CLASS 11 PHYSICS NCERT SOLUTIONS

NCERT Solutions for Class 11 Physics Chapter 12: Kinetic Theory – Free PDF Download

Question: Estimate the fraction of the molecular volume to the actual volume occupied by oxygen gas at STP. Take the diameter of an oxygen molecule to be 3 Å.

Contents

NCERT Class 11 Physics Chapter 12, Kinetic Theory, explains how the macroscopic behaviour of gases — pressure, temperature, and the ideal gas equation — emerges from the random motion of a very large number of molecules. Below are complete, independently worked solutions to all 10 exercise questions from the current (2023 rationalised) NCERT textbook, following the exact answers a CBSE examiner would expect, along with concise revision notes and a short FAQ on the chapter.

NCERT Exercise Solutions – Kinetic Theory (Chapter 12)

12.1 Fraction of molecular volume to actual volume of oxygen at STP

Question: Estimate the fraction of the molecular volume to the actual volume occupied by oxygen gas at STP. Take the diameter of an oxygen molecule to be 3 Å.

Solution: Radius of one O_2 molecule, r = 1.5 Å = 1.5×10^{-10} m.
Volume of one molecule, v = $(4/3)\pi r^3 = (4/3) \times 3.1416 \times (1.5 \times 10^{-10})^3 = 1.414 \times 10^{-29}$ m^3.
At STP, 1 mole (6.022×10^{23} molecules) occupies 22.4 L = 22.4×10^{-3} m^3, so the actual volume available per molecule is
$V_{actual} = (22.4 \times 10^{-3})/(6.022 \times 10^{23}) = 3.72 \times 10^{-26}$ m^3.
Fraction = $v/V_{actual} = (1.414 \times 10^{-29})/(3.72 \times 10^{-26}) \approx$ **3.8×10^{-4}**. This very small ratio is why real gases at ordinary pressures behave close to ideal gases — molecules occupy a negligible fraction of the container's volume.

12.2 Molar volume of an ideal gas at STP

Question: Molar volume is the volume occupied by 1 mol of any (ideal) gas at standard temperature and pressure (STP: 1 atmospheric pressure, 0 °C). Show that it is 22.4 litres.

Solution: From the ideal gas equation, PV = nRT, so for n = 1 mol at STP (P = 1.013 × 10&sup5; Pa, T = 273 K):
V = nRT/P = (1 × 8.314 × 273)/(1.013 × 10&sup5;) = 2269.7/101300 = 0.0224 m^3 = **22.4 litres**, as required.

12.3 PV/T versus P plot for oxygen

Question: Figure 12.8 (Fig. 13.8 in earlier editions) shows a plot of PV/T versus P for 1.00×10^{-3} kg of oxygen gas at two different temperatures.
(a) What does the dotted line signify?
(b) Which is true: $T_1 > T_2$ or $T_1 < T_2$?
(c) What is the value of PV/T where the curves meet on the y-axis?
(d) If we obtained similar plots for 1.00×10^{-3} kg of hydrogen, would we get the same value of PV/T at the point where the curves meet on the y-axis? If not, what mass of hydrogen yields the same value?

Solution:
(a) The dotted line represents the ideal gas behaviour, i.e., PV/T = μR = constant, independent of pressure. Real gases approach this line as P → 0.
(b) At any given pressure, the curve closer to the dotted (ideal) line deviates least from ideal behaviour. Since a gas behaves more ideally at higher temperature (molecules are farther apart on average relative to their size and interactions), the curve for T_1 lies closer to the dotted line, so **$T_1 > T_2$**.
(c) At the y-axis (P → 0), PV/T = μR, independent of temperature. Number of moles μ = mass/molar mass = $(1.00 \times 10^{-3}$ kg$)/(32 \times 10^{-3}$ kg/mol) = 0.03125 mol.
PV/T = μR = 0.03125 × 8.314 = **0.26 J K^{-1}** (approximately).

(d) No. Since PV/T = μR depends on the number of moles μ, not on the mass alone, using the same mass (1.00×10^{-3} kg) of hydrogen (M = 2 g/mol) would give a different, larger μ and hence a different intercept. To get the *same* value of PV/T, hydrogen must have the same number of moles as the oxygen sample: mass of $H_2 = \mu \times M_{H2} = 0.03125 \times 2 \times 10^{-3}$ kg = **6.25×10^{-5} kg (0.0625 g)**.

12.4 Mass of oxygen withdrawn from a cylinder

Question: An oxygen cylinder of volume 30 litres has an initial gauge pressure of 15 atm and a temperature of 27°C. After some oxygen is withdrawn, the gauge pressure drops to 11 atm and the temperature drops to 17°C. Estimate the mass of oxygen taken out of the cylinder. (R = 8.31 J mol^{-1} K^{-1}, molecular mass of O_2 = 32 u.)

