Chapter – 2

Structure of Atom

2marks:

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

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

Answer:

(i) Mass of 1 electron = 9.108 x 10-28 g

Hence,

1 g = 1/(9.108 x 10-28) = 1.098 x 1027 electrons

(ii) Mass of one mole of electron = 9.108 x 10-28 x 6.022 x 1023

We get,

= 5.48 x 10-4 g

Charge on one mole of electron = 1.6 x 10-19 C x 6.022 x 1023

We get,

= 9.63 x 104 C

2. How Many Neutrons and Protons are There in the Following Nuclei? 136C,168O,2412Mg,5626Fe,8838Sr

Answer: 136C:

Atomic mass =13

Atomic number = Number of Protons = 6

Number of neutrons = (Atomic Mass)-(Atomic Number)=13-6=7

168O:

Atomic mass =16

Atomic number = Number of Protons = 8

Number of neutrons = (Atomic Mass)-(Atomic Number) =16-8=8

2412Mg:

Atomic mass =24

Atomic number = Number of Protons = 12

Number of neutrons = (Atomic Mass)-(Atomic Number)=24-12 =12

5626Fe:

Atomic mass = 56

Atomic number = Number of Protons = 26

Number of neutrons = (Atomic Mass)-(Atomic Number) = 56-26=30

8838Sr:

Atomic mass = 88

Atomic number = Number of Protons = 38

Number of neutrons = (Atomic Mass)-(Atomic Number)= 88-38=50

3. Write the Complete Symbol for the Atom with the Given Atomic Number (Z) and Atomic mass (A)

a. Z = 17, A = 35

Answer:

3517Cl

b. Z = 92, A = 233

Ans: 23392U

c. Z = 4, A = 9

Ans: 94Be

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

Answer:

Energy of sodium (E)

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

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

= 4.947 × 105 J mol^{-1}

= 494.7 × 103 J mol^{-1}

= 494 kJ mol^{-1}

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

Answer:

The number of spectral lines produced when an electron in the nth level drops down to the ground state is given by n(n-1)/2.

Given, n=6

∴ Number of spectral lines = 6×5/2 = 15

also given by, ∑(n2 - n1) = ∑(6-1) = ∑5 = 5+4+3+2+1 = 15

6. (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) $E_n = -21.8\times10^{-19}/n^2$ J

∴ $E_5 = -21.8\times10^{-18}/5^2$ J $= 8.72\times10^{-20}$ J

(ii) For H atom, $r_n = 0.529 \times n^2$ Å

$\therefore r_5 = 0.529 \times 5^2 = 13.225$ Å $= 1.3225$ nm

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

Answer:

For the Balmer series, $n_1 = 2$. Hence, $\tilde{v} = R(1/2^2 - 1/n2^2)$

$\tilde{v} = 1/\lambda$ (inversely proportional)

For λ to be maximum, $\tilde{v}$ should be minimum. This can be happened when n2 is minimum i.e. n2 = 3. Hence, $\tilde{v} = (1.097\times10^7\ m^{-1})$

$(1/2^2 - 1/3^2) = 1.097\times10^7\times5/36\ m^{-1} = 1.523\times10^6\ m^{-1}$

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

Answer:

By de Broglie equation,

$\lambda = h/mv = 6.626 \times 10^{-34}$ Js / $(9.11\times10^{-31}$ kg$)$ $(2.05\times10^7\ ms^{-1}) = 3.55 \times10^{-11}$ m

9. The mass of an electron is 9.11×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

Answer:

K.E. = $1/2\ mv^2$

$= \sqrt{\dfrac{2 \times 3.0 \times 10^{-25}\,\text{J}}{9.11 \times 10^{-31}\,\text{kg}}}$ ∴ v = √2 K.E./m

= 812 ms^{-1} By de Broglie equation, λ = h/mv = 6.626 × 10^{-34} Js/ (9.11 × 10^{-31} kg) (812 ms^{-1}) = 8.967 ×10^{-7} m

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

Answer:

For g-orbitals, l = 4.

