Here A/C to question,

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

so,

$\Delta E = (-2.18 \times 10^{-18})/\infty^2 - (-2.18 \times 10^{-18})/2^2$

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

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

Now,

$\Delta E = hc/\lambda$

$\lambda = hc/\Delta E$

$= 6.626 \times 10^{-34}$ Js $\times 3 \times 10^{8}$ m/s $/5.45 \times 10^{-19}$ J

$= 3.647 \times 10^{-7}$ m

$= 364.7 \times 10^{-9}$ m

= 364.7 nm

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

Answer:

According to de Broglie;s equation,

$\lambda = h/mv$

Where,

λ = wavelength of moving particle

m = mass of particle

v = velocity of particle

h = plank' s constant

Substituting the values in given expression

$\lambda = 6.626 \times 10^{-34}$Js / $(9.10939 \times 10^{-31}$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 2.05×10^{7}m / s is 3.548×10^{-11}m.

Question : 21 The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate it's wavelength.

Answer:

Given:

Mass of an electron = 9.1x10 -31 kg

Kinetic energy = 3.0x10 -25 J

h=6.6x10 -34 Js

By using de- Broglie equation :

Wavelength = h/mv = h/ p ,where h = Planck's constant, p= momentum

As P= $\sqrt{2Em}$, by subtituting the value of p we get,

Wavelength = h / $\sqrt{2Em}$ (E= kinectic energy)

Wavelength = 6.6 x 10-34/$\sqrt{2 \times 3.0 \times 10^{-25} \times 9.1 \times 10^{-31}}$

= 6.6 x 10 -34/$\sqrt{54.6 \times 10^{-56}}$

= 6.6 x 10 -34/ 7.4 x 10 - 28

= 0.89 x 10 -6 m

Question : 22 Which of the following are isoelectric species i.e., those having the same number of electrons?

Na^+, K^+, Mg^2+, Ca^2+, S^2-, Ar.

Answer:

Isoelectronic species have the same number of electrons but different atomic numbers

Na^+ , number of e^- = 11 - 1 = 10 e^-

Mg^{2+}, number of e^- = 12 - 2 = 10e^-

K^+, number of e^- = 19 - 1 = 18 e^-

Ca^{2+}, number of e^-= 20 -2 = 18 e^-

S^{2-}, number of e^- = 16 +2 = 18e^-

Ar , number of e^- = 18e^-

Now, we can observe.

Isoelectronic species are

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

(ii)K^+, Ca^{2+},S^{2-} and Ar

Question :23 I. Write the electronic configuration of the following ions: a. H- b.) Na+ c.) O2- d.) F-

II. What are the atomic numbers of the elements whose outermost electrons are represented by a.) 3s1. B.) 2p3 c) 3p5.

III. Which atoms are indicated by the following configurations?

a. **(He) 2s1**
b. **(NE) 3s*2 3p3**
c. **(AR) 4s2 3d1**

Answer:

(i) (a) H^-

Configuration of H = $1s^1$

So, configuration of H^- = $1s^2$

(b) Na^+

Configuration of Na = $1s^2, 2s^2, 2p^6, 3s^1$

So, configuration of Na^+ = $1s^2, 2s^2, 2p^6$

(c) O^{2-}

Configuration of O = $1s^2, 2s^2, 2p^4$

So, configuration of O^-2 = $1s^2, 2s^2, 2p^6$

(d) F^-

Configuration of F = $1s^2, 2s^2, 2p^5$

So, configuration of F^- = $1s^2, 2s^2, 2p^6$

To obtain atomic number of an element fill the orbitals in order of their increasing energies up to the given outer orbital configuration.

(ii) (a) $1s^2, 2s^2, 2p^6, 3s^1$

Number of electrons = 2 + 2 + 6 + 1 = 11

Hence, atomic number = 11

(b) $1s^2, 2s^2, 2p^3$

Number of electrons = 2 + 2 + 3 = 7

Atomic number = 7

(c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

Number of electrons = 2 + 2 + 6 + 2 + 5 = 17

Atomic number = 17

(iii)

(a) $[He]1s^1$, it represents Li (Lithium)

(b) $[Ne]3s^2, 3p^3$, it represents P (phosphorus)

(c) $[Ar]\ 4s^2, 3d^1$, it represents Sc (scandium)

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

Answer:

For g-orbital, l = 4,

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

:For l = 4, minimum value of n = 5.

