Chapter 2: Atomic Structure

Examples

1. Calculate the number of protons, neutrons and electrons in $^{80}_{35}Br$.

Answer: Here $^{80}_{35}Br$ is given that where $Z = 35, A = 80$

i.e. species is neutral

NO. of protons = no. of electrons = $Z = 35$

No. of neutrons = $80-35 = 45$ i.e. Mass no. – No. of protons

2. The number of electrons, protons and neutrons in a species are equal to 18, 16, 16 respectively. Assign the proper symbol to the species

Answer: It is given that atomic number is equal to no. of protons i.e. 16 and at that it is sulphur (S)

As we know that Atomic mass = No. of protons + No. of neutrons

$$= 16+16 = 32$$

Here, species is not neutral because protons is not equal to electrons

Thus, it is negatively charged which is equal to excess electrons = $18-16 = 2$

So, it gives us symbol $^{32}_{16}S^{2-}$

3. The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of $1.368\ \text{kHz}$ Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?

Answer: Here, wave length = λ and it is equal to $\frac{c}{\nu}$

i.e. c is speed of electromagnetic radiation in vacuum and ν is the frequency

Substitute values, we get

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

$$= \frac{3.00\times10^{8}\ \text{ms}^{-1}}{1368\ \text{kHz}}$$

$$= \frac{3.00\times10^{8}\ \text{m s}^{-1}}{1368\times10^{3}\ \text{s}^{-1}}$$

$$= 219.3\ \text{m}$$

Hence, It is a characteristic radioware wave length

4. The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz)$(1\,\text{nm} = 10^{-9}\,\text{m})$.

Answer: Frequency of violet line

$$\nu = \frac{c}{\lambda}$$
$$= \frac{3.00\times10^{8}\ \text{ms}^{-1}}{400\times10^{-9}\ \text{m}}$$
$$= 7.50\times10^{14}\ \text{Hz}$$

Frequency of red light

$$\nu = \frac{c}{\lambda}$$
$$= \frac{3.00\times10^{8}\ \text{ms}^{-1}}{750\times10^{-9}\ \text{m}}$$
$$= 4.00\times10^{14}\ \text{Hz}$$

Hence, range of visible spectrum is from 4.00×10^{14} Hz to 7.50×10^{14} Hz

5. Calculate (a) wavenumber

(b) Frequency of yellow radiation having wavelength 5800Å

Answer: (a) Here, wavenumber denotes as $(\bar{\nu})$

$$\lambda = 5800\text{Å}$$
$$= 5800\times10^{-8}\ \text{cm} = 5800\times10^{-10}\ \text{m}$$
$$\bar{\nu} = \frac{1}{\lambda}$$
$$= \frac{1}{5800\times10^{-10}\ \text{m}}$$
$$= 1.724\times10^{6}\ \text{m}^{-1} = 1.724\times10^{4}\ \text{cm}^{-1}$$

(b) Here, frequency denotes as ν

$$\bar{\nu} = \frac{c}{\lambda}$$
$$= \frac{3\times10^{8}\ \text{m s}^{-1}}{5800\times10^{-10}\ \text{m}}$$
$$= 5.172\times10^{14}\ \text{s}^{-1}$$

6. Calculate energy of one mole of photons of radiation whose frequency is 5×10^{14} Hz.

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Answer: Given that

$$h = 6.626\times10^{-34}\ \text{J s}$$
$$\nu = 5\times10^{14}\ \text{s}^{-1}$$

As energy denotes by E and it is given by expression $E = h\nu$, So

$$E = \left(6.626\times10^{-34}\ \text{J s}\right)\times\left(5\times10^{14}\ \text{s}^{-1}\right)$$
$$= 3.313\times10^{-19}\ \text{J}$$

Now, energy of one mole of photons

$$= \left(3.313\times10^{-19}\ \text{J}\right)\times\left(6.022\times10^{23}\ \text{mol}^{-1}\right)$$
$$= 199.51\ \text{kJ mol}^{-1}$$

7. A 100 watt bulb emits monochromatic light of wavelength 400nm. Calculate the number of photons emitted per second by the bulb.

Answer: Here, power of bulb = 100 watt = 100 J s^{-1}

Energy of one photon $E = h\nu$

$$= \frac{hc}{\lambda}$$
$$= \frac{6.626\times10^{-34}\ \text{J s}\times3\times10^{8}\ \text{m s}^{-1}}{400\times10^{-9}\ \text{m}}$$
$$= 4.969\times10^{-19}\ \text{J}$$

No. of photons emitted $= \dfrac{100\ \text{J s}^{-1}}{4.969\times10^{-19}\ \text{J}}$

$$= 2.012\times10^{20}\ \text{s}^{-1}$$

8. When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of $1.68\times10^{5}\ \text{J mol}^{-1}$. What is the minimum energy needed to remove an electron from sodium? What is the maximum wavelength that will cause a photoelectron to be emitted?

Answer: Energy (E) of 300 nm photon given as $hn = \dfrac{hc}{\lambda}$

$$= \frac{6.626\times10^{-34}\ \text{J s}\times3.0\times10^{8}\ \text{m s}^{-1}}{300\times10^{-9}\ \text{m}}$$
$$= 6.626\times10^{-19}\ \text{J}$$

Energy of one mole of photons = $6.626\times10^{-19}\ \text{J}\times6.022\times10^{23}\ \text{mol}^{-1}$

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$= 3.99\times10^{5}\ \text{J mol}^{-1}$

So, minimum energy needed to remove one mole of electrons from sodium

$= (3.99\text{-}1.68)10^{5}\ \text{J mol}^{-1}$

$= 2.31\times10^{5}\ \text{J mol}^{-1}$

Now, minimum energy for one electron

$$= \frac{2.31\times10^{5}\ \text{J mol}^{-1}}{6.022\times10^{23}\ \text{electrons mol}^{-1}}$$

$= 3.84\times10^{-19}\ \text{J}$

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

$$= \frac{6.626\times10^{-34}\ \text{J s}\times3.0\times10^{8}\ \text{m s}^{-1}}{3.84\times10^{-19}\ \text{J}}$$

$= 517\ \text{nm}$

9. The threshold frequency v_0 for a metal is $7.0\times10^{14}\ s^{-1}$. Calculate the kinetic energy of an electron emitted when radiation of frequency $v = 1.0\times10^{15}\ s^{-1}$ hits the metal.

Answer: Here kinetic energy $= \frac{1}{2}m_c v^2$

$= h\left(v\text{-}v_0\right)$

$= \left(6.626\times10^{-34}\ \text{J s}\right)\left(1.0\times10^{15}\ \text{s}^{-1}\text{-}7.0\times10^{14}\ \text{s}^{-1}\right)$

$= \left(6.626\times10^{-34}\ \text{J s}\right)(10.0\times10^{14}\ \text{s}^{-1}\text{-}7.0\times10^{14}\ \text{s}^{-1})$

$= \left(6.626\times10^{-34}\ \text{J s}\right)\times\left(3.0\times10^{14}\ \text{s}^{-1}\right)$

$= 1.988\times10^{-19}\ \text{J}$

10. What are the frequency and wavelength of a photon emitted during a transition from n=5 state to the n=2 state in the hydrogen atom.