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

∴ For l = 4, minimum value of n = 5

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

Answer:

For the 3d orbital:

Principal quantum number (n) = 3

Azimuthal quantum number (l) = 2

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

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

Answer:

(i) For neutral atom, number of protons = number of electrons.

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

(ii) The electronic configuration of the atom with Z=29 is

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

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

Answer:

H_2^+ = 2-1 = 1 electron

$H_2 = {}_1H + {}_1H$ = 2 electrons

O_2^+ = 16-1 = 15 electrons

4marks:

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

Answer:

From the expression,

We get,

$\nu = \frac{c}{\lambda}$ …….. (i)

Where,

ν = frequency of yellow light

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

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

Substituting the values in expression (i):

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

Thus, frequency of yellow light emitted from the sodium lamp

= 5.17 × 1014 s^{-1}

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

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

2. Find energy of each of the photons which

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

(ii) have wavelength of 0.50 Å.

Answer:

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

E = $h\nu$

Where,

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

ν = frequency of light = 3×10^{15} Hz

Substituting the values in the given expression of E:

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

E = 1.988×10^{-18} J

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

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

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

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

Substituting the values in the given expression of E:

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^{8})}{0.50 \times 10^{-10}} = 3.976 \times 10^{-15}\ \text{J}$$
$$\therefore E = 3.98 \times 10^{-15}\ \text{J}$$

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

Answer:

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

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

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

Where,

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

Substituting the value in the given expression of λ:

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

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

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

Answer:

Energy (E) of a photon = hν

Energy (En) of ‘n’ photons = nhν

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

Where,

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

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

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

Substituting the values in the given expression of n:

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

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

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

Answer:

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

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

Substituting the values in the given expression of E:

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

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

Rate of emission of quanta per second

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

6. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate

threshold frequency ((ν_0)) and work function (W0) of the metal.

Answer:

Threshold wavelength of radian $(\lambda_0) = 6800 Å$ = 6800 × 10–10 m

Threshold frequency (ν_0) of the metal

$= \frac{c}{\lambda_0} = \frac{3\times10^8\ ms^{-1}}{6.8\times10^{-7} m}$ = 4.41 × 1014 s–1

Thus, the threshold frequency (ν_0) of the metal is 4.41 × 1014 s1.

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

= (6.626 × 10^{-3}4 Js) (4.41 × 1014 s^{-1})

= 2.922 × 10^{-19} J

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

Answer:

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

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

Substituting the values in the given expression of E:

$$E = 2.18\times10^{-18}\left[\frac{1}{4^2}-\frac{1}{2^2}\right]$$
$$= 2.18\times10^{-18}\left[\frac{1-4}{16}\right]$$
$$= 2.18\times10^{-18}\times\left(-\frac{3}{16}\right)$$

E = – (4.0875 × 10–19 J)

The negative sign indicates the energy of emission.

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

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

Substituting the values in the given expression of λ:

$$\lambda = \frac{\left(6.626\times10^{-34}\right)\left(3\times10^{8}\right)}{4.0875\times10^{-19}}$$
$$\lambda = 4.8631\times10^{-7}\text{ m}$$
$$= 486.3\times10^{-9}\text{ m}$$
$$= 486\text{ nm}$$

8. 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:

1erg = 10^{-7} J

As ground state electronic energy is –2.18×10^{-11} ergs, this means that $E_{n\,=}$ $-21.8 \times10^{-11}/n^2$ ergs.

$\Delta E = E_5 - E_1 = 2.18\times10^{-11} (1/1^2 - 1/5^2) = 2.18 \times10^{-11} (24/25) = 2.09 \times 10^{-11}$ ergs = 2.09×10^{-18} J

When electron returns to ground state (n=1), energy emitted = 2.09×10^{-11} ergs.

