Chapter 2: Structure of Atom

Exercises — Exercises

Q1 (i) Calculate the number of electrons which will together weigh one gram. (ii) Calculate the mass and charge of one mole of electrons.

SOLUTION

Given: mass of one electron = 9.10938×10^{-28} g; charge on one electron = 1.6022×10^{-19} C; $N_A = 6.022 \times 10^{23}$ mol^{-1}

(i) Number of electrons weighing 1 g:
$n = 1\ \text{g} / (9.10938 \times 10^{-28}\ \text{g}) = 1.098 \times 10^{27}$ electrons

(ii) Mass of 1 mole of electrons:
$m = N_A \times m_e = 6.022 \times 10^{23} \times 9.10938 \times 10^{-28}\ \text{g} = 5.486 \times 10^{-4}\ \text{g} = 5.486 \times 10^{-7}\ \text{kg}$

Charge of 1 mole of electrons (1 Faraday):
$q = N_A \times e = 6.022 \times 10^{23} \times 1.6022 \times 10^{-19}\ \text{C} = 9.648 \times 10^{4}\ \text{C}$

Answer: (i) 1.098×10^{27} electrons; (ii) mass = 5.486×10^{-4} g, charge = 9.648×10^{4} C

Q2 (i) Calculate the total number of electrons present in one mole of methane. (ii) Find (a) the total number and (b) the total mass of neutrons in 7 mg of 14C. (Assume that mass of a neutron = 1.675×10^{-27} kg). (iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3 at STP. Will the answer change if the temperature and pressure are changed?

SOLUTION

Given: $N_A = 6.022 \times 10^{23}$ mol^{-1}; $m_{neutron} = 1.675 \times 10^{-27}$ kg; $m_{proton} = 1.6726 \times 10^{-27}$ kg

(i) Electrons in 1 mol of CH_4:
Each CH_4 molecule has 6 (C) + 4×1 (H) = 10 electrons.
Total = $10 \times N_A = 10 \times 6.022 \times 10^{23} = 6.022 \times 10^{24}$ electrons

(ii) Neutrons in 7 mg of 14C (A = 14, Z = 6, neutrons per atom = 8):
Moles of 14C = $7 \times 10^{-3}\ \text{g} / 14\ \text{g mol}^{-1} = 5.0 \times 10^{-4}$ mol
Number of atoms = $5.0 \times 10^{-4} \times 6.022 \times 10^{23} = 3.011 \times 10^{20}$ atoms
(a) Total neutrons = $8 \times 3.011 \times 10^{20} = 2.409 \times 10^{21}$

(b) Total mass = $2.409 \times 10^{21} \times 1.675 \times 10^{-27}$ kg = 4.035×10^{-6} kg

(iii) Protons in 34 mg of NH_3 (M = 17 g mol^{-1}; each NH_3 has 7 + 3 = 10 protons):
Moles of NH_3 = 34×10^{-3} g / 17 g mol^{-1} = 2.0×10^{-3} mol
Number of molecules = $2.0 \times 10^{-3} \times 6.022 \times 10^{23} = 1.204 \times 10^{21}$
(a) Total protons = $10 \times 1.204 \times 10^{21} = 1.204 \times 10^{22}$
(b) Total mass = $1.204 \times 10^{22} \times 1.6726 \times 10^{-27}$ kg = 2.014×10^{-5} kg

The number of molecules in a fixed mass (34 mg) depends only on the molar mass, not on temperature or pressure. Therefore, the answer will NOT change with temperature and pressure.

Answer: (i) 6.022×10^{24}; (ii)(a) 2.409×10^{21} neutrons, (b) 4.035×10^{-6} kg; (iii)(a) 1.204×10^{22} protons, (b) 2.014×10^{-5} kg; no change with T and P

Q3 How many neutrons and protons are there in the following nuclei? 13/6 C, 16/8 O, 24/12 Mg, 56/26 Fe, 88/38 Sr

SOLUTION

For any nucleus: protons = Z; neutrons = A – Z

13/6 C: Z = 6, A = 13 -> protons = 6, neutrons = 13 – 6 = 7
16/8 O: Z = 8, A = 16 -> protons = 8, neutrons = 16 – 8 = 8
24/12 Mg: Z = 12, A = 24 -> protons = 12, neutrons = 24 – 12 = 12
56/26 Fe: Z = 26, A = 56 -> protons = 26, neutrons = 56 – 26 = 30
88/38 Sr: Z = 38, A = 88 -> protons = 38, neutrons = 88 – 38 = 50

Answer: 13C: 6p, 7n; 16O: 8p, 8n; 24Mg: 12p, 12n; 56Fe: 26p, 30n; 88Sr: 38p, 50n

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Q4 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.

SOLUTION

The complete symbol is written as (A/Z) X where X is the chemical symbol of the element.

(i) Z = 17 is Chlorine (Cl): symbol = 35/17 Cl
(ii) Z = 92 is Uranium (U): symbol = 233/92 U
(iii) Z = 4 is Beryllium (Be): symbol = 9/4 Be

Answer: (i) 35/17 Cl (ii) 233/92 U (iii) 9/4 Be

Q5 Yellow light emitted from a sodium lamp has a wavelength (lambda) of 580 nm. Calculate the frequency (nu) and wavenumber of the yellow light.

SOLUTION

Given: lambda = 580 nm = 580×10^{-9} m; c = 3×10^{8} m s^{-1}

Frequency: nu = c / lambda = (3×10^{8} m s^{-1}) / (580×10^{-9} m)
nu = 5.17×10^{14} Hz

Wavenumber: v-bar = 1 / lambda = 1 / (580×10^{-9} m) = 1.72×10^{6} m^{-1}
In cm^{-1}: v-bar = 1 / (580×10^{-7} cm) = 1.724×10^{4} cm^{-1}

Answer: nu = 5.17×10^{14} Hz; wavenumber = 1.72×10^{6} m^{-1} = 1.724×10^{4} cm^{-1}

Q6 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.

SOLUTION

Given: h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}

(i) nu = 3×10^{15} Hz
E = h × nu = $6.626 \times 10^{-34} \times 3 \times 10^{15} = 1.988 \times 10^{-18}$ J

(ii) lambda = 0.50 Å = 0.50×10^{-10} m
E = hc / lambda = ($6.626 \times 10^{-34} \times 3 \times 10^{8}$) / ($0.50 \times 10^{-10}$)
E = 3.976×10^{-15} J

Answer: (i) E = 1.988×10^{-18} J; (ii) E = 3.976×10^{-15} J

Q7 Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0×10^{-10} s.

SOLUTION

Given: period T = 2.0×10^{-10} s; c = 3×10^{8} m s^{-1}

Frequency: nu = 1 / T = 1 / (2.0×10^{-10} s) = 5.0×10^{9} Hz

Wavelength: lambda = c / nu = (3×10^{8} m s^{-1}) / (5.0×10^{9} s^{-1}) = 6.0×10^{-2} m = 6.0 cm

Wavenumber: v-bar = 1 / lambda = 1 / (6.0×10^{-2} m) = 16.7 m^{-1} = 0.167 cm^{-1}

Answer: nu = 5.0×10^{9} Hz; lambda = 6.0×10^{-2} m; wavenumber = 16.7 m^{-1} (0.167 cm^{-1})

Q8 What is the number of photons of light with a wavelength of 4000 pm that provide 1 J of energy?