Answer: As $n_1 = 5, n_2 = 2$ here this transition gives rise to spectral line in the visible region. So,

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

$= -4.58\times10^{-19}\ \text{J}$

Thus, it’s an emission energy

Frequency of photon $\nu = \frac{\Delta E}{h}$

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$$= \frac{4.58\times10^{-19}\ J}{6.626\times10^{-34}\ J\ s}$$

$= 6.91\times10^{14}\ Hz$

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

$$= \frac{3.0\times10^{8}\ m\ s^{-1}}{6.91\times10^{14}\ Hz}$$

$= 434\ nm$

11. Calculate the energy associated with the first orbit of He^{+}. What is the radius of this orbit?

Answer: $E_n = -\frac{(2.18\times10^{-18}\ J)Z^2}{n^2}\ atom^{-1}$

So, For He^{+}, $n = 1, Z = 2$

i.e. $E_1 = -\frac{(2.18\times10^{-18}\ J)(2^2)}{1^2}$

$= -8.72\times10^{-18}\ J$

So, radius of orbit given as

$$r_n = \frac{(0.0529\ nm)n^2}{Z}$$

As $n = 1, Z = 2$

$$r_n = \frac{(0.0529\ nm)1^2}{2}$$

$= 0.02645\ nm$

12. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of $10\ m\ s^{-1}$.

Answer: According to brogile equation i.e. $\lambda = \frac{h}{m\nu} = \frac{h}{p}$

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

$$= \frac{(6.626\times10^{-34} Js)}{(0.1\ kg)(10\ m\ s^{-1})}$$

$= 6.626\times10^{-34}\ m$ $\quad (J = kg\ m^2\ s^{-2})$

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

Answer: As, K.E. = $\frac{1}{2}mv^2$

$$v=\left(\frac{2\ \text{K.E.}}{m}\right)^{\frac{1}{2}}$$

$$=\left(\frac{2\times3.0\times10^{-25}\ \text{kg m}^2\ \text{s}^{-2}}{9.1\times10^{-31}\ \text{kg}}\right)^{\frac{1}{2}}$$

$$=812\ \text{m s}^{-1}$$

$$\lambda=\frac{h}{mv}=\frac{6.626\times10^{-34}\text{Js}}{(9.1\times10^{-31}\ \text{kg})(812\ \text{m s}^{-1})}$$

$$=8967\times10^{-10}\ \text{m}$$

$$=896.7\ \text{nm}$$

14. Calculate the mass of a photon with wavelength 3.6 Å.

Answer: Here, $\lambda = 3.6$ Å

$$=3.6\times10^{-10}\ \text{m}$$

So, velocity of photon = velocity of light

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

$$=\frac{6.626\times10^{-34}\text{Js}}{(3.6\times10^{-10}\ \text{m})(3\times10^{8}\ \text{ms}^{-1})}$$

$$=6.135\times10^{-29}\ \text{kg}$$

15. A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 Å What is the uncertainty involved in the measurement of its velocity?

Answer: As $\Delta x\ \Delta p=\frac{h}{4\pi}$ or $\Delta x\ m\ \Delta v=\frac{h}{4\pi}$

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$$\Delta v = \frac{h}{4\pi\ \Delta x\ m}$$

$$\Delta v = \frac{6.626\times10^{-34} Js}{4\times3.14\times0.1\times10^{-10}\ m\times9.11\times10^{-31}\ kg}$$

$$= 0.579\times10^{7}\ m\ s^{-1} \qquad \left(1\ J = 1\ kg\ m^{2}\ s^{-2}\right)$$

$$= 5.79\times10^{6}\ m\ s^{-1}$$

16. A golf ball has a mass of 40 g and a speed of $45\ m/s$ if the speed can be measured within accuracy of 2%, calculate the uncertainty in the position.

Answer: As given that uncertainty in speed = 2%

$$\Rightarrow\ 45\times\frac{2}{100} = 0.9\ m\ s^{-1}$$

$$\Rightarrow\ \Delta x = \frac{h}{4\pi m\Delta v}$$

$$= \frac{6.626\times10^{-34} Js}{4\times3.14\times40\ g\times10^{-3} kg\ g^{-1}\left(0.9\ m\ s^{-1}\right)}$$

$$= 1.46\times10^{-33}\ m$$

Approximately it is near to 10^{18} times smaller than the diameter of typical atomic nucleus.

17. What is the total number of orbitals associated with the principle quantum number n=3.

Answer: As it is given that $n = 3$

So, possible values for $l = 0, 1, 2$

Hence, there is $1\ 3s$ orbital i.e. $n = 3, l = 1, m_l = 0$

And there are $3\ 3p$ orbitals i.e. $n = 3, l = 1, m_l = -1, 0, +1$

And there are $5\ 3d$ orbitals i.e. $n = 3, l = 2, m_l = -2, -1, 0, +1+, +2$

So, total orbitals = $1+3+5 = 9$

No. of orbitals = $n^2 = 3^2 = 9$

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

(a) $n = 2, l = 1$
(b) $n = 4, l = 0$
(c) $n = 5, l = 3$
(d) $n = 3, l = 2$

Answer: Here, orbitals are

(a) n = 2, l = 1, orbital = 2p
(b) n = 4, l = 0, orbital = 4s
(c) n = 5, l = 3, orbital = 5f
(d) n = 3, l = 2, orbital = 3d

Exercise

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

Answer: (i) Here mass of one electron = 9.10939×10^{-31} kg

No. of electrons weigh = 9.10939×10^{-31} kg = 1

No. of electrons will weigh = $1\ g\left(1\times10^{-3}\ kg\right)$

$$= \frac{1}{9.10939\times10^{-31}\ kg}\times\left(1\times10^{-3}\ kg\right)$$
$$= 0.1098\times10^{-3+31}$$
$$= 0.1098\times10^{28}$$
$$= 1.098\times10^{27}$$

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

Mass of one mole of electron = $\left(6.022\times10^{23}\right)\times\left(9.10939\times10^{-31}\ kg\right) = 5.48\times10^{-7}$ kg

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

Charge on one mole of electron = $\left(1.6022\times10^{-19}C\right)\left(6.022\times10^{23}\right) = 9.65\times10^{4}C$

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 7mg of ^{14}C

(Assume that mass of a neutron $= 1.675\times10^{-27}$ kg).