As, $E = h\nu = hc/\lambda$

$\Rightarrow \lambda = hc/E = (6.626 \times 10^{-27}$ erg sec) $(3.0 \times 10^{10}$ cm $s^{-1}) / 2.09 \times 10^{-11}$ ergs

$= 9.51 \times 10^{-6}$ cm $= 951 \times 10^{-8}$ cm = 951 Å

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

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

Answer:

Notes: Isoelectronic are the species having same number of electrons.

A positive charge means the shortage of an electron.

A negative charge means gain of electron.

Number of electrons in Na^+ = 11-1 = 10

Number of electrons in K^+ = 19-1 = 18

Number of electrons in Mg^{2+} = 12-2 = 10

Number of electrons in Ca^{2+} = 20-2 = 18

Number of electrons in S^{2-} = 16+2 = 18

Number of electrons in Ar = 18

Hence, the following are isoelectronic species:

1) Na^+ andMg^{2+} (10 electrons each)

2) K^+, Ca^{2+}, S^{2-} and Ar (18 electrons each)

10 (i). What is the lowest value of n that allows g orbitals to exist?

(ii). Give the number of electrons in the species H_2^+, H_2 and O_2^+

Answer:

For g-orbitals, l = 4.

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

∴ For l = 4, minimum value of n = 5

(ii):

H_2^+ = 2-1 = 1 electron

$H_2 = {}_1H + {}_1H$ = 2 electrons

O_2^+ = 16-1 = 15 electrons

11.(i)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.

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

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

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

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

(d) $n = 2$, $l = 1$, $m_l = 0$, $m_s = –½$

(e) $n = 3$, $l = 3$, $m_l = –3$, $m_s = +½$

(f) $n = 3$, $l = 1$, $m_l = 0$, $m_s = +½$

Answer:

(a) 1s

(b) 3p

(c) 4d

(d) 4f

(ii)**:**

(a) Not possible because $n \neq 0$

(b) Possible

(c) not possible because when $n=1, l \neq 1$

(d) Possible

(e) Not possible because when $n=3$, $l \neq 3$

(f) Possible

7marks:

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

Answer:

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

Where,

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

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

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

Substituting the values in the given expression of E:

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

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

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

$$= h\nu - h\nu_0$$
$$= (E - W)\,\text{eV}$$
$$= \left(\frac{4.9695\times10^{-19}}{1.6020\times10^{-19}}\right)\text{eV} - 2.13\ \text{eV}$$

= (3.1020 – 2.13) eV

= 0.9720 eV

Hence, the kinetic energy of emission is 0.97 eV.

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

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

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

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

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

v = 5.84 × 105 ms^{-1}

Hence, the velocity of the photoelectron is 5.84 × 105 ms^{-1}.

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

Answer:

The expression of energy is given by,

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

Where,

Z = atomic number of the atom

n = principal quantum number

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

$$\Delta E = E_\infty - E_5$$
$$= \left[\left\{\frac{-(2.18\times10^{-18}\ \text{J})(1)^2}{(\infty)^2}\right\} - \left\{\frac{-(2.18\times10^{-18}\ \text{J})(1)^2}{(5)^2}\right\}\right]$$
$$= (2.18\times10^{-18}\ \text{J})\left(\frac{1}{(5)^2}\right) \qquad \left(\text{Since } \frac{1}{\infty} = 0\right)$$
$$= 0.0872\times10^{-18}\ \text{J}$$
$$\Delta E = 8.72\times10^{-20}\ \text{J}$$

Hence, the energy required for ionization from n = 5 to n =

∞ is 8.72 × 10–20 J.

Energy required for n1 = 1 to n =

∞,

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

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

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

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

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

3. (i) Write the electronic configurations of the following ions: (a) H^- (b) Na^+ (c) O^{2-} (d) F^-

(ii) What are the atomic numbers of elements whose outermost electrons are represented by (a) $3s^1$ (b) $2p^3$ and (c) $3p^5$?