SOLUTION

Given: lambda = 4000 pm = 4000×10^{-12} m = 4×10^{-9} m; h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}

Energy per photon: E = hc / lambda = $(6.626 \times 10^{-34} \times 3 \times 10^{8}) / (4 \times 10^{-9})$
E = 4.969×10^{-17} J

Number of photons: n = 1 J / (4.969×10^{-17} J) = 2.012×10^{16}

Answer: n = 2.012×10^{16} photons

Q9 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).

SOLUTION

Given: lambda = 4×10^{-7} m; phi = 2.13 eV = $2.13 \times 1.6020 \times 10^{-19}$ J = 3.412×10^{-19} J
h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}; m_e = 9.10938×10^{-31} kg

(i) Energy of photon:
E = hc / lambda = $(6.626 \times 10^{-34} \times 3 \times 10^{8}) / (4 \times 10^{-7}) = 4.970 \times 10^{-19}$ J
E = $4.970 \times 10^{-19} / 1.6020 \times 10^{-19} = 3.102$ eV

(ii) Kinetic energy of emitted electron:
KE = E – phi = 3.102 – 2.13 = 0.972 eV = 1.557×10^{-19} J

(iii) Velocity of photoelectron:
KE = (1/2) $m_e v^2$ => v = sqrt(2 KE / m_e)
v = sqrt($2 \times 1.557 \times 10^{-19} / 9.10938 \times 10^{-31}$) = 5.847×10^{5} m s^{-1}

Answer: (i) 3.10 eV; (ii) KE = 0.972 eV (1.557×10^{-19} J); (iii) v = 5.85×10^{5} m s^{-1}

Q10 Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1}.

SOLUTION

Given: lambda = 242 nm = 242 × 10^{-9} m; h = 6.626 × 10^{-34} J s; c = 3 × 10^{8} m s^{-1}; N_A = 6.022 × 10^{23} mol^{-1}

Energy per photon: E = hc / lambda = (6.626 × 10^{-34} × 3 × 10^{8}) / (242 × 10^{-9})
E = 8.215 × 10^{-19} J per atom

Ionisation energy per mole:
IE = E × N_A = 8.215 × 10^{-19} × 6.022 × 10^{23} = 4.946 × 10^{5} J mol^{-1} = 494.6 kJ mol^{-1}

Answer: IE = 494.6 kJ mol^{-1}

Q11 A 25 watt bulb emits monochromatic yellow light of wavelength of 0.57 μm. Calculate the rate of emission of quanta per second.

SOLUTION

Given: P = 25 W (= 25 J s^{-1}); lambda = 0.57 μm = 0.57 × 10^{-6} m
h = 6.626 × 10^{-34} J s; c = 3 × 10^{8} m s^{-1}

Energy of one quantum: E = hc / lambda = (6.626 × 10^{-34} × 3 × 10^{8}) / (0.57 × 10^{-6})
E = 3.488 × 10^{-19} J

Rate of emission (quanta per second) = P / E = 25 / (3.488 × 10^{-19})
= 7.169 × 10^{19} quanta s^{-1}

Answer: Rate = 7.17 × 10^{19} quanta s^{-1}

Q12 Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (nu_0) and work function (W_0) of the metal.

SOLUTION

Given: $lambda_0$ = 6800 Å = 6800 × 10^{-10} m; h = 6.626 × 10^{-34} J s; c = 3 × 10^{8} m s^{-1}

Electrons emitted with zero velocity => the radiation has exactly the threshold wavelength.

Threshold frequency: nu_0 = c / $lambda_0$ = (3 × 10^{8}) / (6800 × 10^{-10}) = 4.412 × 10^{14}

Hz

Work function: $W_o = h \times nu_o = 6.626 \times 10^{-34} \times 4.412 \times 10^{14} = 2.923 \times 10^{-19}$ J
$W_o = 2.923 \times 10^{-19} / 1.6022 \times 10^{-19} = 1.82$ eV

Answer: $nu_o = 4.41 \times 10^{14}$ Hz; $W_o = 2.92 \times 10^{-19}$ J (1.82 eV)

Q13 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?

SOLUTION

Given: R_H = 1.097×10^7 m^{-1}; $n_1 = 2$ (lower), $n_2 = 4$ (upper)

Using Rydberg equation: $1/lambda = R_H (1/n_1^2 - 1/n_2^2)$
$1/lambda = 1.097 \times 10^7 \times (1/4 - 1/16) = 1.097 \times 10^7 \times (4/16 - 1/16)$
$1/lambda = 1.097 \times 10^7 \times 3/16 = 2.057 \times 10^6$ m^{-1}

$lambda = 1 / (2.057 \times 10^6) = 4.862 \times 10^{-7}$ m = 486.2 nm
This is the blue-green line in the Balmer series (visible region).

Answer: lambda = 486.2 nm (Balmer series, visible blue-green line)

Q14 How much energy is required to ionise a H atom if the electron occupies n = 5 orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from n = 1 orbit).

SOLUTION

Given: $E_n = -2.18 \times 10^{-18} / n^2$ J $atom^{-1}$

Energy at n = 5: $E_5 = -2.18 \times 10^{-18} / 25 = -8.72 \times 10^{-20}$ J

Energy to ionise from n = 5 (to n = inf): $dE = 0 - E_5 = +8.72 \times 10^{-20}$ J

Energy to ionise from n = 1 (ground state): $dE_1 = 0 - E_1 = 2.18 \times 10^{-18}$ J

Ratio: IE(n=1) / IE(n=5) = $2.18 \times 10^{-18} / 8.72 \times 10^{-20} = 25$
So the ionisation energy from n = 5 is 25 times smaller than from n = 1.

Answer: IE from n=5 = 8.72×10^{-20} J $atom^{-1}$; this is 1/25 of IE from n=1 (2.18×10^{-18} J)

Q15 What is the maximum number of emission lines when the excited electron of a H atom in n = 6 drops to the ground state?

SOLUTION

When an electron in H atom drops from n = 6 to the ground state (n = 1), it can make transitions between any two energy levels from n = 6 to n = 1.

Maximum number of spectral lines = $n(n-1)/2 = 6 \times 5 / 2 = 15$

These transitions are: 6->5, 6->4, 6->3, 6->2, 6->1 (5 lines from n=6); 5->4, 5->3, 5->2, 5->1 (4 lines from n=5); and so on down to 2->1.

Answer: Maximum 15 emission lines

Q16 (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.

SOLUTION

Given: $E_1 = -2.18 \times 10^{-18}$ J atom^{-1}; $a_0 = 52.9$ pm (Bohr radius)

(i) $E_n = E_1 / n^2$

$E_5 = -2.18 \times 10^{-18} / 5^2 = -2.18 \times 10^{-18} / 25 = -8.72 \times 10^{-20}$ J atom^{-1}

(ii) Bohr radius for nth orbit: $r_n = n^2 \times a_0$

$r_5 = 5^2 \times 52.9$ pm $= 25 \times 52.9$ pm $= 1322.5$ pm $= 13.225$ Å

Answer: (i) $E_5 = -8.72 \times 10^{-20}$ J atom^{-1}; (ii) $r_5 = 1322.5$ pm

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Q17 Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

SOLUTION

In the Balmer series, $n_1 = 2$. The longest wavelength corresponds to the smallest energy transition, which is $n_2 = 3$ -> $n_1 = 2$.

Given: $R_H = 1.097 \times 10^7$ m^{-1}

Wavenumber: $\bar{\nu} = R_H (1/2^2 - 1/3^2) = 1.097 \times 10^7 \times (1/4 - 1/9)$
$= 1.097 \times 10^7 \times 5/36 = 1.097 \times 10^7 \times 0.1389 = 1.524 \times 10^6$ m^{-1}

$\lambda = 1 / (1.524 \times 10^6) = 6.56 \times 10^{-7}$ m $= 656$ nm (red H-alpha line)

Answer: Wavenumber = 1.524×10^6 m^{-1} ($\lambda = 656$ nm, red H-alpha line)

Q18 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.