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

Answer: (i) Methane represents as CH_4

No. of electrons in one mole molecules of methane = 6.023×10^{23}

$$= 6.022\times10^{23}\times10$$
$$= 6.022\times10^{24}$$

(ii) (a) As no. of atoms in one mole = 6.023×10^{23}

As one atom of ^{14}C have eight neutrons

So, no. of neutrons in 14 g of ^{14}C = $(6.023\times10^{23})\times8$ or $(6.022\times10^{23}\times8)$ neutrons

Thus, No. of neutrons in 7 mg $= \dfrac{6.022\times10^{23}\times8\times7\text{mg}}{1400\text{mg}} = 2.4092\times10^{21}$

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

So, mass of total neutrons in 7 g of ^{14}C

$$= (2.4092\times10^{21})(1.67493\times10^{-27}\ \text{kg})$$
$$= 4.0352\times10^{-6}\ \text{kg}$$

(iii) (a) One mole of NH_3 = $\{1(14)+3(1)\}$g of NH_3 = 17 g of NH_3

$= 6.022\times10^{23}$ molecules of NH_3

Total protons in one molecule of NH_3 = $\{1(7)+3(1)\}$ = 10

No. of protons in 6.023×10^{23} molecules of NH_3

$$= (6.023\times10^{23})(10)$$
$$= 6.023\times10^{24}$$

This implies 17 g of NH_3 have (6.023×10^{24}) protons

No. of protons in 34 mg of NH^3

$$= \frac{6.022\times10^{24}\times34\text{mg}}{17000\text{ mg}} = 1.2046\times10^{22}$$

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

So, Total mass of proton in 34 mg of NH_3

$$= \left(1.67493\times10^{-27}\text{ kg}\right)\left(1.2046\times10^{22}\right)$$
$$= 2.0176\times10^{-5}\text{ kg}$$

Thus, values remain unchanged if temperature and pressure changed because protons, electrons, neutrons is independent of temperature and pressure.

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

$^{13}_{6}C$, $^{16}_{8}O$, $^{24}_{12}Mg$, $^{36}_{26}Fe$, $^{88}_{38}Sr$

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

Here, atomic mass = 13

Atomic number = no. of protons = 6

No. of neutrons = Atomic mass – atomic no. = 13-6 = 7

$^{16}_{8}O$

Here, atomic mass = 16

Atomic number = no. of protons = 8

No. of neutrons = Atomic mass – atomic no. = 16-8 = 8

$^{24}_{12}Mg$

Here, atomic mass = 24

Atomic number = no. of protons = 12

No. of neutrons = Atomic mass – atomic no. = 24-12 = 12

$^{36}_{26}Fe$

Here, atomic mass = 56

Atomic number = no. of protons = 26

No. of neutrons = Atomic mass – atomic no. = 56-26 = 30

$^{88}_{38}Sr$

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Here, atomic mass = 88

Atomic number = no. of protons = 38

No. of neutrons = Atomic mass – atomic no. = 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

Answer: (i) ${}^{35}_{17}Cl$

(ii) ${}^{233}_{92}U$

(iii) ${}^{9}_{4}Be$

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

Answer: We have given that frequency ν and wave number $\bar{\nu}$ of the yellow light and c be the velocity of light = 3×10^{8}m/s

λ be the wavelength = 580 nm = 580×10^{-9} m

As we know that $\lambda = \frac{c}{\nu}$

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

Frequency of yellow light emitted = $5.17\times10^{14}\ s^{-1}$

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

$= \frac{1}{580\times10^{-9}}$

$= 1.72\times10^{6}\ m^{-1}$

6. Find energy of each of the photons which

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

(ii) have wave length of $0.50 Å$

Answer: (i) Energy denotes as E

So, $E = h\nu$ (i)

Here, h = planck constant = 6.626×10^{-34}Js

ν = frequency of light = 3×10^{15} Hz

Substitute values in (i) we get

$$E = \left(6.626\times10^{-34}\right)\left(3\times10^{15}\right)$$
$$E = 1.988\times10^{-18}\ J$$

(ii) Energy denotes as E and wavelength as λ

So, $E = \dfrac{hc}{\lambda}$ (i)

Here, h = planck constant = 6.626×10^{-34}Js

c = frequency of light = 3×10^{8} m/s

Substitute values in (i) we get

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

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

Answer: Here frequency = ν, wavelength = λ, Velocity of light = c

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

$$\lambda = \frac{c}{\nu} \text{ where } c = 3\times10^{8}\text{m/s}$$

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

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

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

Answer: Here, energy E of a photon $= hv$

And energy (E_n) of n photons $= n\,hv$

It is given that

Wavelength $= \lambda = 4000 \text{ pm} = 4000\times10^{-12}$ m

Velocity $= c = 3\times10^8$ m/s

Planck constant $= h = 6.626\times10^{-34}$ Js

This implies $n = \dfrac{E_n\lambda}{hc}$

Substitute values in above equation, we get

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

$= 2.012\times10^{16}$

Thus, no. of photons $= 2.012\times10^{16}$

9. A photon of wavelength 4×10^{-7} m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate

(i) the energy of the photon (eV)

(ii) the kinetic energy of the emission, and

(iii) the velocity of the photoelectron $(1\text{eV} = 1.6020\times10^{-19}\text{ J})$

Answer: (i) Here,

Planck constant $= h = 6.626\times10^{-34}$ Js

Velocity of light $= c = 3\times10^8$ m/s

Wavelength $= \lambda = 4\times10^{-7}$ m

So, energy $(E) = \dfrac{hc}{\lambda}$

Substitute values in above equation, we get

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

(ii) Kinetic energy of emission denotes as E_k

So, $E_k = h\nu - h\nu_0$

$$= (E\text{-}W)\text{eV}$$
$$= \left(\frac{4.9695\times10^{-19}}{1.6020\times10^{-19}}\right)\text{eV} - 2.13\text{eV}$$
$$= (3.1020-2.13)\text{eV}$$
$$= 0.9720\text{eV}$$

(iii) Here, velocity of photo electron denoted by v

Ans it is calculated by using formula i.e.

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

Where $\left(h\nu - h\nu_0\right)$ be kinetic energy and m be the mass

Now, substitute the values we get:

$$v = \sqrt{\frac{2\times\left(0.9720\times1.6020\times10^{-19}\right)\text{J}}{9.10939\times10^{-31}\ \text{kg}}}$$
$$= \sqrt{0.3418\times10^{12}\ \text{m}^2\ \text{s}^{-2}}$$
$$v = 5.84\times10^{5}\ \text{ms}^{-1}$$

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

Answer: Energy of sodium $(E) = \frac{N_A hc}{\lambda}$

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

$= 4.947\times10^{5}\ J\ mol^{-1}$

$= 494.7\times10^{3}\ J\ mol^{-1}$

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

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

Answer: Here P denotes power of bulb which is $25\ Watt = 25\ Js^{-1}$

And energy of one photon $(E) = hv = \frac{hc}{\lambda}$

Substitute values in above equation, we get

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

$= 34.87\times10^{-20}\ J$

So, Rate of emission 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 $6800Å$ Calculate threshold frequency (v_0) and work function W_0 of the metal.