Solution: Using PV = nRT with R = 0.0821 L atm mol^{-1} K^{-1}:
Initial moles: $n_1 = P_1V/RT_1 = (15 \times 30)/(0.0821 \times 300) = 450/24.63 = 18.27$ mol.
Final moles: $n_2 = P_2V/RT_2 = (11 \times 30)/(0.0821 \times 290) = 330/23.81 = 13.86$ mol.
Moles withdrawn, $\Delta n = n_1 - n_2 = 18.27 - 13.86 = 4.41$ mol.
Mass withdrawn = Δn × 32 g/mol = 4.41 × 32 ≈ **141 g**.

12.5 Volume of a rising air bubble

Question: An air bubble of volume 1.0 cm^3 rises from the bottom of a lake 40 m deep at a temperature of 12°C. To what volume does it grow when it reaches the surface, which is at a temperature of 35°C? (1 atm = 1.01 × 10&sup5; Pa, and density of water = 1000 kg m^{-3}; g = 9.8 m s^{-2}.)

Solution: Pressure at the bottom, $P_1 = P_{atm} + \rho gh$ = 1.01 × 10&sup5; + (1000 × 9.8 × 40) = 1.01 × 10&sup5; + 3.92 × 10&sup5; = 4.93 × 10&sup5; Pa.
Pressure at the surface, P_2 = 1.01 × 10&sup5; Pa.
T_1 = 285 K, T_2 = 308 K.
Using $P_1V_1/T_1 = P_2V_2/T_2$:
$V_2 = V_1 \times (P_1/P_2) \times (T_2/T_1) = 1.0 \times (4.93/1.01) \times (308/285) = 1.0 \times 4.881 \times 1.081 \approx$ **5.3 cm^3**.

12.6 Number of air molecules in a room

Question: Estimate the total number of air molecules (inclusive of oxygen, nitrogen, water vapour, and other constituents) in a room of capacity 25.0 m^3 at a temperature of 27°C and 1 atm pressure.

Solution: n = PV/RT = (1.013 × 10&sup5; × 25.0)/(8.314 × 300) = $(2.5325 \times 10^6)/2494.2$ = 1015.4 mol.
Number of molecules $N = n \times N_A = 1015.4 \times 6.022 \times 10^{23} \approx$ **6.11×10^{26} molecules**.

12.7 Average thermal energy of a helium atom

Question: Estimate the average thermal energy of a helium atom at (i) room temperature (27°C), (ii) the temperature on the surface of the Sun (6000 K), and (iii) the temperature of 10 million kelvin (the core of a star).

Solution: Average thermal (translational kinetic) energy per molecule = $(3/2)k_BT$, where $k_B = 1.38 \times 10^{-23}$ J/K.
(i) T = 300 K: $E = 1.5 \times 1.38 \times 10^{-23} \times 300 =$ **6.21×10^{-21} J**.
(ii) T = 6000 K: $E = 1.5 \times 1.38 \times 10^{-23} \times 6000 =$ **1.24×10^{-19} J**.
(iii) $T = 1.0 \times 10^6$ K: $E = 1.5 \times 1.38 \times 10^{-23} \times 1.0 \times 10^6 =$ **2.07×10^{-16} J**.

12.8 Neon, chlorine and uranium hexafluoride in equal vessels

Question: Three vessels of equal capacity have gases at the same temperature and pressure. The first vessel contains neon (monatomic), the second contains chlorine (diatomic), and the third contains uranium hexafluoride (polyatomic). Do the vessels contain equal number of respective molecules? Is the root mean square speed of molecules the same in the three cases? If not, in which case is v_{rms} the largest?

Solution: By Avogadro's law, equal volumes of any gas at the same temperature and pressure contain an **equal number of molecules**, regardless of the gas's atomicity or molar mass. So yes, all three vessels contain the same number of molecules.
However, $v_{rms} = \sqrt{(3RT/M)}$ depends on the molar mass M. Since T is the same for all three, v_{rms} is **not** the same: it is inversely proportional to $\sqrt{M}$. Neon has the smallest molar mass ($M \approx 20.2$ g/mol) compared to chlorine ($M \approx 71$ g/mol) and uranium hexafluoride ($M \approx 352$ g/mol), so **neon has the largest v_{rms}**.

12.9 Temperature for equal rms speeds of argon and helium

Question: At what temperature is the root mean square speed of an atom in an argon gas cylinder equal to the rms speed of a helium gas atom at –20°C? (Atomic mass of Ar = 39.9 u, of He = 4.0 u.)