SOLUTION

Given: $E_1 = -2.18 \times 10^{-11}$ erg $= -2.18 \times 10^{-18}$ J (1 erg $= 10^{-7}$ J)
$h = 6.626 \times 10^{-34}$ J s; $c = 3 \times 10^{8}$ m s^{-1}

$E_n = E_1 / n^2$
$E_1 = -2.18 \times 10^{-11}$ erg $= -2.18 \times 10^{-18}$ J
$E_5 = -2.18 \times 10^{-11} / 25 = -8.72 \times 10^{-13}$ erg $= -8.72 \times 10^{-20}$ J

Energy required to shift from n=1 to n=5:
$dE = E_5 - E_1 = -8.72 \times 10^{-13} - (-2.18 \times 10^{-11})$ erg
$dE = 2.0928 \times 10^{-11}$ erg $= 2.0928 \times 10^{-18}$ J

Wavelength of light emitted on return to ground state (n=5 -> n=1):
$\lambda = hc / dE = (6.626 \times 10^{-34} \times 3 \times 10^{8}) / (2.0928 \times 10^{-18})$
$\lambda = 9.498 \times 10^{-8}$ m = 94.98 nm (UV, Lyman series)

Answer: Energy required = 2.093×10^{-18} J; wavelength emitted = 9.498×10^{-8} m (94.98 nm)

Q19 The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18}) / n^2$ J. Calculate the energy required to remove an electron completely from the n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition?

SOLUTION

Given: $E_n = -2.18 \times 10^{-18} / n^2$ J; $h = 6.626 \times 10^{-34}$ J s; $c = 3 \times 10^{8}$ m s^{-1}

$E_2 = -2.18 \times 10^{-18} / 4 = -5.45 \times 10^{-19}$ J
$E_{inf} = 0$ J

Energy to ionise from n = 2: $dE = E_{inf} - E_2 = 5.45 \times 10^{-19}$ J

Longest wavelength of light that can cause this transition:
$\lambda = hc / dE = (6.626 \times 10^{-34} \times 3 \times 10^{8}) / (5.45 \times 10^{-19})$
$\lambda = 3.647 \times 10^{-7}$ m $= 3.647 \times 10^{-5}$ cm

Answer: Energy required = 5.45 × 10^{-19} J; longest wavelength = 3.647 × 10^{-5} cm (364.7 nm)

Q20 Calculate the wavelength of an electron moving with a velocity of 2.05 × 10^{7} m s^{-1}.

SOLUTION

Given: v = 2.05 × 10^{7} m s^{-1}; m_e = 9.1 × 10^{-31} kg; h = 6.626 × 10^{-34} J s

de Broglie wavelength: lambda = h / (m_e × v)
lambda = 6.626 × 10^{-34} / (9.1 × 10^{-31} × 2.05 × 10^{7})
lambda = 6.626 × 10^{-34} / (1.866 × 10^{-23})
lambda = 3.552 × 10^{-11} m = 0.3552 Å

Answer: lambda = 3.55 × 10^{-11} m (0.355 Å)

Q21 The mass of an electron is 9.1 × 10^{-31} kg. If its K.E. is 3.0 × 10^{-25} J, calculate its wavelength.

SOLUTION

Given: m_e = 9.1 × 10^{-31} kg; KE = 3.0 × 10^{-25} J; h = 6.626 × 10^{-34} J s

From KE = (1/2) m_e v^2:
v = sqrt(2 KE / m_e) = sqrt(2 × 3.0 × 10^{-25} / 9.1 × 10^{-31})
v = sqrt(6.593 × 10^{5}) = 8.12 × 10^{2} m s^{-1}

de Broglie wavelength: lambda = h / (m_e v)
lambda = 6.626 × 10^{-34} / (9.1 × 10^{-31} × 812)
lambda = 6.626 × 10^{-34} / 7.389 × 10^{-28} = 8.97 × 10^{-7} m
lambda = 897 nm

Answer: lambda = 8.97 × 10^{-7} m (897 nm)

Q22 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.

SOLUTION

Count electrons in each species (electron count = Z – charge for cations, Z + |charge| for anions):

Na^+ (Z=11): 11 – 1 = 10 electrons

K^+ (Z=19): 19 – 1 = 18 electrons
Mg^{2+} (Z=12): 12 – 2 = 10 electrons
Ca^{2+} (Z=20): 20 – 2 = 18 electrons
S^{2-} (Z=16): 16 + 2 = 18 electrons
Ar (Z=18): 18 electrons

Isoelectronic pairs (10 electrons): Na^+ and Mg^{2+}
Isoelectronic group (18 electrons): K^+, Ca^{2+}, S^{2-}, and Ar

Answer: Na^+ and Mg^{2+} are isoelectronic (10 e each); K^+, Ca^{2+}, S^{2-}, and Ar are isoelectronic (18 e each)

Q23 (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] $4s_2\ 3d_1$.

SOLUTION

(i) Electronic configurations:
(a) H^- (Z=1, gains 1 e -> 2 electrons): $1s_2$
(b) Na^+ (Z=11, loses 1 e -> 10 electrons): $1s_2\ 2s_2\ 2p_6$
(c) O^{2-} (Z=8, gains 2 e -> 10 electrons): $1s_2\ 2s_2\ 2p_6$
(d) F^- (Z=9, gains 1 e -> 10 electrons): $1s_2\ 2s_2\ 2p_6$

(ii) Atomic numbers:
(a) $3s_1$: configuration is ... $3s_1$ -> Na, Z = 11
(b) $2p_3$: configuration is $1s_2\ 2s_2\ 2p_3$ -> N, Z = 7
(c) $3p_5$: configuration is [Ne] $3s_2\ 3p_5$ -> Cl, Z = 17

(iii) Atoms:
(a) [He] $2s_1$ -> Li, Z = 3
(b) [Ne] $3s_2\ 3p_3$ -> P, Z = 15
(c) [Ar] $4s_2\ 3d_1$ -> Sc, Z = 21

Answer: (i) H^-: $1s_2$; Na^+: $1s_2\ 2s_2\ 2p_6$; O^{2-}: $1s_2\ 2s_2\ 2p_6$; F^-: $1s_2\ 2s_2\ 2p_6$. (ii) Z = 11 (Na); Z = 7 (N); Z = 17 (Cl). (iii) Li; P; Sc

Q24 What is the lowest value of n that allows g orbitals to exist?

SOLUTION

For a g orbital, the azimuthal quantum number l = 4.
The condition is: l can range from 0 to (n – 1).
For l = 4: n – 1 >= 4 => n >= 5.

Therefore, the minimum value of n for g orbitals is n = 5.

Answer: n = 5

Q25 An electron is in one of the 3d orbitals. Give the possible values of n, l and m_l for this electron.

SOLUTION

For the 3d orbital:
n = 3 (principal quantum number)
l = 2 (for d orbital, l = 2)
m_l can be: -2, -1, 0, +1, +2 (2l + 1 = 5 values)

So n = 3, l = 2, m_l ∈ {-2, -1, 0, +1, +2}

Answer: n = 3, l = 2, m_l = -2, -1, 0, +1, or +2

Q26 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.

SOLUTION

For a neutral atom: number of protons = number of electrons = 29.
(i) Number of protons = 29.

Element with Z = 29 is Copper (Cu). Mass number A = 29 + 35 = 64.

