CLASS 11 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 11 Chemistry Chapter 2: Structure of Atom – Free PDF Download

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

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Chapter 2, "Structure of Atom," of the current NCERT Class 11 Chemistry (Part I) textbook builds the quantum mechanical picture of the atom — from Thomson's and Rutherford's early models through Bohr's theory, de Broglie's wave–particle duality, Heisenberg's uncertainty principle, and the quantum numbers that describe electron distribution. Below are complete, worked solutions to every end-of-chapter exercise question (2.1–2.67) from the current 2026-27 reprint of the NCERT textbook.

NCERT Exercise Solutions

Question 2.1

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

Solution: (i) Mass of one electron = 9.10939×10^{-31} kg = 9.10939×10^{-28} g.
Number of electrons in 1 g = $1 \div (9.10939 \times 10^{-28}) = 1.098 \times 10^{27}$ electrons.
(ii) Mass of 1 mole of electrons = $6.022 \times 10^{23} \times 9.10939 \times 10^{-31}$ kg = 5.48×10^{-7} kg = 5.48×10^{-4} g.
Charge on 1 mole of electrons = $6.022 \times 10^{23} \times 1.6022 \times 10^{-19}$ C = 9.65×10^{4} C (this is the Faraday constant).

Question 2.2

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

Solution: (i) CH_4 has 6 + (4 × 1) = 10 electrons per molecule. Electrons in 1 mole = $10 \times 6.022 \times 10^{23} = 6.022 \times 10^{24}$ electrons.
(ii) ^{14}C has 14 – 6 = 8 neutrons per atom. Moles of C in 7 mg = $7 \times 10^{-3}/14 = 5 \times 10^{-4}$ mol ⇒ atoms = $5 \times 10^{-4} \times 6.022 \times 10^{23} = 3.011 \times 10^{20}$.
(a) Total neutrons = $8 \times 3.011 \times 10^{20} = 2.409 \times 10^{21}$.
(b) Mass of neutrons = $2.409 \times 10^{21} \times 1.675 \times 10^{-27}$ kg = 4.036×10^{-6} kg = 4.036×10^{-3} g.
(iii) Molar mass of NH_3 = 17 g/mol. Moles in 34 mg = $34 \times 10^{-3}/17 = 2 \times 10^{-3}$ mol ⇒ molecules = $2 \times 10^{-3} \times 6.022 \times 10^{23} = 1.2044 \times 10^{21}$. Each NH_3 has 7 + 3(1) = 10 protons.
(a) Total protons = $10 \times 1.2044 \times 10^{21} = 1.2044 \times 10^{22}$.
(b) Mass of protons = $1.2044 \times 10^{22} \times 1.6726 \times 10^{-27}$ kg = 2.014×10^{-5} kg = 2.014×10^{-2} g.
The answer will NOT change with temperature and pressure, since it is based on the fixed mass/mole quantity of NH_3, not on gas volume.

Question 2.3

Question: 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: Protons = Z, Neutrons = A – Z.
$^{13}_{6}C$: protons = 6, neutrons = 13 – 6 = 7.
$^{16}_{8}O$: protons = 8, neutrons = 16 – 8 = 8.
$^{24}_{12}Mg$: protons = 12, neutrons = 24 – 12 = 12.
$^{56}_{26}Fe$: protons = 26, neutrons = 56 – 26 = 30.
$^{88}_{38}Sr$: protons = 38, neutrons = 88 – 38 = 50.

Question 2.4

Question: 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: (i) Z = 17 is chlorine ⇒ $^{35}{}_{17}Cl$.
(ii) Z = 92 is uranium ⇒ $^{233}{}_{92}U$.
(iii) Z = 4 is beryllium ⇒ $^{9}{}_{4}Be$.

Question 2.5

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

Solution: $\lambda = 580 \times 10^{-9}$ m.
$\nu = c/\lambda = (3 \times 10^{8})/(580 \times 10^{-9}) = 5.17 \times 10^{14}$ Hz.
Wavenumber = $1/\lambda = 1/(580 \times 10^{-9}$ m$) = 1.724 \times 10^{6}$ m^{-1} (= 1.724×10^{4} cm^{-1}).