Solution: $v_{rms} = \sqrt{(3RT/M)}$. Setting $v_{rms,Ar}(T) = v_{rms,He}(253\ K)$:
$3RT/M_{Ar} = 3R(253)/M_{He}$
$T = 253 \times (M_{Ar}/M_{He}) = 253 \times (39.9/4.0) = 253 \times 9.975 \approx$ **2523.7 K**.

12.10 Mean free path and collision frequency of nitrogen

Question: Estimate the mean free path and collision frequency of a nitrogen molecule in a cylinder containing nitrogen at 2.0 atm and temperature 17°C. Take the radius of a nitrogen molecule to be roughly 1.0 Å. Compare the collision time with the time the molecule moves freely between two successive collisions (molecular mass of N_2 = 28.0 u).

Solution: Diameter d = 2.0 Å = 2.0×10^{-10} m; T = 290 K; P = 2.0 × 1.013 × 10&sup5; Pa = 2.026 × 10&sup5; Pa.
Number density, $n = P/(k_BT) = (2.026 \times 10\&sup5;)/(1.38 \times 10^{-23} \times 290) = 5.06 \times 10^{25}$ molecules/m^3.
Mean free path, $\lambda = 1/(\sqrt{2} \cdot \pi d^2 \cdot n) = 1/(1.414 \times 3.1416 \times (2.0 \times 10^{-10})^2 \times 5.06 \times 10^{25})$
$= 1/(1.414 \times 3.1416 \times 4.0 \times 10^{-20} \times 5.06 \times 10^{25}) \approx$ **1.11×10^{-7} m**.
$v_{rms} = \sqrt{(3RT/M)} = \sqrt{(3 \times 8.314 \times 290/0.028)} = \sqrt{(258{,}330)} \approx 508$ m/s.
Collision frequency $= v_{rms}/\lambda = 508/(1.11 \times 10^{-7}) \approx$ **4.58×10^9 collisions per second**.
Collision time $\tau = d/v_{rms} = (2.0 \times 10^{-10})/508 \approx 3.9 \times 10^{-13}$ s. Free time between collisions $T_2 = \lambda/v_{rms} = (1.11 \times 10^{-7})/508 \approx 2.2 \times 10^{-10}$ s. The ratio $T_2/\tau \approx 560$, i.e., a molecule spends far more time moving freely than colliding — this justifies treating molecular collisions as instantaneous point events in kinetic theory.

Notes and Extra Questions

Key formulas from the chapter:

Ideal gas equation: PV = nRT (or PV = μRT, where μ is the number of moles).
Pressure from kinetic theory: $P = (1/3)\rho v^2_{rms} = (1/3)(nm)v^2_{rms}$, where n is number density and m is the mass of a molecule.
Average kinetic energy per molecule: $(1/2)mv^2_{rms} = (3/2)k_BT$ (this is independent of the nature of the gas — it depends only on temperature).
Root mean square speed: $v_{rms} = \sqrt{(3RT/M)} = \sqrt{(3k_BT/m)}$.
Law of equipartition of energy: each degree of freedom (translational, rotational, or vibrational) contributes

$(1/2)k_BT$ of energy per molecule on average. A vibrational mode contributes k_BT (kinetic + potential) if fully active.
Degrees of freedom: monatomic gas — 3 (translational only); diatomic gas — 5 at moderate temperatures (3 translational + 2 rotational), rising to 7 at high temperatures if vibration is active; polyatomic (non-linear) gas — 6 (3 translational + 3 rotational), plus vibrational modes at higher temperatures.
Specific heats from degrees of freedom: $C_v = (f/2)R$ and $C_p = C_v + R$, where f is the number of degrees of freedom; so $\gamma = C_p/C_v = 1 + 2/f$.
Mean free path: $\lambda = 1/(\sqrt{2}\,\pi d^2 n)$, the average distance a molecule travels between successive collisions, where d is the molecular diameter and n is the number density.

Concepts worth revising carefully: the distinction between v_{rms}, average speed, and most probable speed (they are numerically different even though all three describe molecular speed distribution, per the Maxwell speed distribution introduced conceptually in this chapter); why gas pressure is independent of the nature of the gas at a given number density and temperature; why real gases deviate from the ideal gas law at high pressure and low temperature (finite molecular size and intermolecular attraction, addressed by the Van der Waals equation, mentioned qualitatively in the chapter); and the assumptions of the kinetic theory of gases (point-sized molecules in random motion, elastic collisions, no intermolecular forces except during collision, and Newtonian mechanics applying to each molecule).

Note on the current syllabus: as part of the CBSE/NCERT 2023 content rationalisation, the sub-topic “Specific Heat Capacity of Water” and the exercise questions that accompanied it (which dealt with the anomalous behaviour of water’s specific heat using the vibrational degrees of freedom of the H_2O molecule) have been removed from the current textbook. The exercise numbering above (12.1–12.10) reflects only the questions that remain in the current, examinable 2026-27 edition.