Question :25 An electron is one of the 3d orbitals. Give the possible value of n, l, and m1 for this electron.

Answer:

For the 3d orbital:

Principal quantum number (n) = 3

Azimuthal quantum number (l) = 2

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

Question :26 An atom of an element contains 29 electrons and 35 neutrons. Describe

I. The number of protons

II. The electronic configuration of elements.

Answer:

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

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

ii.) The electronic configuration of the atom is

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

Question :27 Give the number of electrons in the species H2*+, H2 and O2*-

Answer:

H_2^+:

Number of electrons present in hydrogen molecule (H2) = 1 + 1 = 2

: Number of electrons in $H2^+$ = 2 – 1 = 1

H_2:

Number of electrons in H2 = 1 + 1 = 2

O_2^-:

Number of electrons present in O2 = 8 + 8 = 16

: Number of electrons present in $O2^-$ = 16 + 1 = 17.

Question :28

I. **An atomic orbital has n = 3. What are the possible values of n and m1**
II. **List the quantum numbers of electrons for 3d orbitals.**
III. **Which of the following orbitals are possible?**
1p, 2s, 2p and 3f.

Answer:

i.) The possible values of l and m1 are :

l	m_l
0	0
1	-1, 0. +1
2	-2, -1, 0, +1, +2

ii.) The quantum numbers (m1 and l) of electrons for 3d orbital are
l= 2, m1= −2, −1, 0, +1, +2.

iii.) 2s,2p orbitals are possible.

Question :29 Using s, p, d notations, describe the orbital with the following quantum numbers.

a. **N = 1, l = 0**
b. **n = 3, l = 1**
c. **n = 4, l = 2**
d. **n = 4, l = 3**

Answer:

a.) n = 1, l = 0. 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.

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

a. **n = 0, l = 0. m = 0, m*s = +1/2**
b. **n = 1, l = 0. m*l = 0, m*s = -1/2**
c. **n = 1, l = 1. m*l = 0, m*s = +1/2**

d. n = 2, l = 1. m*l = 0, m*s = -1/2

e. n = 3, l = 3. m*l = -3, m*s = +1/2

f. n = 3, l = 1. m*l = 0, m*s = +1/2

Answer:

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

b.) The given set of quantum number is possible.

c.) The given set of quantum number is not possible.

For a given value of n, 'l' can have values from zero to (n- 1) . for n = 1, l = 0 and nor 1.

d.) The given set of quantum number is possible.

e.) The given set of quantum number is not possible. For n = 3,

l = 0 to (3 – 1)

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

f.) The given set of quantum number is possible.

Question :31 How many electrons in an atom may have the following quantum numbers?

a. n = 4, m*s = -1/2

b. n = 3, l = 0

Answer:

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

: For n = 4

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

= 32

The given element has a fully filled orbital as $1s^2 2s^2\ 2p^6 3s^2\ 3p^6 4s^2 3d^{10}$

Hence, all the electrons are paired.

: Number of electrons = 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.

Question :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 orbitals.

Answer:

We know that for the angular momentum of an electron ⇒

mvr = nh / 2 π -----⟶ (i)

Also, By de broglie equation , we know that

λ = h/mv

mv = h / λ ----⟶(ii)

Now putting the value of (ii) to (i)

2 π r = nh(λ/h)

2πr = nλ

Since '2πr' represents here, circumference of the Bohr orbit (r)

Therefore the given sentence of the question is proved.

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

Answer:

For He^+ ion, the wave number associated with the Balmer transition, n = 4 to n = 2 is given by:

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

Where,

n1 = 2

n2 = 4

Z = atomic number of helium

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

= 4R(3/16)

$\bar{\nu} = 3R / 4$

$1 / \lambda = 4 / 3R$

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

$R(1)^2(1 / n_1^2 - 1/ n_2^2) = 3R / 4$

$(1 / n1^2 - 1 / n_2^2) = 3/ 4$

By hit and trail method, the equality given by equation (1) is true only when n1 = 1 and n2 = 2

The transition for n2 = 2 to n1 = 1 in hydrogen spectrum would have the same wavelength as Balmer transition n = 4 to n = 2 of He+ spectrum.