Question 2.6

Question: 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: (i) $E = h\nu = 6.626 \times 10^{-34} \times 3 \times 10^{15} = 1.988 \times 10^{-18}$ J.
(ii) $\lambda = 0.50$ Å $= 5 \times 10^{-11}$ m. $E = hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(5 \times 10^{-11}) = 3.976 \times 10^{-15}$ J.

Question 2.7

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

Solution: $\nu = 1/T = 1/(2.0 \times 10^{-10}) = 5 \times 10^{9}$ Hz.
$\lambda = c/\nu = (3 \times 10^{8})/(5 \times 10^{9}) = 6 \times 10^{-2}$ m.
Wavenumber = $1/\lambda = 16.67$ m^{-1}.

Question 2.8

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

Solution: $\lambda = 4000$ pm $= 4 \times 10^{-9}$ m.
Energy of one photon = $hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(4 \times 10^{-9}) = 4.97 \times 10^{-17}$ J.
Number of photons = $1/(4.97 \times 10^{-17}) = 2.012 \times 10^{16}$ photons.

Question 2.9

Question: 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: (i) $E = hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(4 \times 10^{-7}) = 4.97 \times 10^{-19}$ J $= 4.97 \times 10^{-19}/1.602 \times 10^{-19} =$ 3.10 eV.
(ii) $KE = E_{photon} - W_0 = 3.10 - 2.13 = 0.97$ eV $= 1.554 \times 10^{-19}$ J.
(iii) $KE = ½mv^2 \Rightarrow v = \sqrt{(2KE/m)} = \sqrt{(2 \times 1.554 \times 10^{-19}/9.109 \times 10^{-31})} = 5.84 \times 10^{5}$ m/s.

Question 2.10

Question: 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: $\lambda = 242 \times 10^{-9}$ m.
E per atom $= hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(242 \times 10^{-9}) = 8.216 \times 10^{-19}$ J.
Per mole $= 8.216 \times 10^{-19} \times 6.022 \times 10^{23} = 4.948 \times 10^{5}$ J/mol = 494.8 kJ/mol.

Question 2.11

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

Solution: $\lambda = 0.57 \times 10^{-6}$ m.
Energy of one photon $= hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(0.57 \times 10^{-6}) = 3.489 \times 10^{-19}$ J.
Rate of quanta emission $= 25/(3.489 \times 10^{-19}) = 7.17 \times 10^{19}$ photons/s.

Question 2.12

Question: Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (v_0) and work function (W_0) of the metal.

Solution: $\lambda = 6800$ Å $= 6.8 \times 10^{-7}$ m.
$v_0 = c/\lambda = (3 \times 10^{8})/(6.8 \times 10^{-7}) = 4.41 \times 10^{14}$ Hz.
$W_0 = hv_0 = 6.626 \times 10^{-34} \times 4.41 \times 10^{14} = 2.92 \times 10^{-19}$ J ($\approx$ 1.83 eV).

Question 2.13

Question: 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: Using the Rydberg formula: $1/\lambda = R(1/n_1^2 - 1/n_2^2)$, $R = 1.097 \times 10^{7}$ m^{-1}.
$1/\lambda = 1.097 \times 10^{7}(1/4 - 1/16) = 1.097 \times 10^{7} \times 0.1875 = 2.057 \times 10^{6}$ m^{-1}.
$\lambda = 1/(2.057 \times 10^{6}) = 4.86 \times 10^{-7}$ m = 486 nm.

Question 2.14

Question: 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: $E_n = -2.18 \times 10^{-18}/n^2$ J. Ionization from n = 5 means $E = 2.18 \times 10^{-18}/25 = 8.72 \times 10^{-20}$ J. Ionization enthalpy from n = 1 is 2.18×10^{-18} J, which is 25 times larger than the energy needed from n = 5, since the electron in the n = 5 orbit is already far less tightly bound to the nucleus.