Answer: Given that threshold wavelength $(\lambda_0) = 6800Å = 6800\times10^{-10}\ m$

And threshold frequency $(v_0) = \frac{c}{\lambda_0} = \frac{3\times10^{8}\ ms^{-1}}{6.8\times10^{-7}\ m} = 4.41\times10^{14}\ s^{-1}$

Thus, work function $(W_0) = hv_0$

$= (6.626\times10^{-34}\ J\ s)(4.41\times10^{14}\ s^{-1})$

$= 2.922\times10^{-19}\ J$

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.

Answer: Let from given condition $n_i = 4$ to $n_f = 2$ transition which give rise to spectral line

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

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

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

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

$= -\left(4.0875\times10^{-19}\ J\right)$ which indicates energy of emission

So, wavelength $(\lambda) = \frac{hc}{E}$

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

$\lambda = 4.8631\times10^{-7}\ m$

$= 486.3\times10^{-9}\ m$

$= 486\ nm$

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).

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

where Z = atomic number

n = principal quantum no.

So, for $n_1 = 5$ to $n_2 = \infty$

$\Delta E = E_\infty - E_5$

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

$= \left(2.18\times10^{-18}\ J\right)\left(\frac{1}{(5)^2}\right)$ $\quad [\because \frac{1}{\infty} = 0]$

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

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

Now, energy required for $[n_1 = 1 \text{ to } n = \infty] = 8.72\times10^{-20}$ J

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

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

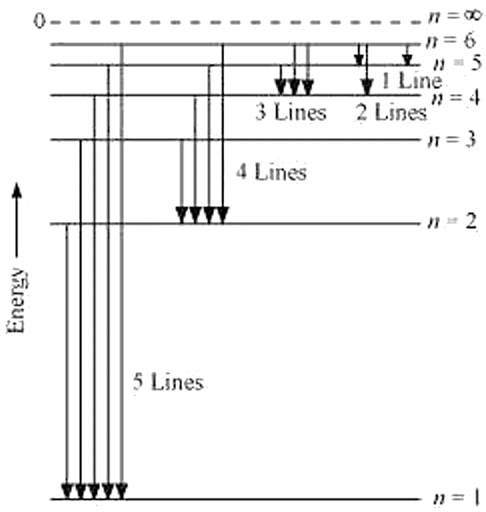
$$= \left(2.18\times10^{-18}\right)[1-0]$$

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

Thus, less energy required in fifth orbital as compared to ground state.

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

Answer:



Thus, total lines in emission spectrum = $5+4+3+2+1 = 15$

No. of spectral lines in n^{th} level to ground state = $\dfrac{n(n-1)}{2}$

As given that $n = 6$

So, no. of spectral lines = $\dfrac{6(6-1)}{2} = 15$

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.

Answer: (i) Energy associated with the fifth orbit is

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

$$= \frac{-2.18\times10^{-18}}{25}$$

$$E_5 = -8.72\times10^{-20}\ \text{J}$$

(ii) Radius of Bohr n^{th} orbit $(r_n) = (0.0529\ \text{nm})n^2$

So, for given $n = 5$

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

$$r_5 = 1.3225\ \text{nm}$$

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

Answer: As for Balmer series $n_i = 2$

Wave number $(\bar{\nu}) = \left[\frac{1}{(2)^2} - \frac{1}{n_f^2}\right]\left(1.097\times10^7\ \text{m}^{-1}\right)$

Wave no. is inversely proportional to wave length

Thus, for longest wavelength $\bar{\nu}$ is to be smallest

So, for $\bar{\nu}$ to be minimum, n_f should be minimum

Thus, for balmer series, transition from $n_i = 2$ to $n_f = 3$ is allowed

Let $n_f = 3$, we get:

$$\bar{\nu} = \left(1.097\times10^7\right)\left[\frac{1}{2^2} - \frac{1}{3^2}\right]$$

$$\bar{\nu} = \left(1.097\times10^7\right)\left[\frac{1}{4} - \frac{1}{9}\right]$$

$$= \left(1.097\times10^7\right)\left(\frac{9-4}{36}\right)$$

$$= \left(1.097\times10^7\right)\left(\frac{5}{36}\right)$$

$$\bar{\nu} = 1.5236\times10^6\ \text{m}^{-1}$$

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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 $-2.18\times10^{-11}\text{ergs}$.

Answer: Here, energy for bohr orbit $E_n = \dfrac{-\left(2.18\times10^{-18}\right)Z^2}{n^2}$

where Z = atomic no.

and ground state energy = $-2.18\times10^{-11}\text{ergs}$

$= -2.18\times10^{-11}\times10^{-7}$ J

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

So, energy required to shift electron from $n = 1$ to $n = 5$ is

$\Delta E = E_5 - E_1$

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

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

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

$= 2.0928\times10^{-18}$ J

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

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

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

19. The electron energy in hydrogen atom is given by $E_n = \dfrac{\left(-2.18\times10^{-18}\right)}{n^2}$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: We have given that $E_n = -\dfrac{2.18\times10^{-18}}{n^2}$ J

So, energy required from $n = 2$ is $\Delta E = E_\infty - E_2$

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

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

$= 0.545\times10^{-18}$ J

$= 5.45\times10^{-19}$ J

$\lambda = \frac{hc}{\Delta E}$ where λ is longest wavelength

$$\lambda = \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{5.45\times10^{-19}}$$

$= 3.647\times10^{-7}$ m

$= 3647\times10^{-10}$ m

$= 3647Å$

20. Calculate the wavelength of an electron moving with a velocity of $2.05\times10^{7}\,ms^{-1}$.

Answer: As we know that by Broglie's equation $\lambda = \frac{h}{mv}$

where λ = wavelength, m = mass, v = velocity, h = planck constant

Now, substituting all values

$$\lambda = \frac{6.626\times10^{-34}Js}{\left(9.10939\times10^{-31}\ kg\right)\left(2.05\times10^{7}\ ms^{-1}\right)}$$

$\lambda = 3.548\times10^{-11}$ 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.

Answer: According to Broglie's equation $\lambda = \frac{h}{mv}$ (i)

As we have kinetic energy (K.E.) = 3.0×10^{-25} J

As we know that K.E. = $\frac{1}{2}mv^2$

So, velocity $(v) = \sqrt{\frac{2\ K.E.}{m}}$

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

$$= \sqrt{6.5866\times10^{4}}\,\text{v}$$

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

Put values in (i) we get

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

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

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

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

Answer: Total number of electrons in Na, K, Ca, S and Ar are 11, 19, 20, 16 and 18.

Now, the number of electron in given species are,

Species	Electrons
Na^{+}	11-1=10
K^{+}	19-1=18
Ca^{2+}	20-2=18
S^{2-}	16+2=18
Ar	18

So, $K^{+}, Ca^{2+}, S^{2-}, Ar$ are isoelectronic species.