(iii) Which atoms are indicated by the following configurations ? (a) [He]$2s^1$ (b) [Ne] $3s^2$ $3p^3$ (c) [Ar] $4s^2$ $3d^1$.

Answer:

(i) (a) $_1H = 1s^1$ · A negative charge means gain of electron.

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

(b) $_{11}Na = 1s^2 2s^2\ 2p^6 3s^1$ · A positive charge means the shortage of an electron.

∴ electronic configuration of $Na^+ = 1s^2 2s^2 2p^6$

(c) $_8O = 1s^2 2s^2 2p^4$

∴ electronic configuration of $O^{2-} = 1s^2\ 2s^2\ 2p^6$ (d) $_9F = 1s^2\ 2s^2\ 2p^5$

∴ electronic configuration of $F^- = 1s^2\ 2s^2\ 2p^6$ (ii) (a) $3s^1$

Completing the electron configuration of the element

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

∴ Number of electrons present in the atom of the element = 2+2+6+1

= 11

∴ Atomic number of the element = 11

(b) $2p^3$

Completing the electron configuration of the element as $1s^2 2s^2 2p^3$

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

∴ Atomic number of the element = 7

(c) $3p^5$

Completing the electron configuration of the element

as $1s^2\ 2s^2\ 2p^6\ 3s^2 3p^5$

∴ Number of electrons present in the atom of the element =

2+2+6+2+5 = 17

∴ Atomic number of the element = 9

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

electronic configuration = $1s^2 2s^1$

∴ Atomic number of the element = 2+1 = 3

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

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

electronic configuration = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

∴ Atomic number of the element = 2+2+6+2+3 = 15

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

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

electronic configuration = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$

∴ Atomic number of the element = 2+2+ 6+2+ 6+2+1 = 21

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

4. (i) An atomic orbital has n = 3. What are the possible values of l and m_l ?

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

(iii) Which of the following orbitals are possible?

1p, 2s, 2p and 3f

(iv). How many electrons in an atom may have the following quantum numbers?

(a) n = 4, m_s = –½ (b) n = 3, l = 0

Answer:

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

∴ For n = 3 , l = 0, 1, 2

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

When l = 0, m = 0

l = 1, m = - 1, 0, 1

l = 2, m = - 2, - 1, 0, 1, 2

l = 3, m = -3, -2, -1, 0, 1, 2, 3

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

∴ For l = 2

m_2 = -2, -1, 0, 1, 2

(iii) 1p is not possible because when n = 1, l = 0. (for p, l = 1)

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

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

3f is not possible because when n=3, l = 0, 1 , 2 (for f, l=3)

(iv)

(i) The total number of electrons in n is given by $2n^2$

n=4, Number of electrons = $2 \times 4^2 = 32$

Half of 32 electrons will have spin quantum number $m_s = -½$ i.e. 16 electrons

(ii) n=3 and l=0 means it is 3s orbital which can have only 2 electrons.

5. 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

Answer:

Let the threshold wavelength to be λ_0 nm = $\lambda_0\times10^{-9}$ m .

Following equation holds true for photoelectric emission in given case:

K.E. = 1/2 $mv^2 = h(\nu - \nu_0)$

$\Rightarrow 1/2\ mv^2 = h\nu - h\nu_0$

$\Rightarrow h\nu_0 = h\nu - 1/2\ mv^2$

$\Rightarrow hc/\lambda_0 = hc/\lambda - 1/2\ mv^2$

$$\Rightarrow \lambda_0 = \frac{1}{\frac{1}{\lambda} - \frac{1}{2}\frac{m}{hc}v^2} = \frac{1}{\frac{1}{\lambda} - \frac{1}{2}\frac{9.1\times10^{-31}}{6.6\times10^{-34}\times3\times10^{8}}v^2}$$

(a) Substituting the value of λ and v from the above given data, we get three values of λ_0 as,

$\lambda_{0(1)}$ = 541 nm

$\lambda_{0(2)}$ = 546 nm

$\lambda_{0(3)}$ = 542 nm

Threshold frequency = λ_{av} = $\{\lambda_{0(1)} + \lambda_{0(2)} + \lambda_{0(3)}\}/3$ = (541+ 546+ 542)/3 = 543 (approx 540)

(b) Part of this question can't be solved due to incorrect value of v i.e 5.35.