Question 2.15

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

Question 2.16

Question: (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: (i) $E_5 = E_1/n^2 = -2.18 \times 10^{-18}/25 = -8.72 \times 10^{-20}$ J.
(ii) $r_n = r_1 \times n^2$, where r_1 (Bohr radius) = 52.9 pm. $r_5 = 52.9 \times 25 = 1322.5$ pm.

Question 2.17

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

Solution: The longest wavelength (smallest energy gap) in the Balmer series corresponds to the $n = 3 \rightarrow n = 2$ transition.
Wavenumber = $R(1/4 - 1/9) = 1.097 \times 10^{7} \times 0.1389 = 1.524 \times 10^{6}\ m^{-1}$ (= $1.524 \times 10^{4}\ cm^{-1}$).

Question 2.18

Question: 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: $E = 2.18 \times 10^{-11}$ erg $\times (1 - 1/25) = 2.18 \times 10^{-11} \times 0.96 = 2.093 \times 10^{-11}$ erg $= 2.093 \times 10^{-18}$ J (since 1 erg = 10^{-7} J).
This is the energy absorbed to reach n = 5, and the same magnitude is emitted when the electron falls back from n = 5 to n = 1.
$\lambda = hc/E = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(2.093 \times 10^{-18}) = 9.50 \times 10^{-8}$ m ≈ 95 nm.

Question 2.19

Question: 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: Energy to remove electron from n = 2 (to n = ∞) = $2.18 \times 10^{-18}/4 = 5.45 \times 10^{-19}$ J.
$\lambda = hc/E = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(5.45 \times 10^{-19}) = 3.647 \times 10^{-7}$ m = 3.647×10^{-5} cm.

Question 2.20

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

Solution: Using de Broglie's equation: $\lambda = h/mv = (6.626 \times 10^{-34})/(9.109 \times 10^{-31} \times 2.05 \times 10^{7}) = 3.548 \times 10^{-11}$ m = 35.48 pm.

Question 2.21

Question: 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: $KE = \frac{1}{2}mv^2 \Rightarrow v = \sqrt{(2KE/m)} = \sqrt{(2 \times 3.0 \times 10^{-25}/9.1 \times 10^{-31})} = 811.7$ m/s.
$\lambda = h/mv = (6.626 \times 10^{-34})/(9.1 \times 10^{-31} \times 811.7) = 8.96 \times 10^{-7}$ m ≈ 896 nm.

Question 2.22

Question: 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: Electron counts: $Na^+ = 10$, $K^+ = 18$, $Mg^{2+} = 10$, $Ca^{2+} = 18$, $S^{2-} = 18$, Ar = 18.
Isoelectronic groups: {Na^+, Mg^{2+}} with 10 electrons each; {K^+, Ca^{2+}, S^{2-}, Ar} with 18 electrons each.

Question 2.23

Question: (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] $3s^23p^3$ (c) [Ar] $4s^23d^1$.

Solution: (i) (a) H^-: $1s^2$. (b) Na^+: $1s^22s^22p^5$. (c) O^{2-}: $1s^22s^22p^5$. (d) F^-: $1s^22s^22p^5$.
(ii) (a) $3s^1 \rightarrow$ Na, Z = 11. (b) $2p^3 \rightarrow$ N, Z = 7. (c) $3p^5 \rightarrow$ Cl, Z = 17.
(iii) (a) [He]$2s^1 \rightarrow$ Li (Z = 3). (b) [Ne]$3s^23p^3 \rightarrow$ P (Z = 15). (c) [Ar]$4s^23d^1 \rightarrow$ Sc (Z = 21).

Question 2.24

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

Solution: g orbitals correspond to l = 4. Since the maximum value of l for a given n is (n−1), the minimum n for l = 4 to exist is n = 5.

Question 2.25

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

Solution: For a 3d orbital: n = 3, l = 2, and m_l can be any one of −2, −1, 0, +1, +2.