Question :34 Calculate the energy required for the process

He+(g) → He2+(g) + e-

The ionisation energy for the H atom in the ground state is 2.18 x 10*-18 J/ atom.

Answer:
According to Bohr's theory , we know,
energy of electron in unielectron atomic system ,
$E_n = -2\pi^2 m Z^2 e^4 / n^2 h^2$
For H - atom,
Ionisation energy is the energy required to remove a electron from ground state to infinity.
$I.E = 0 - (-2\pi^2 m 1^2 e^4 / 1^2 h^2)$ [because n = 1, Z = 1]
$= 2\pi^2 m e^4 / h^2$
After putting $\pi = 3.14$
$m = 9.1 \times 10^{-31}$ Kg
$e = 1.602 \times 10^{-19}$ C
$h = 6.626 \times 10^{-34}$ Js
Then, we get $I.E = 2.18 \times 10^{-18}$ J/atom
Now, for He^+ atom,
$IE = 0 - \{-2\pi\ m(2)^2 e^4 / 1^2 h^2\}$ [because n = 1, Z = 2]
$I.E = 4 \times \{2\pi^2 m e^4 / h^2\} = 4 \times 2.18 \times 10^{-18}$ J/atom
$= 8.72 \times 10^{-18}$ J/atom
Hence, the energy required for the given process is 8.72×10^{-18} J/atom

Question :35 If the diameter of a carbon atom is 0.5 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:
1m = 100 cm
1 cm = 10^{-2}m
Length of the scale = 20 cm
$= 20 \times 10^{-2}$ m
Diameter of a carbon atom = 0.15 nm
$= 0.15 \times 10^{-9}$m
One carbon atom occupies 0.15×10^{-9}m
Number of carbon atoms that can be placed in a straight line
$= 20 \times 10^{-2}\text{m} / 0.15 \times 10^{-9}\text{m}$
$= 133.33 \times 10^{7}$
1.33×10^{9}

Question :36 2 x 10*8 atoms of carbon are arranged side by side. Calculate the radius of carbon atom if length of this arrangement is 2.4 CM.

Answer:

Light of the given arrangement = 2.4 cm

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

: Diameter of carbon atom

= 2.4×10^{-2}m / 2×10^{8}m

= 1.2×10^{-10}m

Radius of carbon atom = Diameter / 2

= 1.2×10^{-10}m / 2

= 6.0×10^{-10}m

Question :37 The diameter of zinc atom is 2.6 angstrom. Calculate

a. Radius of zinc atom in pm

b. Number of atoms present in a length of 1.6 CM if the zinc atoms are arranged side by side lengthwise.

Answer:

a.) Radius of zinc = Diameter / 2

2.6 A^{0} / 2

= 1.3×10^{-10}m

= 130×10^{-12}m

= 130pm

b.) Length of the arrangement = 1.6 cm

= 1.6×10^{-2}m

Diameter of zinc atoms present in the arrangement

= 1.6×10^{-2}m / 2.6×10^{-10}m

= 0.6153×10^{8}m

= 6.153×10^{7}

Question :38 A certain particle carries 2.5 x 10*-16 C of static electric charge. Calculate the number of electrons present in it.

Answer:

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

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

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

= 1(2.5×10^{-16}C / 1.6022×10^{-19}C)

= $1.560 \times 10^3 C$
= 1560 C.

Question :39 In Milikan's experiment, static electric charge on the Lil drops has been obtained by shining X – days. If the static electric charge on the oil drops is -1.282 x 10*-18 C, calculate the number of electrons present on it.

Answer:
Charge on the oil drop = $1.282 \times 10^{-18}C$
Charge on an electron = $1.6022 \times 10^{-19}C$
: Number of electrons present on the oil drop
= $1.282 \times 10^{-18}C / 1.6022 \times 10^{-19}C$
= 0.8001 x 10
= 8.0.

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

Answer:
A thin foil of lighter atoms will not give the same results as given with the foil of heavier atoms. Lighter atoms would be able to carry very little positive charge. Hence, they will not cause enough deflection of a- particles.

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

Answer:
The general convention of representing an element along with its atomic mass and atomic number(Z) is $^A{}_Z$ X.
Hence, $^{79}{}_{35}Br$ is acceptable but $_{79}{}^{35}Br$ is not acceptable.
^{79}Br can be written but ^{35}Br cannot be written because the atomic number of an element is constant, but the atomic mass of an element depends upon the relative abundance of its isotopes. Hence, it is necessary to mention the atomic mass of an element.