(ii) Electronic configuration of Cu (Z = 29):
Expected: [Ar] $3d^9\ 4s^2$, but Cu shows anomalous configuration due to extra stability of half-filled and fully-filled subshells.
Actual: [Ar] $3d^{10}\ 4s^1$
i.e., $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

Answer: (i) 29 protons; (ii) Cu: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$ (anomalous, fully-filled 3d)

Q27 Give the number of electrons in the species H^{2+}, H_2 and O^{2+}.

SOLUTION

H^{2+} : formed from H_2 by removal of one electron. H_2 has 2 electrons. H^{2+} has 2 – 1 = 1 electron.

H_2 : each H has 1 electron, so H_2 has 1 + 1 = 2 electrons.

O^{2+} : each O atom has 8 electrons, so O_2 has 16 electrons. O^{2+} has 16 – 1 = 15 electrons.

Answer: H^{2+}: 1 electron; H_2: 2 electrons; O^{2+}: 15 electrons

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Q28 (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

SOLUTION

(i) For n = 3, l can be 0, 1, or 2.
l = 0 (3s): m_l = 0
l = 1 (3p): m_l = -1, 0, +1
l = 2 (3d): m_l = -2, -1, 0, +1, +2

(ii) For 3d orbital: n = 3, l = 2.
m_l values: -2, -1, 0, +1, +2

(iii) Possible orbitals:
1p: Not possible (for n=1, l can only be 0; l=1 requires n >= 2)
2s: Possible (n=2, l=0 is valid)
2p: Possible (n=2, l=1 is valid since l < n)
3f: Not possible (f means l=3, requires n >= 4; n=3 allows l only up to 2)

Answer: (i) l = 0,1,2; m_l as listed. (ii) l=2, m_l = -2,-1,0,+1,+2. (iii) 2s and 2p are possible; 1p and 3f are not possible.

Q29 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.

SOLUTION

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

Answer: (a) 1s; (b) 3p; (c) 4d; (d) 4f

Q30 Explain, giving reasons, which of the following sets of quantum numbers are not possible. (a) n=0, l=0, m_l=0, m_s=+1/2 (b) n=1, l=0, m_l=0, m_s=-1/2 (c) n=1, l=1, m_l=0, m_s=+1/2 (d) n=2, l=1, m_l=0, m_s=-1/2 (e) n=3, l=3, m_l=-3, m_s=+1/2 (f) n=3, l=1, m_l=0, m_s=+1/2

SOLUTION

(a) n=0, l=0, m_l=0, m_s=+1/2: NOT POSSIBLE. n must be a positive integer (n >= 1). n=0 is not allowed.

(b) n=1, l=0, m_l=0, m_s=-1/2: POSSIBLE. n=1, l=0 (valid since l < n), m_l=0 (valid for l=0), m_s=-1/2 (valid). This describes the 1s orbital.

(c) n=1, l=1, m_l=0, m_s=+1/2: NOT POSSIBLE. For n=1, l can only be 0 (l must be < n). l=1 is not allowed when n=1.

(d) n=2, l=1, m_l=0, m_s=-1/2: POSSIBLE. n=2, l=1 (valid: l < n), m_l=0 (valid: |m_l| <= l), m_s=-1/2 (valid). This is a 2p electron.

(e) n=3, l=3, m_l=-3, m_s=+1/2: NOT POSSIBLE. For n=3, l can be 0, 1, or 2 (l must be < n). l=3 requires n >= 4.

(f) n=3, l=1, m_l=0, m_s=+1/2: POSSIBLE. n=3, l=1 (valid), m_l=0 (valid), m_s=+1/2 (valid). This is a 3p electron.

Answer: Not possible: (a) n cannot be 0; (c) l=1 requires n>=2; (e) l=3 requires n>=4. Possible: (b), (d), (f).

Q31 How many electrons in an atom may have the following quantum numbers? (a) n=4, m_s=-1/2 (b) n=3, l=0

SOLUTION

(a) n=4, m_s=-1/2:
Total electrons in n=4 shell = $2n^2 = 2 \times 16 = 32$.
Half of them have m_s = -1/2: 32/2 = 16 electrons.

(b) n=3, l=0 (3s subshell):
For l=0, m_l = 0 only. Each m_l value holds 2 electrons (m_s = +1/2 and -1/2).
Number of electrons = 2.

Answer: (a) 16 electrons; (b) 2 electrons

Q32 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.

SOLUTION

For the nth Bohr orbit, the angular momentum is quantised:
$m_e v r_n = n h / (2 \pi)$ ... (1)

The de Broglie wavelength of the electron: $\lambda = h / (m_e v)$ ... (2)

From (2): $m_e v = h / \lambda$. Substituting in (1):
$(h / \lambda) \times r_n = n h / (2 \pi)$
$r_n / \lambda = n / (2 \pi)$
$2 \pi r_n = n \lambda$

The circumference of the nth Bohr orbit is $2 \pi r_n = n \lambda$.
This shows the circumference is exactly n times the de Broglie wavelength — an integral multiple.

Answer: Circumference $2 \pi r_n = n \times \lambda$; the Bohr orbit contains exactly n de Broglie wavelengths.

Q33 What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 of He^+ spectrum?

SOLUTION

For He^+ (Z=2), the Balmer transition n=4 -> n=2:
$1/\lambda = R_H \times Z^2 \times (1/n_1^2 - 1/n_2^2) = R_H \times 4 \times (1/4 - 1/16)$
$= R_H \times 4 \times 3/16 = (3/4)\, R_H$

For H (Z=1), we need: $R_H \times (1/n_1^2 - 1/n_2^2) = (3/4)\, R_H$
$\Rightarrow 1/n_1^2 - 1/n_2^2 = 3/4 = 1/1 - 1/4 = 1 - 1/4$
$\Rightarrow n_1 = 1, n_2 = 2$

The transition n = 2 -> n = 1 in hydrogen (Lyman series) has the same wavelength.

Answer: Transition n=2 -> n=1 in hydrogen (Lyman series, UV region)

Q34 Calculate the energy required for the process He+(g) -> He_2+(g) + e-. The ionization energy for the H atom in the ground state is 2.18×10^{-18} J $atom^{-1}$.

SOLUTION

Given: IE(H, ground state) = 2.18×10^{-18} J $atom^{-1}$

For a hydrogen-like ion, IE scales as Z^2:
$IE(He^+) = Z^2 \times IE(H) = (2)^2 \times 2.18 \times 10^{-18}$ J
$IE(He^+) = 4 \times 2.18 \times 10^{-18} = 8.72 \times 10^{-18}$ J $atom^{-1}$

Energy per mole: $IE(He^+) = 8.72 \times 10^{-18} \times 6.022 \times 10^{23} / 1000 = 5250$ kJ mol^{-1}

Answer: $IE(He^+) = 8.72 \times 10^{-18}$ J $atom^{-1}$ (5250 kJ mol^{-1})

Q35 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.

SOLUTION

Given: diameter d = 0.15 nm = 0.15×10^{-9} m; length L = 20 cm = 0.20 m

Number of atoms = L / d = $0.20 / (0.15 \times 10^{-9}) = 1.33 \times 10^9$

Answer: 1.33×10^9 carbon atoms

Q36 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.

SOLUTION

Given: $n = 2 \times 10^8$ atoms; total length $L = 2.4\ cm = 2.4 \times 10^{-2}\ m$

Diameter of one atom: $d = L / n = (2.4 \times 10^{-2}) / (2 \times 10^8) = 1.2 \times 10^{-10}\ m$

Radius $= d / 2 = 0.6 \times 10^{-10}\ m = 6.0 \times 10^{-11}\ m = 60\ pm = 0.60\ Å$

Answer: Radius $= 6.0 \times 10^{-11}$ m (60 pm or 0.60 Å)

Q37 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.