Question 2.26

Question: 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: (i) Since the atom is neutral, protons = electrons = 29 (Z = 29, element is copper, Cu).
(ii) Electronic configuration: $1s^22s^22p^53s^23p^53d^{10}4s^1$ (exceptional stable configuration due to a fully filled 3d subshell).

Question 2.27

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

Solution: H_2 has 1 + 1 = 2 electrons. H_2^+ (one electron removed) has 1 electron. O_2 has 8 + 8 = 16 electrons, so O_2^+ has 15 electrons.

Question 2.28

Question: (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, 2. l = 0 ⇒ m_l = 0; l = 1 ⇒ m_l = −1, 0, +1; l = 2 ⇒ m_l = −2, −1, 0, +1, +2.
(ii) For 3d: l = 2, m_l = −2, −1, 0, +1, +2 (five orbitals).
(iii) 1p is not possible (n = 1 allows only l = 0). 2s is possible (n = 2, l = 0). 2p is possible (n = 2, l = 1). 3f is not possible (n = 3 allows only l = 0, 1, 2, not l = 3). So only 2s and 2p are possible.

Question 2.29

Question: 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 → 1s. (b) n = 3, l = 1 → 3p. (c) n = 4, l = 2 → 4d. (d) n = 4, l = 3 → 4f.

Question 2.30

Question: 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

Solution: (a) Not possible – n cannot be zero (n = 1, 2, 3…).
(b) Possible – represents a valid 1s electron with spin down.
(c) Not possible – for n = 1, l can only be 0 (max l = n−1); l = 1 is invalid.
(d) Possible – represents a valid 2p electron (l = 1 allows m_l = 0).
(e) Not possible – for n = 3, maximum l = 2; l = 3 is invalid.
(f) Possible – represents a valid 3p electron.

Question 2.31

Question: How many electrons in an atom may have the following quantum numbers?
(a) n = 4, m_s = −½ (b) n = 3, l = 0

Solution: (a) For n = 4, total electrons = $2n^2$ = 32. Exactly half of these have m_s = −½, so 16 electrons.
(b) n = 3, l = 0 is the 3s subshell, which is a single orbital holding a maximum of 2 electrons.

Question 2.32

Question: 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: Bohr's quantization condition: mvr = nh/2π, so mv = nh/2πr.

De Broglie's relation: $\lambda = h/mv$, so $mv = h/\lambda$.
Equating the two expressions for mv: $h/\lambda = nh/2\pi r \Rightarrow 2\pi r = n\lambda$.
This shows the circumference ($2\pi r$) of the Bohr orbit equals an integral multiple (n) of the de Broglie wavelength.

Question 2.33

Question: 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), wavenumber = $RZ^2(1/4 - 1/16) = 1.097 \times 10^7 \times 4 \times 0.1875 = 8.228 \times 10^6\ m^{-1}$.
For hydrogen (Z = 1): $R(1/n_1^2 - 1/n_2^2) = 8.228 \times 10^6 \Rightarrow 1/n_1^2 - 1/n_2^2 = 0.75$.
Trying $n_1 = 1$, $n_2 = 2$: 1 – 0.25 = 0.75 ✓.
So the n = 2 → n = 1 transition in the hydrogen spectrum has the same wavelength.

Question 2.34

Question: Calculate the energy required for the process $He^+(g) \rightarrow He^{2+}(g) + e^-$. The ionization energy for the H atom in the ground state is 2.18×10^{-18} J $atom^{-1}$.

Solution: This is the ionization energy of the hydrogen-like He^+ ion (Z = 2): $E = 2.18 \times 10^{-18} \times Z^2$ J/atom = $2.18 \times 10^{-18} \times 4 = 8.72 \times 10^{-18}$ J/atom.
Per mole: $8.72 \times 10^{-18} \times 6.022 \times 10^{23} = 5.25 \times 10^6$ J/mol = 5250 kJ/mol.