Question :42 An element with mass number 81 contains 31.7% more neutrons as

compared to protons. Assign the atomic symbol

Answer:

We know that mass number of the element,

A = p + n = 81

Let the number of protons, p = x

Then, number of neutrons,

n = x + 31.7 / 100x = 1.317 x

(As number of neutrons are 31.7% more than the protons.)

Hence, from Eq. (i)

x + 1.317 x = 81

x = 81 / 2.317 = 34.958 = 35 (approx)

Therefore, number of protons = 35 and the symbol is $_{35}^{81}Br$

(Number of protons = Atomic number)

Question :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 ion.

Answer:

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

Then,

Number of neutrons present

= x + 11.1% of x

= x + 0.111 x

= 1.111 x

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

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

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

Therefore 37 = 1.111x + x -1

Or 2.111 x = 38

Or x = 18

Therefore no of protons = atomic no = x -1 = 18 – 1 = 17

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

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

Answer:
Let the number of electrons in ion = x
Number of neutrons = x + 0.304x = 1.304x
Number of electrons in neutral atom = x + 3
Number of protons in neutral atom = x + 3
Mass number = 56
(x+3) + (1.304x) = 56
X = 23
Number of protons = x+3 = 26
Symbol of ions = $_{26}^{56}$ Fe3+

Question :45 Arrange the following type of radiation 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 spacee**
e. **X – rays**

Answer:
The increase in order of frequency is as follows:
Radiation from FM radio < radiation from micro wave oven < X – rays < cosmic rays
The increasing order of wavelength is as follows:
Cosmic rays < X – rays < radiation from microwave ovens < amber light < radiation of FM radio.

Question :46 Nitrogen lase produces a radiation at a wavelength of 337.1 NM. If the number of protons emitted is 5.6 x 10*24, calculate the power of this laser.

Answer:
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):
= $(5.6 \times 10^{24})(6.626 \times 10^{-34} Js)(3 \times 10^{8}\ m / s) / 337.1 \times 10^{-9} m$
= 3.33 × 106 J

Hence, the power of the laser is 3.33 × 106 J.

Question :47 Neon gas is generally used in the sign boards. If it emits strongly at 616nm, calculate

a. **The frequency of emission**
b. **Distance traveled by this radiation in 30 s**
c. **Energy of quantum**
d. **Number of quanta present if it produces 2J of energy.**

Answer:

λ = 616nm or = 616x10-9m

a) Frequency , $v = c/\lambda = 3.8 x 10^8 / 616 \times 10^{-9} = 4.87 x 10^{14} s^{-1}$

b) Velocity of the radiation = 3×10^8 m/s

Therefore distance travelled in 30s = $30 \times 3 \times 10^8 = 9.0 \times 10^9$ m

c) $E = hv = hc / \lambda = (6.626 \times 10^{-34})(3.0 \times 10^8) / (616 \times 10^{-9}) = 32.27 \times 10^{-20}$ J

d) No of quanta in 2J of energy = $2/32.27 \times 10^{-28} = 6.2 \times 10^{18}$

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

Answer:

From the expression of energy of one photon, $E = hc / \lambda$

Where,

λ = wavelength of radiation

h = Plank's constant

c = velocity of radiation

Substituting the values in the given expression of E:

Energy of 1 photon = $hv = hc/\lambda = (6.626 \times 10^{-34})(3 \times 10^8) / 600 \times 10^{-9} = 3.313 \times 10^{-19}$ J

Total energy received = $3.15 x 10^{-18}$ J

Therefore no of photons received = $3.15 \times 10^{-18} / 3.313 \times 10^{-19} = 9.51$

Question :49 Lifetime of the molecules in the excited states are often measured by

using pulsed radiations source of duration nearly in the bank second range. If the radiation source has the duration 2ns and the number of photons emitted spring the pulse source is 2.5 x 10*15 ,calculate the energy of the source.