SOLUTION

Given: diameter $d = 2.6\ Å = 2.6 \times 10^{-10}\ m$; $1\ Å = 100\ pm$

(a) Radius $= d / 2 = 2.6 / 2 = 1.3\ Å = 1.3 \times 100\ pm = 130\ pm$

(b) Length $L = 1.6\ cm = 1.6 \times 10^{-2}\ m$

Number of atoms $= L / d = (1.6 \times 10^{-2}) / (2.6 \times 10^{-10}) = 6.15 \times 10^7$

Answer: (a) radius = 130 pm; (b) 6.15×10^7 atoms

Q38 A certain particle carries 2.5×10^{-16} C of static electric charge. Calculate the number of electrons present in it.

SOLUTION

Given: $q = 2.5 \times 10^{-16}\ C$; charge of one electron $e = 1.6022 \times 10^{-19}\ C$

Number of electrons $= q / e = (2.5 \times 10^{-16}) / (1.6022 \times 10^{-19}) = 1560.4 \approx 1560$

Answer: 1560 electrons

Q39 In Millikan'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.

SOLUTION

Given: q = 1.282 × 10^{-18} C (magnitude); e = 1.6022 × 10^{-19} C

Number of electrons = q / e = (1.282 × 10^{-18}) / (1.6022 × 10^{-19}) = 8.00 ≈ 8

Answer: 8 electrons

Q40 In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the alpha-particles. If the thin foil of light atoms like aluminium etc. is used, what difference would be observed from the above results?

SOLUTION

In Rutherford's experiment, heavy atoms like gold are used because their nuclei are large and carry high positive charge (Z_{Au} = 79), causing significant deflection of alpha-particles.

If a thin foil of a light element like aluminium (Z_{Al} = 13) were used:
1. The nuclear charge (+13 for Al vs +79 for Au) is much smaller, so the Coulomb repulsion on alpha-particles is weaker.
2. The nucleus of Al is smaller than that of Au.
3. As a result, far fewer alpha-particles would be deflected at large angles.
4. Most alpha-particles would pass through with very small or no deflection.
5. The observation would still confirm the nuclear model (most of atom is empty space), but the scattering pattern would show many fewer large-angle deflections, making it harder to precisely determine nuclear size.

Conclusion: Fewer large-angle scatterings would be observed; the basic nuclear model would still hold but with less pronounced effect.

Answer: Fewer alpha-particles would be deflected at large angles due to smaller nuclear charge and size of Al; the experiment would still support the nuclear model but with much less pronounced scattering.

Q41 Symbols 79/35 Br and 79Br can be written, whereas symbols 35/79 Br and 35Br are not acceptable. Answer briefly.

SOLUTION

Standard nuclide notation places the mass number (A) as a superscript at the top-left and the atomic number (Z) as a subscript at the bottom-left of the element symbol: (A on top, Z below) X.

79/35 Br (A = 79 on top, Z = 35 below): acceptable. It follows the convention exactly — mass number 79 as the top-left superscript and atomic number 35 as the bottom-left subscript.

79Br (A = 79 on top-left, Z omitted): acceptable. The atomic number is fixed by the element symbol itself — every bromine atom has Z = 35 — so writing Z is optional. Giving only the mass number still identifies the isotope unambiguously.

35/79 Br (35 on top, 79 below): not acceptable. The mass number and atomic number are in the wrong positions (swapped). Placing 35 as the mass number and 79 as the atomic number is inconsistent with bromine (Z of Br must be 35, and A must exceed Z).

35Br (35 as top-left superscript, Z omitted): not acceptable. Written this way it means mass number 35, which is impossible for bromine (a bromine nucleus has at least 44 neutrons, so A cannot be 35 = Z). Only the atomic number can be dropped, never the mass number, and it cannot be mis-placed as the mass number.

Key point: the mass number always occupies the top-left position and the atomic number the bottom-left; the atomic number may be omitted (implied by the symbol), but the mass number may not be omitted or written in the wrong place.

Answer: 79/35 Br is correct (A top-left, Z bottom-left) and 79Br is allowed because Z = 35 is implied by the symbol Br. 35/79 Br is wrong (A and Z positions swapped) and 35Br is wrong (35 read as the mass number is impossible for Br). The mass number's top-left position is mandatory; only the atomic number may be dropped.

Q42 An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

SOLUTION

Let number of protons = p (= Z) and neutrons = n.
Given: n = p + 0.317p = 1.317p
Also: p + n = 81 => p + 1.317p = 81 => 2.317p = 81 => p = 81 / 2.317 = 34.96 ≈ 35

Z = 35, neutrons = 81 – 35 = 46.
Check: 46 is (46–35)/35 × 100 = 31.4% more than 35 (≈ 31.7%). Confirmed.

Element with Z = 35 is Bromine (Br).
Symbol: 81/35 Br

Answer: 81/35 Br (Bromine)

Q43 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.

SOLUTION

One unit of negative charge: number of electrons = number of protons + 1 = p + 1.
Let protons = p. Electrons = p + 1. Neutrons = n.

Given: n = 1.111 × (p + 1)
Also: p + n = 37 => p + 1.111(p+1) = 37 => p + 1.111p + 1.111 = 37 => 2.111p = 35.889 => p = 17

Z = 17, electrons = 18, neutrons = 37 – 17 = 20.
Check: 20/18 = 1.111 => 11.1% more neutrons than electrons. Confirmed.

Element with Z = 17 is Chlorine (Cl). The ion has one negative charge.
Symbol: $^{37}Cl^-$ (chloride ion)

Answer: $^{37}_{17}Cl^-$ (chloride ion)

Q44 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.

SOLUTION

Three units of positive charge: electrons = protons – 3 = p – 3.
Let protons = p. Electrons = p – 3. Neutrons = n.

Given: n = 1.304 × (p – 3)
Also: p + n = 56 => p + 1.304(p – 3) = 56 => p + 1.304p – 3.912 = 56 => 2.304p = 59.912 => p = 26

Z = 26, electrons = 23, neutrons = 56 – 26 = 30.
Check: 30/23 = 1.304 => 30.4% more neutrons than electrons. Confirmed.

Element with Z = 26 is Iron (Fe). The ion has 3 positive charges.
Symbol: 56/26 Fe^{3+} (ferric ion)

Answer: 56/26 Fe^{3+} (iron(III) ion)

Q45 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.

SOLUTION

Frequency ranges in the electromagnetic spectrum (approximate):
FM radio: ~10^{8} Hz (lowest frequency)
Microwave: ~10^{10} – 10^{12} Hz
Amber light (visible): ~5 × 10^{14} Hz
X-rays: ~10^{17} – 10^{19} Hz
Cosmic rays: ~10^{20} – 10^{24} Hz (highest frequency)

Increasing order of frequency:
FM radio < Microwave < Amber light < X-rays < Cosmic rays
(c) < (a) < (b) < (e) < (d)

Answer: FM radio (c) < Microwave (a) < Amber light (b) < X-rays (e) < Cosmic rays (d)

Q46 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.