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)** $2p^{3}$ **(c)** $3p^{5}$

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

(a) $[He]2s^{1}$ **(b)** $[Ne]3s^{2}3p^{3}$ **(c)** $[Ar]4s^{2}3d^{1}$

Answer: (i) (a)

H^{-} ion

Here, electronic configuration of H atom = $1s^{1}$

Negative charge indicates gain of an electron

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

(i) (b)

Na^+ ion

Here, electronic configuration of Na atom = $1s^2 2s^2 2p^6 3s^1$

Positive charge indicates loss of an electron

So, electronic configuration of Na^+ = $1s^2 2s^2 2p^6 3s^0$ or $1s^2 2s^2 2p^6$

(i) (c)

O^{2-} ion

Here, electronic configuration of O atom = $1s^2 2s^2 2p^4$

Negative charge indicates gain of two electron

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

(i) (d)

F^- ion

Here, electronic configuration of F atom = $1s^2 2s^2 2p^5$

Negative charge indicates gain of an electron

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

(ii) (a)

Completing electronic configuration of element $1s^2 2s^2 2p^6 3s^1$

No. of electrons present in atom = $2+2+6+1 = 11$

Atomic no. = 11

(ii) (b)

Completing electronic configuration of element $1s^2 2s^2 2p^3$

No. of electrons present in atom = $2+2+3 = 7$

Atomic no. = 7

(ii) (c)

Completing electronic configuration of element $1s^2 2s^2 2p^5$

No. of electrons present in atom = $2+2+5=9$

Atomic no. = 9

(iii) (a)

Electronic configuration $[He]\ 2s^1 = 1s^2 2s^1$

Atomic no. = 3

Thus, element with the electronic configuration $[He]\ 2s^1$ = lithium (Li)

(iii) (b)

Electronic configuration $[Ne]\ 3s^2 3p^3 = 1s^2 2s^2 2p^6 3s^2 3p^3$

Atomic no. = 15

Thus, element with the electronic configuration $[Ne]\ 3s^2 3p^3$ = phosphorus (P)

(iii) (c)

Electronic configuration $[Ar]\ 4s^2 3d^1 = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$

Atomic no. = 21

Thus, element with the electronic configuration $[Ar]\ 4s^2 3d^1$ = scandium (Sc)

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

Answer: Here, for g orbitals $I = 4$

Since for any value of n of principal quantum no.

Azimuthal quantum no. (I) have the value from 0 to $(n-1)$

So, for $I = 4$, minimum value of $n = 5$

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

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Answer: Here, for $3d$ orbital

Principal quantum no. $(n) = 3$

Azimuthal quantum no. $(l) = 2$

Magnetic quantum no. (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.

Answer: (i) Atom have to be neutral

So, no. of protons = no. of electrons

Thus, no. of protons = 29

(ii) Electronic configuration of the element is

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

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

Answer: H_2^+

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

No. of electrons in H_2^+ = 2-1 = 1

H_2

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

O_2^+

No. of electrons in O_2 = 8+8 = 16

No. of electrons in O_2^+ = 16-1 = 15

28. (i) An atomic orbital has n=3 What are the possible values of l, m_l

(ii) List the quantum numbers (l, m_l) of electrons for 3d orbital.

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(iii) Which of the following orbitals are possible? 1p, 2s, 2p, 3f.

Answer: (i) As it is given that $n = 3$

For n , l have values from 0 to $(n-1)$

So, for $n = 3, l = 0, 1, 2$

For l, m_l have $(2l+1)$ values

For $l = 0, m = 0; l = 1, m = -1,0,1; l = 2, m = -2,-1,0,1,2$

For $n = 3, l = 0,1,2, m_0 = 0, m_1 = -1,0,1, m_2 = -2,-1,0,1,2$

(ii) For $3d$ orbital $l = 2$

For l, m_l have $(2l+1)$ values which is 5

For $l = 2, m_2 = -2,-1,0,1,2$

(iii) Here from given orbitals only $2s, 2p$ are possible

$1p, 3f$ can't exist

For p orbitals, $l = 1$

For n, l have values from zero to $(n-1)$

For $l = 1$, minimum value of $n = 2$

Now, for f orbital, $l = 4$

For $l = 4$ minimum value of $n = 5$

Thus, $1p, 3f$ don't exist

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) Here orbital is $1s$

(b) Here orbital is $3p$

(c) Here orbital is $4d$

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

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

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

Answer: (a) This set is not possible because principal quantum no. (n) can not be zero

(b) This set is possible

(c) This set is not possible because for n, value of 1 vary from 0 to n-1 and for n=l-1=0 and 1

(d) This set is possible

(e) This set is not possible because for n, value of l vary from 0 to 3-1 i.e. 0 to 2 i.e. 0, 1, 2

(f) This set 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=1

Answer: (a) For value of n total no. of electrons = $2n^2$

For $n = 4$, total no. of electrons = $2(4)^2 = 32$

It is fully filled orbital as $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

Thus, all electrons are paired

No. of electrons = 16

(b) Here, given values indicates that electrons are present in 3s orbital

So, no. of electrons = 2

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

Answer: Here, angular momentum of electron is

$$mvr = n\frac{h}{2\pi} \quad \ldots\ldots(1) \quad \text{where } n = 1,2,3,\ldots$$

From de Broglie's equation, we have

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

Now, substitute value of mv in (1), we get

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

$2\pi r = n\lambda$ where $2\pi r$ is circumference of Bohr 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.

Answer: From Balmer transition $n = 4$ to $n = 2$ of He^+ ion, wave no. = $\bar{v}$ and it is given by

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

Here $n_1 = 2, n_2 = 4$

$$\text{So, } \bar{v} = \frac{1}{\lambda} = R(2)^2\left(\frac{1}{4} - \frac{1}{16}\right)$$

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

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

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

Accordingly, hydrogen have same wavelength as He^+

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

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

So, by hit and trail method, above equality is true only when $n_1 = 1, n_2 = 2$

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Thus, for $n_2=2$ to $n=1$ in hydrogen have same wavelength as Balmer transition from $n=4$ to $n=2$ of He^+

34. Calculate the energy required for the process

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

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

Answer: Energy is given by $E_n = -2.18\times10^{-18}\left(\frac{Z^2}{n^2}\right)J$

$$\Delta E = E_\infty - E_1$$
$$= 0 - \left[-2.18\times10^{-18}\left\{\frac{(1)^2}{(1)^2}\right\}\right]J$$
$$\Delta E = 2.18\times10^{-18}\ J$$

For the given condition i.e. $He^+(s) \rightarrow He^{2+}_{(g)} + e^-$

So, electron is removed from $n=1$ to $n=\infty$

$$\Delta E = E_\infty - E_1$$
$$= 0 - \left[-2.18\times10^{-18}\left\{\frac{(2)^2}{(1)^2}\right\}\right]$$
$$\Delta E = 8.72\times10^{-18}\ J$$

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.

Answer: As we know that $1m = 100\ cm$ or $1cm = 10^{-2}m$

Length of scale = $20\ cm = 20\times10^{-2}\ m$

Diameter of carbon atom = $0.15\ nm = 0.15\times10^{-9}\ m$

One carbon atom contains $0.15\times10^{-9}\ m$

Now, No. of carbon atoms placed in straight line = $\frac{20\times10^{-2}\ m}{0.15\times10^{-9}\ m}$

$$= 133.33\times10^7$$
$$= 1.33\times10^9$$

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.