Answer:

Frequency of radiation (v),

Frequency = $1/2.x10^{-9} = 0.5\ x10^{9}$ / sec

Energy = Nhv

Where,

N = number of photons emitted

h = Planck's constant

v = frequency of radiation

= $(2.5x10^{5})(6.626\ x\ 10^{-34})(0.5\ x10^{9})$

= $8.28\ x10^{-10}$ J

Question :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 know,

Frequency = speed of light /wavelength

For wavelength (λ_1) = 589 nm

Frequency (v_1) = 3 × 10^8/589 × 10^-9

= 5.093×10^{14} Hz

For wavelength (λ_2) = 589.6 nm

Frequency (v_2) = 3 × 10^8/589.6 × 10^-9

=5.088×10^{14} Hz

Now, energy difference (ΔE) = $E_1 - E_2$

$\Delta E = E_1 - E_2 = hv_1 - hv_1$

= $h(v_1 - v_2)$

= 6.626 × 10^-34 × (5.093 - 5.088) × 10^{14}

= 6.626 × 0.005 × 10^(-34 + 14) J

= 3.31 × 10^-22 J

Question :51 The work function for caesium atom is 1.9 eV. Calculate

a. **The threshold wavelength**
b. **The threshold frequency of the radiation.**

If the caesium element is irradiated with a wavelength 500nm, calculate the kinetic energy and the velocity of the ejected photoelectric.

Answer:

A) Work function of caesium (WO) = hv_0

Therefore $v_0 = W_0 /h = 1.9 \times 1.602 \times 10^{-19} / 6.626 \times 10^{-34}$

$= 4.59 \times 10^{14}$ /sec

B) $\lambda_0 = c/v_0 = 3 \times 10^8 / 4.59 \times 10^{14} = 6.54v\times 10^{-7}$ m

C) K.E of ejected electron = $h(v - v_0) = hc(1/\lambda - 1/\lambda_0)$

$=(6.626 \times 3 \times 10^{-26}) (1/500 \times 10^{-9} - 1/654 \times 10^{-9})$

$=(6.626 \times 3 \times 10^{-26}) / 10^{-9} (154 / 500 \times 654)$

$= 9.36 \times 10^{-20}$ J

K.E = $1mv^2 /2 = 9.36 \times 10^{-20}$ J

$=9.1\times 10^{-31}/2 = 9.36 \times 10^{-20}$ J

Or

$V^2 = 20.55 \times 10^{10}\ m^2\ s^{-2}$

Or

$v = 4.53 \times 10^5$ m / s

Question :52 Following results are observed when sodium metal is irradiated with different wavelength. Calculate

a. **Threshold wavelength**
b. **Planck's constant.**

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

Answer

Let the threshold wavelength be λ_0 nm or $\lambda_0 \times 10^{-9}$m.

$h(v-v_0) = \frac{1}{2} mv^2$

$hc(1/ \lambda - 1/ \lambda_0) = \frac{1}{2} mv^2$ ……..(a)

$hc(1 / 500 \times 10^9 - 1/ \lambda_0 \times 10^{-9}m) = \frac{1}{2} m (2.55 \times 10^{-5} \times 10^{-2}\ m / s)^2$

$hc / 10^{-9}m([1 / 500 - 1/ \lambda_0] = \frac{1}{2} m(2.55 \times 10^3 m /s)^2$

Similarly,

$hc / 10^{-9}m\ [1/450 – 1/\lambda_0] = 1/2m(3.45 \times 10^3 m / s)^2$ …………(2)

$hc / 10^{-9} m[1/400 – 1/\lambda_0] = 1/2m(5.35 \times 10^3 m / s)^2$ …………(3)

Dividing equation (3) by equation (1):

$[\lambda_0 – 400 / 400\ \lambda_0 / \lambda_0 – 500 / 500\ \lambda_0] = (5.35 \times 10^{-3}\ m / s)^2\ (2.55 \times 10^{-3} m / s)^2$

$5\ \lambda_0 – 2000 / 4\ \lambda_0 – 2000 = 4.40177$

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

$\lambda_0 = 539.8$ nm

$\lambda_0 = 540$ nm

Threshold wavelength = 540 n

Planck's constant:

Substituting this value in equation (a),we get

= h x (3x108)/10-9 [1/400-1/540]

= [(9.11x10-31)(5.20x106)2] / 2

$=6.66 \times 10^{-34}$ Js

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

Answer:

E = hc/λ

Here, h = 6.626 × 10^-34 Js

c = 3 × 10^8 m/s

λ = 256.7 nm = 2.567 × 10^-7 m

E = 6.626 × 10^-34 Js × 3 × 10^8 m/s/2.567 × 10^-7 m

= 7.74 × 10^-19 J/1.602 × 10^-19 J/eV

= 4.83 eV

E = 4.83 eV

Now, potential gives the kinetic energy to electron.