Question 2.35

Question: 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: Diameter = 1.5×10^{-10} m; scale length = 0.2 m.
Number of atoms = $0.2/(1.5 \times 10^{-10}) = 1.333 \times 10^9$ atoms.

Question 2.36

Question: 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: Diameter = length/number = $0.024\ m/(2 \times 10^8) = 1.2 \times 10^{-10}$ m.
Radius = diameter/2 = 6×10^{-11} m = 60 pm.

Question 2.37

Question: 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: Diameter = 2.6×10^{-10} m = 260 pm.
(a) Radius = 130 pm.
(b) Number of atoms = $0.016\ m/(2.6 \times 10^{-10}\ m) = 6.154 \times 10^7$ atoms.

Question 2.38

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

Solution: Number of electrons = $(2.5 \times 10^{-16})/(1.6 \times 10^{-19}) = 1562.5 \approx 1563$ electrons.

Question 2.39

Question: 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.

Solution: Number of electrons = $(1.282 \times 10^{-18})/(1.602 \times 10^{-19}) = 8$ electrons (exactly).

Question 2.40

Question: 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?

Solution: Heavy atoms like gold have a large, dense, highly positively charged nucleus, which causes significant deflection of α-particles, including occasional large-angle (back-)scattering – this is what revealed the existence of a small dense nucleus. If light atoms like aluminium were used instead, their nucleus has much smaller mass and charge, so the α-particles would undergo far less deflection, and very few (if any) large-angle scattering events would be observed, making the existence of a dense nucleus much harder to detect.

Question 2.41

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

Solution: The atomic number of an element is always fixed (Br always has Z = 35), so it is implied by the element's symbol itself and can be dropped – hence writing just the mass number with the symbol (^{79}Br) is acceptable and unambiguous. However, the mass number varies between isotopes of the same element and must always be specified to identify a particular isotope, so it cannot be dropped. Writing "35Br" is unacceptable because it wrongly implies a mass number of 35, which is incorrect and misleading since bromine's fixed atomic number is 35, not its mass number.

Question 2.42

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

Solution: Let protons = p; neutrons = 1.317p. $p + 1.317p = 81 \Rightarrow 2.317p = 81 \Rightarrow p \approx 35$.
Z = 35 is bromine (Br); neutrons = 81 – 35 = 46.
Symbol: $^{81}{}_{35}Br$.

Question 2.43

Question: 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: Let protons = p, electrons = p + 1 (extra electron from –1 charge). Neutrons = 1.111(p+1).

p + 1.111(p+1) = 37 ⇒ 2.111p + 1.111 = 37 ⇒ p ≈ 17 (chlorine, Cl).
Neutrons = 37 − 17 = 20; electrons = 18.
Symbol: $^{37}_{17}Cl^{-}$.

Question 2.44

Question: 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: Let protons = p, electrons = p − 3 (lost 3 electrons for +3 charge). Neutrons = 1.304(p−3).
p + 1.304(p−3) = 56 ⇒ 2.304p − 3.912 = 56 ⇒ p ≈ 26 (iron, Fe).
Neutrons = 56 − 26 = 30; electrons = 23.
Symbol: $^{56}_{26}Fe^{3+}$.

Question 2.45

Question: 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 increases as wavelength decreases across the electromagnetic spectrum: radio waves Increasing order: (c) FM radio

Question 2.46

Question: 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: Energy of one photon = $hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(337.1 \times 10^{-9}) = 5.899 \times 10^{-19}$ J.
Total energy = $5.6 \times 10^{24} \times 5.899 \times 10^{-19} = 3.3 \times 10^{6}$ J.
Taking this as the energy emitted per second, Power = 3.3×10^{6} W (≈ 3.3×10^{3} kW).

Question 2.47

Question: 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.

Solution: $\lambda = 616 \times 10^{-9}$ m.
(a) $v = c/\lambda = (3 \times 10^{8})/(616 \times 10^{-9}) = 4.87 \times 10^{14}$ Hz.
(b) Distance = $c \times t = 3 \times 10^{8} \times 30 = 9 \times 10^{9}$ m.
(c) $E = hv = 6.626 \times 10^{-34} \times 4.87 \times 10^{14} = 3.226 \times 10^{-19}$ J.
(d) Number of quanta = $2/(3.226 \times 10^{-19}) = 6.20 \times 10^{18}$.