Answer: Given that

Length = 2.4 cm

No. of carbon atoms present = 2×10^{8}

So, diameter = $\frac{2.4\times10^{-2}\ \text{m}}{2\times10^{8}} = 1.2\times10^{-10}\ \text{m}$

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

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

$= 6.0\times10^{-11}\ \text{m}$

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

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

$= \frac{2.6\ \text{A}}{2}$

$= 1.3\times10^{-10}\ \text{m}$

$= 130\times10^{-12}\ \text{m} = 130\ \text{pm}$

(b) Length = 1.6 cm = 1.6×10^{-2} m

Diameter = 2.6×10^{-10} m

No. of zinc atoms present = $\frac{1.6\times10^{-2}\ \text{m}}{2.6\times10^{-10}\ \text{m}}$

$= 0.6153\times10^{8}\ \text{m}$

$= 6.153\times10^{7}$

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

Answer: As we know charge on one electron = 1.6022×10^{-19}C

This implies 1.6022×10^{-19}C is carried by one electron

No. of electrons which carry charge of $2.5\times10^{-16}C$

$$= \frac{1}{1.6022\times10^{-19}C}\left(2.5\times10^{-16}C\right)$$

$= 1.560\times10^{3}C$

$= 1560\ C$

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 $-1.282\times10^{-18}C$ calculate the number of electrons present on it.

Answer: As charge on oil drop = $1.282\times10^{-18}C$

Given that charge on one electron = $1.6022\times10^{-19}C$

No. of electrons present = $\frac{1.282\times10^{-18}C}{1.6022\times10^{-19}C}$

$= 0.8001\times10^{1}$

$= 8.0$

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

Answer: As thin foil of light atoms not give same result as foil of heavy atom

So, light atom carry positive charge

Thus, they carry deflection of a-particles which is positively charged

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

Answer: Let us denote atomic mass = A and atomic number = Z

i.e. general convention represents as $^{A}_{z}X$

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

So, ^{79}Br can be written but ^{35}Br it can't be written because Z is constant but A depends on relative abundance of its isotopes

Thus, it is necessary to mention A

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

Answer: Let no. of protons = x

So, no. of neutron = $x + 31.7\%$ of x

$= x + 0.317x$

$= 1.317x$

Now, we have given that mass no. = 81

No. of protons + no. of neutrons = 81

This implies $x + 1.317x = 81$

$2.317x = 81$

$$x = \frac{81}{2.317} = 34.95$$

$x = 35$

As atomic no. is defined as no. of protons which is present in its nucleus

Atomic no. = 35

And atomic symbol = ${}^{81}_{35}Br$

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.

Answer: Let no. of electrons = x

So, no. of neutrons = $x + 11.1\%$ of x

$= x + 0.111x$

$= 1.111x$

No. of electrons = $(x-1)$

So, no. of protons = $(x-1)$

We have given that mass no. = 37

Thus $(x-1) + 1.111x = 37$

$2.111x = 38$

$x = 18$

Hence symbol of the ion = ${}^{37}_{17}Cl^{-}$

44. An ion with mass number 56 possesses three unit of positive charge and 30.4% more neutrons than the electrons. Assign the symbol to this ion.

Answer: Let no. of electrons in A^{3+} be x

No. of neutrons = $x + 30.4\%$ of $x = 1.304x$

As ion is tripositive

This implies that no. of electrons in neutral atom = $x+3$

So, no. of protons in neutral atom = $x+3$

As it is given that mass no. = 56

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

Thus, no. of protons = $x+3 = 23+3 = 26$

Hence symbol = ${}^{56}_{26}Fe^{3+}$

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

Answer: Here is the increasing order of frequency for these radiations

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

And here is the increasing order of wavelength for these radiations

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.

Answer: Here power of laser = energy which emits photons

Power = $E = \dfrac{Nhc}{\lambda}$

where N = no. of protons

h = planck constant

λ = wavelength

c = velocity

Put all values to find E , we get:

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

$= 0.3302\times10^{7}\ \text{J}$

$= 3.33\times10^{6}\ \text{J}$

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 travelled by this radiation in 30 s

(c) energy of quantum

(d) number of quanta present if it produces $2\ \text{J}$ of energy.

Answer: We have given that wavelength = $616\ \text{nm} = 616\times10^{-9}\ \text{m}$

(a) Let frequency of emission denotes by ν

$\nu = \frac{c}{\lambda}$ where c = velocity, λ = wavelength

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

$= 4.87\times10^{8}\times10^{9}\times10^{-3}\ \text{s}^{-1}\text{v}$

$= 4.87\times10^{14}\ \text{s}^{-1}$

(b) As we know velocity $(c) = 3.0\times10^{8}\ \text{ms}^{-1}$

So, distance travelled in $30\ \text{s}$

$= (3.0\times10^{8}\ \text{ms}^{-1})(30\ \text{s})$

$= 9.0\times10^{9}\ \text{m}$

(c) Energy $(E) = h\nu$

$= (6.626\times10^{-34}\ \text{J s})(4.87\times10^{14}\ \text{s}^{-1})$

$= 32.27\times10^{-20}\text{J}$

(d) Energy of one photon = $32.27\times10^{-20}\ \text{J}$

So, $32.27\times10^{-20}\ \text{J}$ is present in one quantum

No. of quanta in $2\ \text{J} = \frac{2\ \text{J}}{32.27\times10^{-20}\ \text{J}}$

$= 6.19\times10^{18}$

$= 6.2\times10^{18}$

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.

Answer: As $E = \dfrac{hc}{\lambda}$

where λ = wavelength

h = planck constant

c = velocity

Putting values we get

$$E = \frac{\left(6.626\times10^{-4}\text{Js}\right)\left(3\times10^{8}\ \text{ms}^{-1}\right)}{\left(600\times10^{-9}\ \text{m}\right)}$$

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

No. of photons $= \dfrac{3.15\times10^{-18}\ \text{J}}{3.313\times10^{-19}\ \text{J}} = 9.5 = 10$ (approx)

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

Answer: As frequency of radiation $(v) = \dfrac{1}{2.0\times10^{-9}\ \text{s}} = 5.0\times10^{8}\ \text{s}^{-1}$

Energy $(E) = Nhv$

where N = no. of photons

h = planck constant

v = frequency

Putting values we get

$$E = \left(2.5\times10^{15}\right)\left(6.626\times10^{-34}\ \text{J s}\right)\left(5.0\times10^{8}\ \text{s}^{-1}\right)$$

$= 8.282\times10^{-10}$ J

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

Answer: The frequency of first transition is

$$\nu_1 = \frac{c}{\lambda_1} = \frac{3.0\times10^{8}}{589\times10^{-9}} = 5.093\times10^{14}\,s^{-1}$$

The frequency of second transition is

$$\nu_2 = \frac{c}{\lambda_2} = \frac{3.0\times10^{8}}{589.6\times10^{-9}} = 5.088\times10^{14}\,s^{-1}$$

The energy difference between two transitions is

$$\Delta E = h\left(\nu_2 - \nu_1\right) = 6.626\times10^{-34}\times(5.093-5.088)\times10^{14} = 3.313\times10^{-22}\text{ J.}$$

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

(c) calculate the kinetic energy and the velocity of the ejected photoelectron.