$eV_0 = ½\ mv^2_{max.} = K.E.$

K.E = e × 0.35V = 0.35 eV

So, work function = Energy - K.E

= 4.83 eV - 0.35 eV

= 4.48 eV

Question :54 If the photon of the wavelength 150pm strikes an atom and one of its

inner bound electrons is ejected out with a velocity of 1.5 x 10*7 m/s. Calculate the energy with which it is bound to the nucleus.

Answer:

We know, $E = hv = hc/\lambda$

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

$c = 3 \times 10^{8}$ m/s

$\lambda = 150$ pm $= 1.5 \times 10^{-10}$ m

$E = 6.626 \times 10^{-34} \times 3 \times 10^{8}/1.5 \times 10^{-10}$ J

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

K.E of ejected electron $= 1/2\ mv^2$

$= 1/2 \times 9.11 \times 10^{-31} \times (1.5 \times 10^{7})^2$ J

$= 1.015 \times 10^{-16}$ J

Now, energy with which the electron was bound to the nucleus = work function for the metal = w_0

$W_0 = hv - 1/2\ mv^2$

= Energy - K.E of ejected electron

$= 13.25 \times 10^{-16}$ J $- 1.025 \times 10^{-16}$ J

$= 12.225 \times 10^{-16}$ J

Hence, energy with which the electron was bound to the nucleus $= 12.225 \times 10^{-16}$ J

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

Answer:

According to question,

$v = 3.29 \times 10^{15}\ (1/3^2 - 1/n^2)$Hz

We know,

Frequency = speed of light/wavelength

Given , here

Wavelength = 1285 nm $= 1.285 \times 10^{-6}$ m

Speed of light $= 3 \times 10^{8}$ m/s

Frequency (v) $= 3 \times 10^{8}/1.285 \times 10^{-6}$ Hz

$3.29 \times 10^{15}\ (1/3^2 - 1/n^2)$ Hz $= 3 \times 10^{8}/1.285 \times 10^{-6}$ Hz

$(1/9 - 1/n^2) = 3 \times 10^{8}/1.285 \times 10^{-6} \times 3.29 \times 10^{15}$

$0.1111 - 1/n^2 = 0.0709$

$1/n^2 = 0.1111 - 0.0709 = 0.0402 \approx 0.04$

$1/n^2 = 1/25$
$1/n^2 = 1/5^2$
n = 5
The electrons jumps from n = 5 to n = 3 .e.g., the transition occurs in paschen series and lies infrared region.
Hence, the radiation 1285 nm lies in the infrared region.

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

Answer:
We know, from Bohr's theory ,
Radius of nth orbit H like species,
$R_n = 52.9(n^2)$pm / Z
Where, n is nth orbit and Z is atomic number .
Given, r1 = 1.3225 nm = 1322.5 pm
Now,
1322.5pm = $52.9(n_1^2)$pm / Z ……..(1)
Again, r2 = 211.6 pm
2116. pm = $52.9(n_2^2)$ pm / Z . …..(2)
Now, divide equations (1) and (2)
$r_1 / r_2 = 1322.5 / 211.6 = n_1^2 / n_2^2$
$n_1^2 / n_2^2 = 6.25 = (2.5)^2$
$n1 / n_2 = 2.5$
It is possible when , n1 = 5 and n2 = 2 so, the transition is from 5th orbit to 2nd orbit .
It belongs to Balmer's series.
Now, use formula,
$1/\lambda = 1.09677 \times 10^7 (1/n1^2 - 1/n2^2)$
$= 1.09677 \times 10^7(1/2^2 - 1/5^2)$
$= 1.09677 \times 10^7 \times 21/100$
$= 2.303 \times 10^6$ m^-1
$\lambda = 1/2.303 \times 10^6$ m
$= 434 \times 10^{-9}$ m = 434 nm
Hence, wavelength = 434 nm
It belongs to visible region

Question :57 Dual behaviour of matter proposed by de Broglie led to the discovery of electrons 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 x 10*6 m/s, calculate de Broglie wavelength associated with this electron.