SOLUTION

Given: lambda = 337.1 nm = 337.1×10^{-9} m; n = 5.6×10^{24} photons
h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}

Energy per photon: E = hc / lambda = $(6.626 \times 10^{-34} \times 3 \times 10^{8}) / (337.1 \times 10^{-9})$
E = 5.897×10^{-19} J

Total energy emitted: E_total = n × E = $5.6 \times 10^{24} \times 5.897 \times 10^{-19}$
E_total = 3.302×10^{6} J

Answer: Total energy (power output) = 3.30×10^{6} J

Q47 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 and (d) number of quanta present if it produces 2 J of energy.

SOLUTION

Given: lambda = 616 nm = 616×10^{-9} m; h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}

(a) Frequency: nu = c / lambda = $(3 \times 10^{8}) / (616 \times 10^{-9}) = 4.870 \times 10^{14}$ Hz

(b) Distance in 30 s = c × t = $3 \times 10^{8} \times 30 = 9.0 \times 10^{9}$ m

(c) Energy of one quantum: E = h × nu = $6.626 \times 10^{-34} \times 4.870 \times 10^{14} = 3.227 \times 10^{-19}$ J

(d) Number of quanta for 2 J: n = $2 / (3.227 \times 10^{-19}) = 6.198 \times 10^{18}$

Answer: (a) 4.87×10^{14} Hz; (b) 9.0×10^{9} m; (c) 3.22×10^{-19} J; (d) 6.20×10^{18} quanta

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Q48 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.

SOLUTION

Given: E_total = 3.15×10^{-18} J; lambda = 600 nm = 600×10^{-9} m
h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}

Energy per photon: E = hc / lambda = $(6.626 \times 10^{-34} \times 3 \times 10^{8}) / (600 \times 10^{-9})$

$E = 3.313 \times 10^{-19}$ J

Number of photons: $n = E_{total} / E = (3.15 \times 10^{-18}) / (3.313 \times 10^{-19}) \approx 9.51 \approx 9.5$

Answer: Approximately 9 to 10 photons (9.51 using given data)

Q49 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.

SOLUTION

Given: $t = 2$ ns $= 2 \times 10^{-9}$ s; $n = 2.5 \times 10^{15}$ photons; $h = 6.626 \times 10^{-34}$ J s

The minimum frequency associated with a pulse of duration t is given by $\nu = 1/t$:
$\nu = 1 / (2 \times 10^{-9}) = 5.0 \times 10^{8}$ Hz

Energy per photon: $E = h \times \nu = 6.626 \times 10^{-34} \times 5.0 \times 10^{8} = 3.313 \times 10^{-25}$ J

Total energy of the source: $E_{total} = n \times E = 2.5 \times 10^{15} \times 3.313 \times 10^{-25}$
$E_{total} = 8.28 \times 10^{-10}$ J

Answer: Energy of the source = 8.28×10^{-10} J

Q50 The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

SOLUTION

Given: $\lambda_1 = 589$ nm $= 589 \times 10^{-9}$ m; $\lambda_2 = 589.6$ nm $= 589.6 \times 10^{-9}$ m
$h = 6.626 \times 10^{-34}$ J s; $c = 3 \times 10^{8}$ m s^{-1}

Frequency of first transition:
$\nu_1 = c / \lambda_1 = (3 \times 10^{8}) / (589 \times 10^{-9}) = 5.093 \times 10^{14}$ Hz

Frequency of second transition:
$\nu_2 = c / \lambda_2 = (3 \times 10^{8}) / (589.6 \times 10^{-9}) = 5.088 \times 10^{14}$ Hz

Energy difference between the two excited states:
$dE = h \times (\nu_1 - \nu_2) = 6.626 \times 10^{-34} \times (5.093 - 5.088) \times 10^{14}$
$dE = 6.626 \times 10^{-34} \times 5.18 \times 10^{11} = 3.43 \times 10^{-22}$ J

Answer: $\nu_1 = 5.093 \times 10^{14}$ Hz; $\nu_2 = 5.088 \times 10^{14}$ Hz; $dE = 3.43 \times 10^{-22}$ J

Q51 The work function for caesium atom is 1.9 eV. Calculate (a) the threshold wavelength and (b) the threshold frequency of the radiation. If the caesium element is irradiated with a wavelength 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

SOLUTION

Given: phi = 1.9 eV = $1.9 \times 1.6022 \times 10^{-19} = 3.044 \times 10^{-19}$ J
h = 6.626×10^{-34} J s; c = 3×10^{8} m s^{-1}; m_e = 9.10938×10^{-31} kg

(a) Threshold wavelength: $\lambda_0 = hc / \phi$
$\lambda_0 = (6.626 \times 10^{-34} \times 3 \times 10^{8}) / (3.044 \times 10^{-19}) = 6.53 \times 10^{-7}$ m = 653 nm

(b) Threshold frequency: $\nu_0 = \phi / h = (3.044 \times 10^{-19}) / (6.626 \times 10^{-34}) = 4.596 \times 10^{14}$ Hz

Irradiation at λ = 500 nm = 500×10^{-9} m:
$E_{photon} = hc / \lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8}) / (500 \times 10^{-9}) = 3.976 \times 10^{-19}$ J
KE = $E_{photon} - \phi = 3.976 \times 10^{-19} - 3.044 \times 10^{-19} = 9.31 \times 10^{-20}$ J

Velocity: $v = \sqrt{2\,KE / m_e} = \sqrt{2 \times 9.31 \times 10^{-20} / 9.10938 \times 10^{-31}}$
$v = \sqrt{2.044 \times 10^{11}} = 4.52 \times 10^{5}$ m s^{-1}

Answer: (a) λ_0 = 653 nm; (b) $\nu_0 = 4.60 \times 10^{14}$ Hz; KE = 9.31×10^{-20} J; v = 4.52×10^{5} m s^{-1}

Q52 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 × 10^{-5} (m s^{-1}): 2.55, 4.35, 5.35

SOLUTION

Given: m_e = 9.1×10^{-31} kg; c = 3×10^{8} m s^{-1}
λ_1 = 500 nm, v_1 = 2.55×10^{5} m s^{-1}
λ_2 = 450 nm, v_2 = 4.35×10^{5} m s^{-1}
λ_3 = 400 nm, v_3 = 5.35×10^{5} m s^{-1}

Kinetic energies:
$KE_1 = 0.5 \times 9.1 \times 10^{-31} \times (2.55 \times 10^{5})^2 = 2.96 \times 10^{-20}$ J
$KE_3 = 0.5 \times 9.1 \times 10^{-31} \times (5.35 \times 10^{5})^2 = 1.302 \times 10^{-19}$ J

Frequencies:
$\nu_1 = c / \lambda_1 = 3 \times 10^{8} / 500 \times 10^{-9} = 6.00 \times 10^{14}$ Hz
$\nu_3 = c / \lambda_3 = 3 \times 10^{8} / 400 \times 10^{-9} = 7.50 \times 10^{14}$ Hz

(b) Planck's constant using pair (1,3):
$h = (KE_3 - KE_1) / (nu_3 - nu_1) = (1.302 \times 10^{-19} - 2.96 \times 10^{-20}) / (7.50 - 6.00) \times 10^{14}$
$h = 1.006 \times 10^{-19} / 1.50 \times 10^{14} = 6.71 \times 10^{-34}$ J s

(a) Threshold frequency: $nu_0 = nu_1 - KE_1/h = 6.00 \times 10^{14} - (2.96 \times 10^{-20})/(6.71 \times 10^{-34})$
$nu_0 = 6.00 \times 10^{14} - 4.41 \times 10^{13} = 5.56 \times 10^{14}$ Hz
$lambda_0 = c / nu_0 = 3 \times 10^8 / 5.56 \times 10^{14} = 5.40 \times 10^{-7}$ m = 540 nm

Answer: (a) threshold wavelength ~540 nm; (b) $h \approx 6.71 \times 10^{-34}$ J s

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Q53 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.