Question 2.48

Question: 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: Energy of one photon = $hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^{8})/(600 \times 10^{-9}) = 3.313 \times 10^{-19}$ J.
Number of photons = $(3.15 \times 10^{-18})/(3.313 \times 10^{-19}) \approx 9.51$, i.e. approximately 9–10 photons.

Question 2.49

Question: 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: Rate of photon emission (taken as the effective frequency) = (number of photons)/(duration) = $(2.5 \times 10^{15})/(2 \times 10^{-9}) = 1.25 \times 10^{24}\ s^{-1}$.
Energy of source = h × (this frequency) = $6.626 \times 10^{-34} \times 1.25 \times 10^{24} = 8.28 \times 10^{-10}$ J.

Question 2.50

Question: 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: $v_1 = c/\lambda_1 = (3 \times 10^8)/(589 \times 10^{-9}) = 5.093 \times 10^{14}$ Hz.
$v_2 = c/\lambda_2 = (3 \times 10^8)/(589.6 \times 10^{-9}) = 5.088 \times 10^{14}$ Hz.
$\Delta E = h(v_1 - v_2) = 6.626 \times 10^{-34} \times 5.1 \times 10^{11} = 3.38 \times 10^{-22}$ J.

Question 2.51

Question: 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: $W_0 = 1.9 \times 1.602 \times 10^{-19} = 3.044 \times 10^{-19}$ J.
(a) $\lambda_0 = hc/W_0 = (6.626 \times 10^{-34} \times 3 \times 10^8)/(3.044 \times 10^{-19}) = 6.53 \times 10^{-7}$ m = 653 nm.
(b) $v_0 = W_0/h = (3.044 \times 10^{-19})/(6.626 \times 10^{-34}) = 4.594 \times 10^{14}$ Hz.
At λ = 500 nm: $E_{photon} = hc/\lambda = 3.976 \times 10^{-19}$ J.
$KE = E_{photon} - W_0 = 3.976 \times 10^{-19} - 3.044 \times 10^{-19} = 9.32 \times 10^{-20}$ J.
$v = \sqrt{(2KE/m)} = \sqrt{(2 \times 9.32 \times 10^{-20}/9.109 \times 10^{-31})} = 4.52 \times 10^5$ m/s.

Question 2.52

Question: 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 (m/s, $\times 10^5$): 2.55, 4.35, 5.35

Solution: Frequencies of incident light: v_1 = c/500nm = 6×10^{14} Hz; v_2 = c/450nm = 6.667×10^{14} Hz; v_3 = c/400nm = 7.5×10^{14} Hz.
Kinetic energies: $KE_1 = \frac{1}{2}mv_1^2 = \frac{1}{2}(9.109 \times 10^{-31})(2.55 \times 10^5)^2 = 2.96 \times 10^{-20}$ J; $KE_3 = \frac{1}{2}(9.109 \times 10^{-31})(5.35 \times 10^5)^2 = 1.30 \times 10^{-19}$ J.
Using Einstein's equation, $hv = W_0 + KE$, the slope of KE vs v gives h:
$h = (KE_3 - KE_1)/(v_3 - v_1) = (1.30 \times 10^{-19} - 2.96 \times 10^{-20})/(7.5 \times 10^{14} - 6 \times 10^{14}) \approx 6.7 \times 10^{-34}$ J s (close to the accepted value 6.626×10^{-34} J s).
$W_0 = hv_1 - KE_1 \approx 3.7 \times 10^{-19}$ J.
$\lambda_0 = hc/W_0 \approx 5.4 \times 10^{-7}$ m = 540 nm.