Answer: Given that work W_0 for caesium atom = 1.9 eV

(a) As $W_0 = \frac{hc}{\lambda_0}$ or $\lambda_0 = \frac{hc}{W_0}$

where λ_0 = threshold wavelength

h = planck constant

c = velocity

Putting values we get:

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

$$\lambda_0 = 6.53\times10^{-7}\ m$$

Thus, threshold wavelength λ_0 = 653 nm

(b) As $W_0 = h\nu_0$ or $\nu_0 = \frac{W_0}{h}$

where ν_0 = threshold frequency

h = planck constant

Putting values we get:

$$v_0 = \frac{1.9\times1.602\times10^{-19}\ J}{6.626\times10^{-34}\ J\ s}$$

$$\left(\because\ 1eV = 1.602\times10^{-19}\ J\right)$$

$$v_0 = 4.593\times10^{14}\ s^{-1}$$

(c) Here wavelength $\lambda = 500$ nm

Kinetic energy = $h(v-v_0)$

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

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

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

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

$$= 9.3149\times10^{-20}\ J$$

So, kinetic energy of ejected photoelectron = $9.3149\times10^{-20}\ J$

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

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

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

$$v = 4.52\times10^{5}\ ms^{-1}$$

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
$v\times10^{-5}(cm\ s^{-1})$	2.55	4.35	5.35

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Answer: Let threshold wavelength λ_0 nm$\left(=\lambda_0\times10^{-9}\ \text{m}\right)$

Kinetic energy is

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

From this three different equalities it formed

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

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

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

Similarly,

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

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

Divide (iii) by (i) we get:

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

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

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

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

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

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

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

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.

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Answer: Let E = energy, W_0 = work function

So, by principle of conservation:

$$E = W_0 + K.E.$$
$$\text{or } W_0 = E - K.E.$$
$$\because E = \frac{hc}{\lambda}$$

Putting all values we get:

$$E = \frac{(6.626\times10^{-34} Js)(3.0\times10^{8}\ ms^{-1})}{256.7\times10^{-9}\ m}$$
$$= 7.744\times10^{-19}\ J = \frac{7.744\times10^{-19}}{1.602\times10^{-19}} eV$$
$$E = 4.83\ eV$$

Potential applied to silver metal change into kinetic energy

K.E. = $0.35\ V = 0.35\ eV$

$$W_0 = E - K.E.$$
$$= 4.83\ eV - 0.35\ eV$$
$$= 4.48\ eV$$

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

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 $0.5\times10^{7}\ ms^{-1}$ calculate the energy with which it is bound to the nucleus.

Answer: Energy $(E) = \frac{hc}{\lambda}$

$$= \frac{(6.626\times10^{-34} Js)(3.0\times10^{8}\ ms^{-1})}{(150\times10^{-12}\ m)}$$
$$= 1.3252\times10^{-15}\ J$$
$$= 13.252\times10^{-16}\ J$$

$$\text{K.E.} = \frac{1}{2} m_e v^2$$

$$= \frac{1}{2}\left(9.10939\times10^{-31}\text{ kg}\right)\left(1.5\times10^{7}\text{ ms}^{-1}\right)^2$$

$$= 10.2480\times10^{-17}\text{ J}$$

$$= 1.025\times10^{-16}\text{ J}$$

So, energy with which electron is bound = E - K.E.

$$= 13.252\times10^{-16}\text{ J}-1.025\times10^{-16}\text{ J}$$

$$= 12.227\times10^{-16}\text{ J}$$

$$= \frac{12.227\times10^{-16}}{1.602\times10^{-19}}\text{eV}$$

$$= 7.6\times10^{3}\text{eV}$$

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

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

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

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

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

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

55. Emission transitions in the Paschen series end at orbit n=3 and start from orbit n and can be represented as $v = 3.29\times10^{15}(\text{Hz})\left[\frac{1}{3^2}-\frac{1}{n^2}\right]$ Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

Answer: It is given that wavelength = 1285 nm = 1285×10^{-9} m

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

$$\because\ v = \frac{c}{\lambda} = \frac{3.0\times10^{8}\text{ ms}^{-1}}{1285\times10^{-9}\text{ m}}$$

$$= 2.33\times10^{14}\text{ s}^{-1}$$

Substitute value of v in given equation

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

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

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

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

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

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

$$n=4.98=5$$

So, spectrum lie in 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.

Answer: We have given that

$$r=\frac{0.529n^2}{Z}\text{Å}$$

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

Now, for radius $(r_1)=1.3225\text{ nm}=1.32225\times10^{-9}\text{ m}$

$$=1322.25\times10^{-12}\text{ m}=1322.25\text{ pm}$$

$$n_1^2=\frac{r_1Z}{52.9}=\frac{1322.25Z}{52.9}$$

Similarly

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

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

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

$$\frac{n_1}{n_2}=2.5=\frac{25}{10}=\frac{5}{2}$$

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So, we get $n_1 = 5, n_2 = 2$

Hence, transition vary from fifth orbit to second orbit

So, wave number $(\overline{v}) = 1.097\times10^7\left(\frac{1}{2^2}-\frac{1}{5^2}\right)m^{-1}$

$= 1.097\times10^7\ m^{-1}\left(\frac{21}{100}\right)$

$= 2.303\times10^6\ m^{-1}$

Hence wavelength $\lambda = \frac{1}{\overline{v}}$

$= \frac{1}{2.303\times10^6\ m^{-1}}$

$= 0.434\times10^{-6}\ m$

$= 434\ nm$

57. Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is $1.6\times10^6\ ms^{-1}$ calculate de Broglie wavelength associated with this electron.

Answer: According to question by de Broglie's equation, we get

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

$\lambda = \frac{6.626\times10^{-34}Js}{(9.10939\times10^{-31}\ kg)(1.6\times10^6\ ms^{-1})}$

$= 4.55\times10^{-10}\ m$

$= 455\ pm$

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.

Answer: By using de Broglie's equation, we get

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

where v = velocity

h = planck constant

m = mass

λ = wavelength

Putting all values we get:

$$v = \frac{6.626\times10^{-34}\text{Js}}{\left(1.67493\times10^{-27}\ \text{kg}\right)\left(800\times10^{-12}\ \text{m}\right)}$$

$= 4.94\times10^{2}\ \text{ms}^{-1}$

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

59. If the velocity of the electron in Bohr's first orbit is $2.19\times10^{6}\ \text{ms}^{-1}$ calculate the de Broglie wavelength associated with it.