SOLUTION

Given: lambda = 256.7 nm = 256.7×10^{-9} m; stopping potential $V_0 = 0.35$ V
$h = 6.626 \times 10^{-34}$ J s; $c = 3 \times 10^8$ m s^{-1}; $e = 1.6022 \times 10^{-19}$ C

Energy of incident photon:
$E = hc / lambda = (6.626 \times 10^{-34} \times 3 \times 10^8) / (256.7 \times 10^{-9}) = 7.744 \times 10^{-19}$ J
$= 7.744 \times 10^{-19} / 1.6022 \times 10^{-19} = 4.833$ eV

KE of ejected electron = $eV_0 = 1.6022 \times 10^{-19} \times 0.35 = 5.608 \times 10^{-20}$ J = 0.35 eV

Work function: phi = E – KE = $7.744 \times 10^{-19} - 5.608 \times 10^{-20} = 7.183 \times 10^{-19}$ J
$phi = 7.183 \times 10^{-19} / 1.6022 \times 10^{-19} = 4.483$ eV

Answer: Work function of silver = 7.18×10^{-19} J (4.48 eV)

Q54 If the photon of the wavelength 150 pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of 1.5×10^7 m s^{-1}, calculate the energy with which it is bound to the nucleus.

SOLUTION

Given: lambda = 150 pm = 150×10^{-12} m; $v = 1.5 \times 10^7$ m s^{-1}
$h = 6.626 \times 10^{-34}$ J s; $c = 3 \times 10^8$ m s^{-1}; $m_e = 9.10938 \times 10^{-31}$ kg

Energy of incident photon:
$E_{photon} = hc / lambda = (6.626 \times 10^{-34} \times 3 \times 10^8) / (150 \times 10^{-12}) = 1.325 \times 10^{-15}$ J

Kinetic energy of ejected electron:
KE = (1/2) m_e v^2 = 0.5 × 9.10938 × 10^{-31} × (1.5 × 10^7)^2 = 1.025 × 10^{-16} J

Binding energy = E_photon – KE = 1.325 × 10^{-15} – 1.025 × 10^{-16}
= 1.323 × 10^{-15} – 0.1025 × 10^{-15} = 1.223 × 10^{-15} J

Answer: Binding energy = 1.22 × 10^{-15} J

Q55 Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as v = 3.29 × 10^{15} (Hz) [1/3^2 – 1/n^2]. Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.

SOLUTION

Given: lambda = 1285 nm = 1285 × 10^{-9} m; c = 3 × 10^8 m s^{-1}

nu = c / lambda = (3 × 10^8) / (1285 × 10^{-9}) = 2.335 × 10^{14} Hz

Using the given formula: nu = 3.29 × 10^{15} × [1/9 – 1/n^2]
2.335 × 10^{14} = 3.29 × 10^{15} × [1/9 – 1/n^2]
1/9 – 1/n^2 = 2.335 × 10^{14} / 3.29 × 10^{15} = 0.07098
1/n^2 = 1/9 – 0.07098 = 0.11111 – 0.07098 = 0.04013
n^2 = 1/0.04013 = 24.92 ≈ 25
n = 5

The transition is from n = 5 to n = 3 in the Paschen series.
lambda = 1285 nm is in the infrared (IR) region of the spectrum.

Answer: n = 5; the spectral line at 1285 nm is in the infrared region (Paschen series)

Q56 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.

SOLUTION

Given: r_1 = 1.3225 nm = 1322.5 pm; r_2 = 211.6 pm; a_0 = 52.9 pm

Bohr radius formula: r_n = n^2 × a_0
n_1^2 = r_1 / a_0 = 1322.5 / 52.9 = 25.0 => n_1 = 5
n_2^2 = r_2 / a_0 = 211.6 / 52.9 = 4.0 => n_2 = 2

Transition: n = 5 -> n = 2 (Balmer series)

Using Rydberg equation: 1/lambda = R_H (1/n_2^2 – 1/n_1^2)

$1/\lambda = 1.097 \times 10^7 \times (1/4 - 1/25) = 1.097 \times 10^7 \times 21/100$
$1/\lambda = 1.097 \times 10^7 \times 0.21 = 2.304 \times 10^6\ m^{-1}$
$\lambda = 4.341 \times 10^{-7}\ m = 434.1\ nm$

This belongs to the Balmer series; 434 nm is in the visible (violet) region.

Answer: lambda = 434.1 nm; Balmer series, visible (violet) region

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Q57 Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is $1.6 \times 10^6\ m\ s^{-1}$, calculate de Broglie wavelength associated with this electron.

SOLUTION

Given: $v = 1.6 \times 10^6\ m\ s^{-1}$; $m_e = 9.10938 \times 10^{-31}\ kg$; $h = 6.626 \times 10^{-34}\ J\ s$

de Broglie wavelength: $\lambda = h / (m_e \times v)$
$\lambda = 6.626 \times 10^{-34} / (9.10938 \times 10^{-31} \times 1.6 \times 10^6)$
$\lambda = 6.626 \times 10^{-34} / 1.458 \times 10^{-24}$
$\lambda = 4.55 \times 10^{-10}\ m = 0.455\ nm = 4.55\ Å$

Answer: $\lambda = 4.55 \times 10^{-10}\ m$ (0.455 nm)

Q58 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.

SOLUTION

Given: $\lambda = 800\ pm = 800 \times 10^{-12}\ m$; $m_n = 1.675 \times 10^{-27}\ kg$; $h = 6.626 \times 10^{-34}\ J\ s$

From de Broglie equation: $\lambda = h / (m_n \times v) \Rightarrow v = h / (m_n \times \lambda)$
$v = 6.626 \times 10^{-34} / (1.675 \times 10^{-27} \times 800 \times 10^{-12})$
$v = 6.626 \times 10^{-34} / 1.340 \times 10^{-36}$
$v = 494.5\ m\ s^{-1}$

Answer: $v = 494.5\ m\ s^{-1}$ (~495 $m\ s^{-1}$)

Q59 If the velocity of the electron in Bohr's first orbit is 2.19×10^{6} m s^{-1}, calculate the de Broglie wavelength associated with it.

SOLUTION

Given: $v = 2.19 \times 10^{6}$ m s^{-1}; $m_e = 9.10938 \times 10^{-31}$ kg; $h = 6.626 \times 10^{-34}$ J s

de Broglie wavelength: $\lambda = h / (m_e \times v)$

$\lambda = 6.626 \times 10^{-34} / (9.10938 \times 10^{-31} \times 2.19 \times 10^{6})$

$\lambda = 6.626 \times 10^{-34} / 1.995 \times 10^{-24}$

$\lambda = 3.321 \times 10^{-10}$ m = 3.321 Å

Answer: $\lambda = 3.32 \times 10^{-10}$ m (3.32 Å)

Q60 The velocity associated with a proton moving in a potential difference of 1000 V is 4.37×10^{5} m s^{-1}. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

SOLUTION

Given: $v = 4.37 \times 10^{5}$ m s^{-1}; $m = 0.1$ kg; $h = 6.626 \times 10^{-34}$ J s

de Broglie wavelength for hockey ball:

$\lambda = h / (m \times v) = 6.626 \times 10^{-34} / (0.1 \times 4.37 \times 10^{5})$

$\lambda = 6.626 \times 10^{-34} / 4.37 \times 10^{4}$

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

This wavelength is immeasurably small (far smaller than an atomic nucleus), confirming that wave properties of macroscopic objects are negligible.