Answer:

From the de Broglie equation,

$\lambda = h / mv$

$\lambda = 6.626 \times 10^{-34}$ Js / $(9.10939 \times 10^{-31}$kg$)(1.6 \times 10^{8}$m / s)

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

$\lambda = 455$ pm

De – Broglie's wavelength associated with the electron is 455 pm.

Question :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 800pm. Calculate the characteristics velocity associated with the neutrons.

Answer:

When the mass of the neutron is 1.675 x10-27 kg

We know that $\lambda = h/mv$ (let it be 1) and

$v = h/m\lambda$ (let it be 2)

On Substituting of both 1 and 2, we get the following

$= 6.626 \times 10^{-34} / (1.675 \times 10^{-27})(800 \times 10^{-12})$

$= 4.94 \times 10^{4}$ m/sec.

Hence,= 4.94×10^{4} m/sec is the characteristic velocity associated with the neutron.

Question :59 If the velocity of the electron in Bhor's first orbit is 2.19 x 10*6 m/s, calculate the de Broglie wavelength associated with it.

Answer:

$v = 2.19 * 10^{6}$ m/s

$\lambda = h / p$

h = Planck's constant = $6.636 * 10^{-24}$ units

m = mass of an electron = $9.11 * 10^{-21}$ kg

Bohr's first orbit : n = 1, K shell in an atom... like Hydrogen atom..

$\lambda = 6.636 * 10^{-24} / [\ 9.11 * 10^{-31} * 2.19 * 10^{6}\ ]$ meters

$= 0.3326 * 10^{-9}$ meters

Question :60 The velocity associated with a proton moving in a potential difference if 1000 V is 4.37 x 10*5 m/s .If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

Answer:

According to the de Broglie's equation,

$\lambda = h / mv$

Substituting the values in the expression,

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

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

Question :61 If the position of the electron is measured within an accuracy of $\pm$ 0.002nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is h/4 π_m x 0.05nm, is there any problem in defining thus value.

Answer:

From Heisenberg's uncertainty principle,

$\Delta x \times \Delta p = h / 4\pi$

$\Delta p = h / 4 \pi \Delta x$

Where,

Δx = uncertainty in position of the electron

Δp = uncertainty in momentum of the electron

Substituting the values in the expression of Δp

$\Delta p = 1 / 0.002$nm x 6.626×10^{-34} Js / 4 x 3.14

= 1 / 2 x 10^{-12}m x 6.626×10^{-34} Js / 4 x 3.14

= 2.637×10^{-23} Js / m

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

Uncertainty in the momentum of the electron = 2.637×10^{-23} kg m / s

Actual momentum = h / 4 π_m x 0.05nm

= 6.626×10^{-31}Js / 4 x 3.14 x 5.0 x 10^{-11}m

= 1.055×10^{-24}kg m/s

Since the magnitude of the actual momentum is smaller than the uncertainty, the value can not be defined

Question :62 The quantum numbers of six electrons are given below. Arrange them

in order of increasing energies. If any of these combinations has/have the same energy lists:

1. **n = 4, l= 2, m*l = -2, m*s = -1/2**
2. **n = 3, l = 2, m*l = 1, m*s = +1/2**
3. **n = 4, l = 1, m*l = 0, m*s = +1/2**
4. **n = 3, l = 2, m*l = -2, m*s = -1/2**
5. **n = 3, l = 1, m*l = -1, m*s = +1/2**
6. **n = 4, l = 1, m*l = 0, m*s = +1/2**

Answer:

Quantum number – they are the index numbers which gives the complete address of an electron

Here n =principal quantum number

l = azimuthal quantum number

ms = Spin 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 orbitals respectively.

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

Question :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 electrons experiences the lowest effective nuclear charge?

Answer:

Bromine (Atomic No. 35) = 1s2 2s2 2p6 3s2 2p6 3d10 4s2 4p5

2p orbital has 6 electrons.

3p orbital has 6 electrons and 4p orbital has 5 electrons.

4p electron experience lowest effective nuclear charge because of the maximum of screening or shielding effect.

Question :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**

Answer:

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

i.) The electron present in 2s orbital will experience greater nuclear charge than the electron in 3s orbital.

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

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

Question :65 The unpaired electrons in AL and Sister are present in 3p orbitals. Which electrons will experience more effective nuclear charge from the nucleus?