Question 2.53

Question: 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: KE = $eV_0 = 1.6 \times 10^{-19} \times 0.35 = 5.6 \times 10^{-20}$ J.
$E_{photon} = hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^8)/(256.7 \times 10^{-9}) = 7.746 \times 10^{-19}$ J.
$W_0 = E_{photon} - KE = 7.746 \times 10^{-19} - 5.6 \times 10^{-20} = 7.186 \times 10^{-19}$ J (≈ 4.49 eV).

Question 2.54

Question: 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: $E_{photon} = hc/\lambda = (6.626 \times 10^{-34} \times 3 \times 10^8)/(1.5 \times 10^{-10}) = 1.325 \times 10^{-15}$ J.
KE of ejected electron = $½mv^2 = ½(9.109 \times 10^{-31})(1.5 \times 10^7)^2 = 1.025 \times 10^{-16}$ J.
Binding energy = $E_{photon} - KE = 1.325 \times 10^{-15} - 1.025 \times 10^{-16} = 1.2225 \times 10^{-15}$ J.

Question 2.55

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

Solution: $v = c/\lambda = (3 \times 10^8)/(1285 \times 10^{-9}) = 2.335 \times 10^{14}$ Hz.
$2.335 \times 10^{14} = 3.29 \times 10^{15}[1/9 - 1/n^2] \Rightarrow 1/9 - 1/n^2 = 0.0710 \Rightarrow 1/n^2 = 0.0402 \Rightarrow n^2 \approx 25 \Rightarrow n = 5$.
This falls in the infrared region of the spectrum (Paschen series).

Question 2.56

Question: 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: Using $r_n = r_1n^2$ (r_1 = 52.9 pm): starting radius 1322.5 pm $\Rightarrow n^2 = 25 \Rightarrow n = 5$. Ending radius 211.6 pm $\Rightarrow n^2 = 4 \Rightarrow n = 2$.
Transition n = 5 → n = 2 belongs to the Balmer series (visible region).
Wavenumber = $R(1/4 - 1/25) = 1.097 \times 10^7 \times 0.21 = 2.304 \times 10^6$ m^{-1}.
$\lambda = 1/(2.304 \times 10^6) = 4.34 \times 10^{-7}$ m = 434.1 nm.

Question 2.57

Question: 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 1.6×10^6 m s^{-1}, calculate de Broglie wavelength associated with this electron.

Solution: $\lambda = h/mv = (6.626 \times 10^{-34})/(9.109 \times 10^{-31} \times 1.6 \times 10^6) = 4.549 \times 10^{-10}$ m = 454.9 pm.

Question 2.58

Question: 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: $v = h/(m\lambda) = (6.626 \times 10^{-34})/(1.675 \times 10^{-27} \times 8 \times 10^{-10}) = 494.5$ m/s.

Question 2.59

Question: 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: $\lambda = h/mv = (6.626 \times 10^{-34})/(9.109 \times 10^{-31} \times 2.19 \times 10^{6}) = 3.32 \times 10^{-10}$ m = 332 pm.

Question 2.60

Question: 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: $\lambda = h/mv = (6.626 \times 10^{-34})/(0.1 \times 4.37 \times 10^{5}) = 1.516 \times 10^{-38}$ m – an immeasurably small wavelength, which illustrates why macroscopic objects do not show observable wave behaviour.

Question 2.61

Question: If the position of the electron is measured within an accuracy of ± 0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is $h/4\pi_m \times 0.05$ nm, is there any problem in defining this value?

Solution: $\Delta x = 2 \times 10^{-12}$ m.
$\Delta p \geq h/(4\pi\Delta x) = (6.626 \times 10^{-34})/(4 \times 3.1416 \times 2 \times 10^{-12}) = 2.637 \times 10^{-23}$ kg m/s.
Given momentum $p = h/(4\pi \times 0.05 \times 10^{-9}) = 1.05 \times 10^{-24}$ kg m/s.
Since Δp (2.637×10^{-23}) is greater than p itself (1.05×10^{-24}), the uncertainty exceeds the value of the momentum, so yes – there is a real problem in meaningfully defining this momentum value at this level of positional accuracy, consistent with Heisenberg's uncertainty principle.