Students can assume this value as 5.20 if they want to solve this question.

(i). 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:

Total length = 2.4 cm

Total number of atoms along the length = 2×10^{8}

∴ Diameter of each atom = $2.4 \text{ cm}/2 \times 10^{8} = 1.2\times10^{-8}$ cm

∴ Radius of the atom = Diameter/2 = 1.2×10^{-8} cm/2 = 0.6×10^{-8} cm

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

Answer:

(a) Radius of zinc atom = 2.6Å /2 = 1.3Å = 1.3×10^{-10} m = 130×10^{-12} pm

(b) Given length = 1.6 cm = 1.6×10^{-2} m

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

∴ No. of atoms present along the length = $1.6 \times 10^{-2}/ 2.6 \times 10^{-10} = 6.154 \times 10^{7}$

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

Answer:

Charge on one electron = 1.602×10^{-19} C

∴ Number of electrons carrying 2.5×10^{-16} C charge = $2.5 \times 10^{-16}/ 1.602 \times 10^{-19}$ =1560

(iv)In Milikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the static electric charge

on the oil drop is –1.282×10^{-18} C, calculate the number of electrons present on it.

Answer:

As in the above question,

Number of electrons present in oil drop = $-1.282 \times 10^{-18} / 1.602\times10^{-19} = 8$

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

Answer:

Heavy atoms have nucleus carrying large amount of positive charge. Therefor, some α-particles will easily deflected back. Also a number of α-particles are deflected through small angles because of large positive charge.

If light atoms are used, their nuclei will have small positive charge, hence the number of α-particles getting deflected even through small angles will be negligible.

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

Answer:

$^{35}Br_{79}$ is not acceptable because atomic number should be written as subscript, while mass number should be written as superscript. $_{35}Br$ is not acceptable because atomic number of an element is fixed.

However, mass number is not fixed as it depends upon the isotopes taken. Hence, it is essential to indicate mass number.

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

Answer:

Mass number = protons + neutrons = p + n = 81 (given)

Let p be x, then neutrons = x + (31.7 /100)x = 1.317 x

∴ x + 1.317 x = 81

⇒ 2.317 x = 81

⇒ x = 81/2.317 = 35

Thus, protons = 35 = atomic number.

The symbol of the element is $^{81}Br_{35}$ or $^{81}_{35}Br$

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

Answer :

Let the number of electrons in the ion = x

Then, number of neutrons = x+(11.1 x/ 100) = 1.111 x

Number of electrons in the neutral atom = x-1 (ion possesses one unit of negative charge)

∴ Number of protons = x-1

Mass number = No. of protons + No. of neutrons

∴ 1.111 x + x - 1 = 37

⇒ 2.111x = 38

⇒ x = 18

∴ No. of protons = Atomic no. = x-1 = 18-1 = 17

The symbol of the ion is $^{37}_{17}Cl^{-1}$

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

Answer:

Let the number of electrons in the ion = x

Then, number of neutrons = x+(30.4 x/ 100) = 1.304 x

Number of electrons in the neutral atom = x+3 (ion possesses 3 units of positive charge)

∴ Number of protons = x+3

Mass number = No. of protons + No. of neutrons

∴ 1.304 x + x +3 = 56

⇒ 2.304x = 53

⇒ x = 23

∴ No. of protons = Atomic no. = x+3 = 23+3 = 26

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

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

Answer:

The increasing order of frequency is as follows:

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

(iv)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:

E = Nhν = Nhc/λ = $(5.6 \times 10^{24}) \times (6.626 \times 10^{-34}\ Js \times 3.0 \times 10^{8}\ ms^{-1}) / 337.1 \times 10^{-9}\ m = 3.3 \times 10^{6}\ J$

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

Answer:

$\lambda = 616\ nm = 616 \times 10^{-9}\ m$

(a) Frequency, $\nu = c/\lambda = 3.0 \times 10^{8}\ ms^{-1} / 616 \times 10^{-9}\ m = 4.87 \times 10^{14}\ s^{-1}$ (b) Velocity of the radiation = $3.0 \times 10^{8}\ ms^{-1}$

∴ Distance travelled in 30 s = $30 \times 3 \times 10^{8}\ m = 9.0 \times 10^{9}\ m$

(c) $E = h\nu = hc/\lambda = (6.626 \times 10^{-34}\ Js \times 3.0 \times 10^{8}\ ms^{-1}) / 616 \times 10^{-9}\ m = 32.27 \times 10^{-20}\ J$

(d) No. of quanta in 2J of energy = $2J / 32.27 \times 10^{-20}\ J = 6.2 \times 10^{18}$

(ii)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:

Energy of one photon = $h\nu = hc/\lambda = (6.626 \times 10^{-34}\ Js \times 3.0 \times 10^{8}\ ms^{-1}) / 600 \times 10^{-9}\ m = 3.313 \times 10^{-19}\ J$

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

∴ No. of photons received = $3.15 \times 10^{-18}\ J/3.313 \times 10^{-19}\ J = 9.51$ (approx 10)

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

Answer:

Frequency = $1/2 \times10^{-19}\ s = 0.5 \times10^{9}\ s^{-1}$ Energy = $Nh\nu = (2.5 \times 10^{15}) \times (6.626 \times 10^{-34}\ Js)\times (0.5 \times10^{9}\ s^{-1}) = 8.28 \times10^{-10}\ J$

(iv)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:

λ_1 = 589 nm = 589×10^{-9} m

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

λ_2 = 589.6 nm = 589.6×10^{-9} m

∴ $\nu_2 = c/\lambda_2 = 3.0\times10^{8}\ ms^{-1}/ 589.6 \times 10^{-9}\ m = 5.088 \times 10^{14}\ s^{-1}$

$\Delta E = E_2 - E_1 = h(\nu_2 - \nu_1) = (6.626 \times 10^{-34}$ Js$) \times (5.093 - 5.088) \times 10^{14}$ s^{-1} $= 3.31 \times 10^{-22}$ J

(v)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 500nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

Answer:

(a) Work function $(W_0) = h\nu_0$

$\therefore \nu_0 = W_0/h = 1.9 \times 1.602 \times 10^{-19}$ J/ 6.626×10^{-34} Js $= 4.59 \times 10^{14}$ s^{-1}

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

(b) $\lambda_0 = c/\nu_0 = 3.0 \times 10^8$ ms^{-1}/4.59×10^{14} s^{-1} $= 6.54 \times 10^{-7}$ m $= 654 \times 10^{-9} =$ 654 nm

(c) K.E. of ejected electron $= h(\nu - \nu_0) = hc\,(1/\lambda - 1/\lambda_0)$

$= (6.626 \times 10^{-34}$ Js$\times 3.0 \times 10^8$ ms$^{-1}) \times (1/500 \times 10^{-9}$ m $- 1/654 \times 10^{-9}$ m$)$

$= (6.626 \times 3.0 \times 10^{-26}/10^{-9}) \times (154/500 \times 654)$ J $= 9.36 \times 10^{-20}$ J

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

$\therefore 1/2 \times (9.11 \times 10^{-31}$ kg$)\, v^2 = 9.36 \times 10^{-20}$ kgm^{2}s^{-2}

$\Rightarrow v^2 = 2.055 \times 10^{11}$ m^{2}s^{-2} $= 20.55 \times 10^{10}$ m^{2}s^{-2}

$\Rightarrow v = 4.53 \times 10^5$ ms^{-1}

Fill in the blanks:

1.In Bohr's model, electrons revolve in fixed orbits called __________ around the nucleus.