Answer:

Nuclear charge is defined as the net positive charge experienced by an electron in a multi – electron atom. The higher the atomic number, the higher is the nuclear charge. Silicon has 14 protons while aluminium has 13 protons. Hence, silicon has large nuclear charge of (+14) than aluminium, which has a nuclear charge than aluminium.

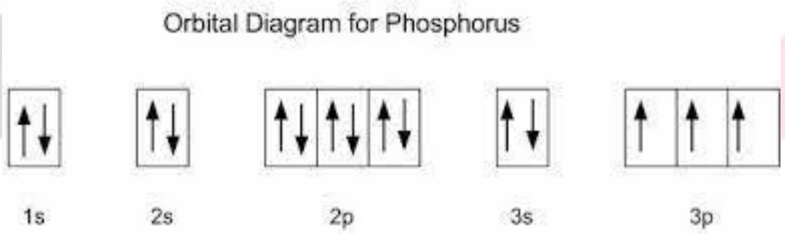
Question :66 Indicate the number of unpaired electrons in :

- **a. P**
- **b. Si**
- **c. Cr**
- **d. Fe**
- **e. Kr**

Answer:

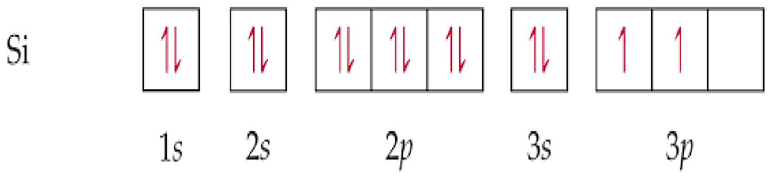
(a) Phosphorus (P): 1s2 2s2 2p6 3s2 3p3

No of unpaired electron = 3



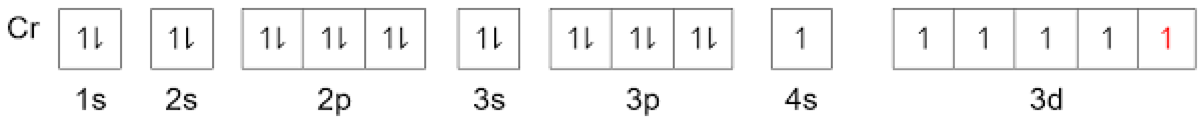
(b) Silicon (Si): 1s2 2s2 2p6 3s2 3p2

No of unpaired electron = 2 (since p orbital can have maximum 6 electron)



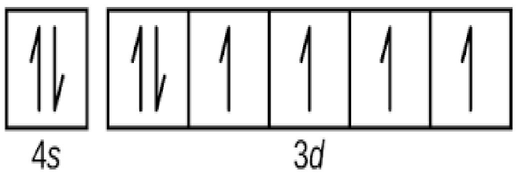
(c) Chromium (Cr): 1s2 2s2 2p6 3s2 3p6 4s1 3d5

No of unpaired electrons = 6 (since 1 electron is to be added to 4s & 5 electron to be added to 3d orbital.

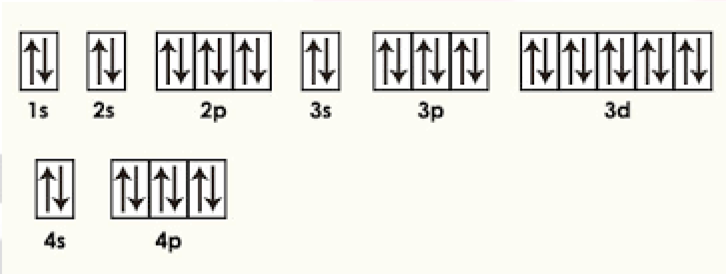


(d) Iron (Fe): 1s2 2s2 2p6 3s2 3p6 4s2 3d6

No of unpaired electrons = 4



(e) Krypton (Kr): 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6



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

Question :67

a. **How many sub – shells are associated with n = 4?**
b. **How many electrons will be present in the sub she'll having Ms value of -1/2 for n = 4?**

Answer:

a.) n = 4

For a given value of ‘n’, ‘l’ can have values from zero to (n-1)

:l = 0, 1, 2, 3

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

For n = 4

Number of orbitals = 16

If each orbital is taken fully, then it will have 1 electron with m_s value of (-1/2)

: Number of electrons with m_s values of (-1/2) is 16.