Answer: By using de Broglie’s equation, we get

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

where v = velocity

h = planck constant

m = mass

λ = wavelength

Putting all values we get:

$$\lambda = \frac{6.626\times10^{-34}\text{Js}}{\left(9.10939\times10^{-31}\ \text{kg}\right)\left(2.19\times10^{6}\ \text{ms}^{-1}\right)}$$

$= 3.32\times10^{-10}\ \text{m}$

$= 3.32\times10^{-10}\text{m}\times\frac{100}{100}$

$= 332\times10^{-12}\text{m}$

$= 332$ pm

60. The velocity associated with a proton moving in a potential difference of 1000 V is $4.37\times10^{5}\ \text{ms}^{-1}$ If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

Answer: From de Broglie expression we get

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

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

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

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 $\frac{h}{4}nm\times0.05\ nm$ is there any problem in defining this value.

Answer: According to Heisenberg's uncertainty principle

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

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

where Δx = uncertainty in position of electron

Δp = uncertainty in momentum of electron

Put values we get:

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

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

$$= 2.637\times10^{-23}Jsm^{-1}$$

$$\Delta p = 2.637\times10^{-23}kgms^{-1} \qquad \left(1\ J = 1kgms^{2}\ s^{-1}\right)$$

$$\text{Now, actual momentum} = \frac{h}{4\pi_m\times0.05\ nm}$$

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

$$= 1.055\times10^{-24}kgms^{-1}$$

As magnitude of actual momentum is smaller than uncertainty

Thus, value can't 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:

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1. $n = 4, l = 2, m_l = -2, m_s = \frac{-1}{2}$

2. $n = 3, l = 2, m_l = 1, m_s = \frac{1}{2}$

3. $n = 4, l = 1, m_l = 0, m_s = \frac{1}{2}$

4. $n = 3, l = 2, m_l = -2, m_s = \frac{-1}{2}$

5. $n = 3, l = 1, m_l = -1, m_s = \frac{1}{2}$

6. $n = 4, l = 1, m_l = 0, m_s = \frac{1}{2}$

Answer: Here,

For $n = 4, l = 2$, orbital occupied = $4d$

For $n = 3, l = 2$, orbital occupied = $3d$

For $n = 4, l = 1$, orbital occupied = $4p$

Thus, six electrons i.e. 1,2,3,4,5,6 are present in $4d, 3d, 4p, 3d, 3p, 4p$ orbitals respectively

So, increasing order 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?

Answer: Nuclear charge experienced by electron is dependent on the distance between nucleus and orbital where electron is present. Since distance increases, so effective nuclear charge also decreases. Further p-orbitals, $4p$ orbitals are farthest from nucleus of bromine atom with $+35$ charge. Thus, electrons in $4p$ orbital experience lowest effective nuclear charge and these electrons are shielded by electrons which are present in $2p, 3p$ orbitals with s-orbitals. So, they will experience 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

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Answer: (i) Electrons present in $2s$ orbital will experience larger effective nuclear charge as compared to the electrons in $3s$ orbital because it is closer to the nucleus

(ii) Electrons present in $4d$ orbital will experience larger effective nuclear charge as compared to the electrons in $4f$ orbital because it is closer to the nucleus

(iii) Electrons present in $3p$ orbital will experience larger effective nuclear charge as compared to the electrons in $3f$ orbital because it is closer to the nucleus

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

Answer: Higher atomic number will get higher nuclear charge. As we know Silicon has 14 protons while aluminium has 13 protons. Thus, from this we get that silicon has larger nuclear charge i.e. $+14$ than aluminium which has nuclear charge $+13$. Hence, electrons in $3p$ orbital of silicon will experience more effective nuclear charge than aluminium.

66. Indicate the number of unpaired electrons in:

(a) P

(b) Si

(c) Cr

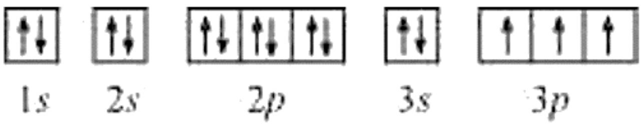
(d) Fe

(e) Kr

Answer: (a) Atomic no. = 15

Electronic configuration of $P = 1s^2 2s^2 2p^6 3s^2 3p^3$

Orbital picture of P



Thus, phosphorus has 3 unpaired electrons

(b) Atomic no. = 14

Electronic configuration of $Si = 1s^2 2s^2 2p^6 3s^2 3p^3$

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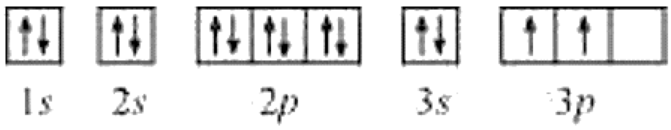
Orbital picture of Si

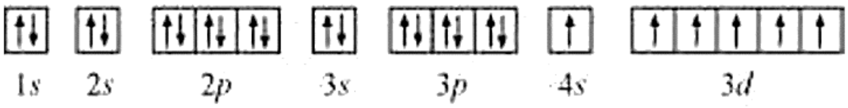


Thus, silicon has 2 unpaired electrons

(c) Atomic no. = 24

Electronic configuration of Cr = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

Orbital picture of Cr

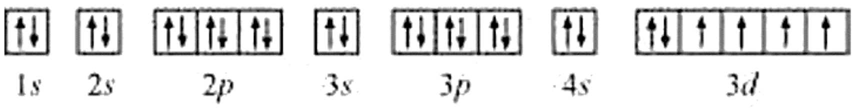


Thus, chromium has 6 unpaired electrons

(d) Atomic no. = 26

Electronic configuration of Fe = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

Orbital picture of Fe

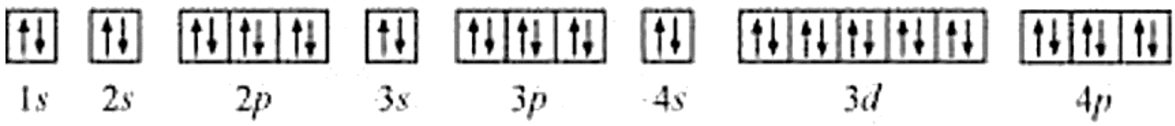


Thus, Iron has 4 unpaired electrons

(e) Atomic no. = 36

Electronic configuration of Kr = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

Orbital picture of Kr



Thus, krypton are fully occupied, so no unpaired electrons.

67. (a) How many sub-shells are associated with n=4

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(b) How many electrons will be present in the sub-shells having m_s value of $\frac{-1}{2}$ for n=4.

Answer: (a) Given that $n = 4$

For given values of n, l have values from 0 to (n-1)

So, $l = 0, 1, 2, 3$

Hence, 4 sub shells associated with $n = 4$ i.e. s, p, d, f

(b) No. of orbital in n^{th} shell = n^2

For $n = 4$

No. of orbitals = 16

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

So, no. of electrons = 16