Answer: shells

2.The maximum number of electrons that can occupy the third shell in an atom is __________.

Answer: 18

3.The subatomic particle with a positive charge found in the nucleus is called a __________.

Answer: proton

4.The atomic number of an element is determined by the number of __________ in its nucleus.

Answer: protons

5.Isotopes of an element have the same number of __________ but different numbers of neutrons.

Answer: protons

6.The mass number of an atom is the sum of its __________ and __________.

Answer: protons, neutrons

7.The region around the nucleus where an electron is likely to be found is called an __________.

Answer: orbital

8.The total number of orbitals in the p subshell is __________.

Answer: 3

9.According to the Aufbau principle, electrons occupy the lowest energy orbitals first, filling up one __________ before moving to the next.

Answer: subshell

10.The Pauli exclusion principle states that no two electrons in an atom can have the same set of __________ quantum numbers.

Answer: four

Multiple choice:

1.In Bohr's atomic model, electrons are arranged in:

a) Random orbits

b) Circular orbits

c) Elliptical orbits

d) Spiral orbits

Answer:

b) Circular orbits

2.The nucleus of an atom is composed of:

a) Electrons

b) Protons and electrons

c) Neutrons and electrons

d) Protons and neutrons

Answer:

d) Protons and neutrons

3.The atomic number of an element is equal to the number of:

a) Electrons

b) Neutrons

c) Protons

d) Electrons and protons combined

Answer:

c) Protons

4.Isotopes of an element have the same number of:

a) Electrons

b) Protons

c) Neutrons

d) Both b) and c)

Answer:

d) Both b) and c)

5.The mass number of an atom is the sum of the number of:

a) Electrons and protons

b) Neutrons and electrons

c) Protons and neutrons

d) Neutrons, protons, and electrons

Answer:

c) Protons and neutrons

6.The quantum number that specifies the shape of an orbital is:

a) Principal quantum number (n)

b) Azimuthal quantum number (l)

c) Magnetic quantum number (m)

d) Spin quantum number (s)

Answer:

b) Azimuthal quantum number (l)

7.The region around the nucleus where an electron is likely to be found is called:

a) Shell

b) Subshell

c) Orbital

d) Electron cloud

Answer:

c) Orbital

8.The maximum number of electrons that can occupy the p subshell is:

a) 2

b) 6

c) 8

d) 18

Answer:

b) 6

9.According to the Aufbau principle, electrons fill orbitals starting from the one with the:

a) Lowest energy

b) Highest energy

c) Lowest principal quantum number

d) Highest principal quantum number

Answer:

a) Lowest energy

10.The Pauli exclusion principle states that no two electrons in an atom can have the same set of:

a) Principal quantum numbers

b) Azimuthal quantum numbers

c) Magnetic quantum numbers

d) Spin quantum numbers

Answer:

d) Spin quantum numbers

Summary:

The chapter on the "Structure of Atom" delves into the fundamental building blocks of matter, exploring the intricate details of atomic structure. The chapter begins with the historical development of atomic models, ranging from Dalton's indivisible atom to Rutherford's nuclear model. The Bohr model introduces quantized energy levels and orbits for electrons, providing insights into the electronic configuration of atoms.

The concept of quantum numbers and the Schrödinger equation are introduced to explain the probability distribution of electrons in atomic orbitals. The chapter emphasizes the importance of electron configurations, detailing how electrons fill up various subshells and orbitals following the Aufbau principle, Pauli exclusion principle, and Hund's rule.

Isotopes, atoms of the same element with different neutron numbers, are discussed, along with the calculation of atomic mass considering isotopic abundance. The chapter also covers the periodic table and the trends in atomic size, ionization energy, and electron affinity.