Question 2.62

Question: 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. $n = 4, l = 2, m_l = -2, m_s = -½$
2. $n = 3, l = 2, m_l = 1, m_s = +½$
3. $n = 4, l = 1, m_l = 0, m_s = +½$
4. $n = 3, l = 2, m_l = -2, m_s = -½$
5. $n = 3, l = 1, m_l = -1, m_s = +½$
6. $n = 4, l = 1, m_l = 0, m_s = +½$

Solution: Using the (n+l) rule (lower n+l = lower energy; for equal n+l, lower n is lower energy):
Electron 1 (4d): n+l = 6. Electron 2 (3d): n+l = 5. Electron 3 (4p): n+l = 5. Electron 4 (3d): n+l = 5. Electron 5 (3p): n+l = 4. Electron 6 (4p): n+l = 5.
Order of increasing energy: 5 < 2 = 4 < 3 = 6 < 1.
Electrons 2 and 4 (both 3d) have the same energy; electrons 3 and 6 (both 4p) have the same energy.

Question 2.63

Question: 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?

Solution: Electrons in orbitals farther from the nucleus (higher n) are more shielded by inner electrons and experience a lower effective nuclear charge. Since 4p is the outermost of the three (highest n, most shielded), electrons in the 4p orbital experience the lowest effective nuclear charge.

Question 2.64

Question: 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 (lower n, closer to nucleus, less shielded) experiences the larger effective nuclear charge.
(ii) 4d (lower l, more penetrating for the same n) experiences the larger effective nuclear charge than 4f.
(iii) 3p (lower l, more penetrating for the same n) experiences the larger effective nuclear charge than 3d.

Question 2.65

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

Solution: Si (Z = 14) has one more proton than Al (Z = 13), with similar shielding for the 3p electrons in both. Therefore, the unpaired 3p electron in Si experiences a greater effective nuclear charge than the unpaired 3p electron in Al.

Question 2.66

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

Solution: (a) P (Z=15): $[Ne]3s^2 3p^3 \rightarrow$ 3 unpaired electrons.
(b) Si (Z=14): $[Ne]3s^2 3p^2 \rightarrow$ 2 unpaired electrons.
(c) Cr (Z=24): $[Ar]3d^5 4s^1$ (exceptional configuration) $\rightarrow$ 6 unpaired electrons (5 in 3d + 1 in 4s).
(d) Fe (Z=26): $[Ar]3d^6 4s^2 \rightarrow$ 4 unpaired electrons (by Hund's rule in the 3d subshell).
(e) Kr (Z=36): $[Ar]3d^{10} 4s^2 4p^6 \rightarrow$ 0 unpaired electrons (all subshells fully filled).

Question 2.67

Question: (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, l can be 0, 1, 2, 3, giving 4 subshells: 4s, 4p, 4d, 4f.
(b) Total electrons for n = 4 = $2n^2$ = 32. Exactly half of these have $m_s = -1/2$, so 16 electrons.

Notes and Extra Questions

This chapter traces the historical development of atomic models — Thomson, Rutherford, and Bohr — before moving to the quantum mechanical model built on de Broglie's wave–particle duality and Heisenberg's uncertainty principle. Key numerical tools used throughout the exercises include the photoelectric effect equation ($E = h\nu = W_0 + KE$), the Rydberg/Bohr energy formula for hydrogen-like species ($E_n = -2.18 \times 10^{-18} Z^2/n^2$ J), the de Broglie wavelength relation ($\lambda = h/mv$), and the four quantum numbers (n, l, m_l, m_s) that together define an electron's state and govern the Aufbau principle, Pauli exclusion principle, and Hund's rule used to write electronic configurations. Students should be comfortable converting between wavelength, frequency, wavenumber, and energy, and applying unit conversions (nm, pm, Å, eV, erg) carefully, since most numerical errors in this chapter arise from unit mismatches rather than conceptual mistakes.