Answer: $\lambda = 1.52 \times 10^{-38}$ m (negligibly small; wave nature is not observable for macroscopic objects)

Q61 If the position of the electron is measured within an accuracy of +0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is h/(4 pi m × 0.05 nm), is there any problem in defining this value.

SOLUTION

Given: $\Delta x = 0.002$ nm = 2×10^{-12} m; $h = 6.626 \times 10^{-34}$ J s

By Heisenberg's uncertainty principle:

$\Delta x \times \Delta p \geq h / (4 \pi)$

$\Delta p \geq h / (4 \pi \times \Delta x) = 6.626 \times 10^{-34} / (4 \times 3.1416 \times 2 \times 10^{-12})$

delta_p >= $6.626 \times 10^{-34} / (2.513 \times 10^{-11}) = 2.636 \times 10^{-23}$ kg m s^{-1}

Stated momentum: p = h / (4 pi × 0.05 nm) = $6.626 \times 10^{-34} / (4 \times 3.1416 \times 0.05 \times 10^{-9})$
$p = 6.626 \times 10^{-34} / 6.283 \times 10^{-10} = 1.055 \times 10^{-24}$ kg m s^{-1}

Comparison: delta_p = 2.636×10^{-23} >> p = 1.055×10^{-24}
The uncertainty in momentum (2.636×10^{-23} kg m s^{-1}) is about 25 times larger than the stated momentum value.

Yes, there is a problem: the uncertainty in momentum is far greater than the momentum itself, making it impossible to define the given momentum value meaningfully.

Answer: delta_p = 2.636×10^{-23} kg m s^{-1}; this is ~25× larger than the stated momentum, so the value cannot be defined precisely.

Q62 The quantum numbers of six electrons are given below. Arrange them in order of increasing energies. If any of these combination(s) has/have the same energy lists: 1. n=4, l=2, m_l=-2, m_s=-1/2; 2. n=3, l=2, m_l=1, m_s=+1/2; 3. n=4, l=1, m_l=0, m_s=+1/2; 4. n=3, l=2, m_l=-2, m_s=-1/2; 5. n=3, l=1, m_l=-1, m_s=+1/2; 6. n=4, l=1, m_l=0, m_s=+1/2

SOLUTION

Use the (n + l) rule: lower (n + l) = lower energy; for equal (n+l), lower n = lower energy.

Electron 1: n=4, l=2 -> n+l=6, subshell 4d
Electron 2: n=3, l=2 -> n+l=5, subshell 3d
Electron 3: n=4, l=1 -> n+l=5, subshell 4p
Electron 4: n=3, l=2 -> n+l=5, subshell 3d
Electron 5: n=3, l=1 -> n+l=4, subshell 3p
Electron 6: n=4, l=1 -> n+l=5, subshell 4p

Order of n+l values: e_5 (n+l=4) < e_2,e_4 (n+l=5, n=3) < e_3,e_6 (n+l=5, n=4) < e_1 (n+l=6)

Among n+l=5 with n=3: electrons 2 and 4 are both in 3d; they have the same energy.
Among n+l=5 with n=4: electrons 3 and 6 have identical quantum numbers (n=4, l=1, m_l=0, m_s=+1/2); they have the same energy (note: two electrons cannot

have all four identical quantum numbers; the question data appears duplicated for e_3 and e_6).

Increasing order of energy: $e_5 < e_2 = e_4 < e_3 = e_6 < e_1$

Answer: Increasing energy: $e_5 < e_2 = e_4 < e_3 = e_6 < e_1$

Q63 The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electrons in 4p orbital. Which of these electrons experiences the lowest effective nuclear charge?

SOLUTION

Effective nuclear charge (Z_eff) experienced by an electron depends on shielding by inner electrons.
Electrons in higher principal quantum number shells (further from the nucleus) experience more shielding from inner electrons and thus a lower Z_eff.

2p electrons (n=2): shielded only by $1s^2$ (2 electrons)
3p electrons (n=3): shielded by all electrons in n=1, 2 shells (10 electrons) plus $3s^2$
4p electrons (n=4): shielded by all electrons in n=1, 2, 3 shells (28 electrons)

The 4p electrons are furthest from the nucleus and most shielded; they experience the lowest effective nuclear charge.

Answer: The 4p electrons experience the lowest effective nuclear charge (most shielded by inner electrons).

Q64 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.

SOLUTION

(i) 2s and 3s: The 2s orbital is closer to the nucleus and less shielded than 3s.
2s experiences a larger effective nuclear charge.

(ii) 4d and 4f: Both have n=4 but 4f has l=3 (more nodes, more diffuse, more shielded by inner electrons) compared to 4d (l=2).
4d orbital experiences a larger effective nuclear charge than 4f.

(iii) 3d and 3p: Both have n=3 but 3p (l=1) has greater penetration toward the nucleus than 3d (l=2).

3p electrons are less shielded and experience a larger effective nuclear charge than 3d.

Answer: (i) 2s; (ii) 4d; (iii) 3p

Q65 The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus?

SOLUTION

Al (Z=13): configuration [Ne] $3s_2$ $3p_1$ -> nuclear charge = 13
Si (Z=14): configuration [Ne] $3s_2$ $3p_2$ -> nuclear charge = 14

Both Al and Si have their unpaired electrons in the 3p subshell with the same shielding from inner electrons ([Ne] core = 10 electrons + $3s^2$ = 12 electrons).

Silicon has a higher nuclear charge (Z=14) than aluminium (Z=13).
The shielding is essentially the same for the 3p electron in both elements.
Therefore, the 3p electron in Si experiences a greater effective nuclear charge than in Al.

Answer: The 3p electron in Si (Z=14) experiences more effective nuclear charge than in Al (Z=13).

Q66 Indicate the number of unpaired electrons in: (a) P, (b) Si, (c) Cr, (d) Fe and (e) Kr.

SOLUTION

(a) P (Z=15): [Ne] $3s_2$ $3p_3$. Three 3p electrons occupy three separate orbitals (Hund's rule): all unpaired. Unpaired = 3

(b) Si (Z=14): [Ne] $3s_2$ $3p_2$. Two 3p electrons each occupy separate orbitals. Unpaired = 2

(c) Cr (Z=24): Anomalous configuration [Ar] $3d_5$ $4s_1$ (half-filled 3d for extra stability).
$3d^5$: 5 unpaired electrons + $4s^1$: 1 unpaired. Total unpaired = 6

(d) Fe (Z=26): [Ar] $3d_6$ $4s_2$. The $3d^6$ has 4 unpaired electrons (one pair + four singles by Hund's rule). Unpaired = 4

(e) Kr (Z=36): [Ar] $3d_{10}$ $4s_2$ $4p_6$. All subshells fully filled. Unpaired = 0

Answer: (a) P: 3; (b) Si: 2; (c) Cr: 6; (d) Fe: 4; (e) Kr: 0

Q67 (a) How many subshells are associated with n = 4? (b) How many electrons will be present in the subshells having m_s value of -1/2 for n = 4?

SOLUTION

(a) For n = 4, the possible values of l are 0, 1, 2, 3 (i.e., l = 0 to n-1).
These correspond to 4s, 4p, 4d, 4f subshells.
Total subshells = 4

(b) Total electrons in n = 4 shell = $2n^2 = 2 \times 16 = 32$ electrons.
Half of all electrons in any shell have $m_s = -1/2$.
Number of electrons with $m_s = -1/2$ in n=4 = 32/2 = 16

Answer: (a) 4 subshells (4s, 4p, 4d, 4f); (b) 16 electrons with $m_s = -1/2$