03 The Gaseous State

Set A **Answer key**

A From the lesson

13 MARKS

1 (a)

25 + 273 = **298 K** [1]

(b)

$250 \times 10^{-6}\ m^3$ = **$2.50 \times 10^{-4}\ m^3$** [1]

(c)

2.00 × 101 325 Pa = **2.03×10^5 Pa** [1]

2 (a)

$T_1 = 27 + 273 = 300$ K, $T_2 = 127 + 273 = 400$ K [1]

Apply $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$V_2 = 2.00 \times \frac{400}{300}$

= **$2.67\ dm^3$** [1]

(b)

V is proportional to T only when T is in kelvin. 127 °C is not 127 ÷ 27 times as hot as 27 °C. [1]

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$T = 23 + 273 = 296$ K, $V = 246 \times 10^{-6}\ m^3$

Using $pV = nRT = \frac{m}{M_r}RT$,

$M_r = \frac{0.320 \times 8.31 \times 296}{1.00 \times 10^5 \times 246 \times 10^{-6}}$

= **32.0** [1]

The gas is **O_2**. [1]

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Apply $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$, with $T_1 = 300$ K and $T_2 = 400$ K [1]

$p_2 = 100 \times \frac{2.00}{0.500} \times \frac{400}{300}$

= **533 kPa** [1]

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p_{O_2} = mole fraction of O_2 × total pressure

= 0.21 × 101

= **21.2 kPa** [1]

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NH_3 deviates more.

1. NH_3 molecules form hydrogen bonds with each other, which are stronger than the instantaneous dipole–induced dipole interactions between CH_4 molecules. [1]
2. The intermolecular forces of attraction between NH_3 molecules are less negligible, so NH_3 deviates more from ideal behaviour. [1]

B Practice

27 MARKS

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B 87.5 kPa [1]

Using $p_1V_1 = p_2V_2$,

$50 \times 10 = p_{small}(10 + 30)$, so $p_{small} = \frac{500}{40} = 12.5$ kPa

$100 \times 30 = p_{large}(10 + 30)$, so $p_{large} = \frac{3000}{40} = 75$ kPa

$p_{total} = 12.5 + 75 = 87.5$ kPa

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C 1 g of H_2 and 2 g of D_2 in a total volume V [1]

Pressure exerted by 2 g of H_2, $p = \frac{(2 \div 2(1.0))RT}{V} = \frac{RT}{V}$

A Pressure exerted = $\frac{(2 \div 2(2.0))RT}{V} = \frac{RT}{2V}$

B Pressure exerted = $\frac{(4 \div 2(2.0))RT}{V/2} = \frac{2RT}{V}$

C Pressure exerted = $\frac{(1 \div 2(1.0))RT}{V} + \frac{(2 \div 2(2.0))RT}{V} = \frac{RT}{V}$

D Pressure exerted = $\frac{(2 \div 2(1.0))RT}{2V} + \frac{(1 \div 2(2.0))RT}{2V} = \frac{5RT}{8V}$

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(a)

Amount of oxygen = $\frac{120}{16.0 \times 2} = 3.75$ mol

Amount of hydrogen = $\frac{5.00}{1.0 \times 2} = 2.50$ mol

Mole fraction of oxygen, $x_{O_2} = \frac{3.75}{3.75 + 2.50} = 0.600$ [1]

$p_{O_2} = x_{O_2} \times p_T$

$3.00 = 0.600 \times p_T$

$p_T =$ **5.00 atm** [1]

(b)

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

H_2 is limiting and is used up.

Amount of oxygen remaining = $3.75 - \frac{2.50}{2} = 2.50$ mol [1]

The water condenses on cooling, so the final pressure is due to 2.50 mol of O_2 only.

Pressure = $\frac{2.50}{3.75} \times 3.00 =$ **2.00 atm** [1]

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Low pressure and high temperature [1]

At low pressure, the gas molecules are far apart. The volume of the gas molecules therefore becomes insignificant compared to the volume of the container *or* the intermolecular forces of attraction are negligible. [1]

At high temperature, the gas molecules move faster (higher kinetic energy) and the intermolecular forces of attraction become negligible. [1]

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(a)

Using the ideal gas equation,

$$p = \frac{nRT}{V} = \frac{1.00 \times 8.31 \times 298}{0.250 \times 10^{-3}}$$

$$= \mathbf{9.91 \times 10^6\ Pa}$$ [1]

(b)

The actual pressure is lower than the pressure calculated from the ideal gas equation because significant intermolecular forces of attraction (instantaneous dipole–induced dipole interactions) exist between CO_2 molecules. [1]

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	C	H
% by mass	85.7	14.3
A_r	12.0	1.0
no. of moles	7.142	14.3
simplest ratio	1	2

Empirical formula of A: $\mathbf{CH_2}$ [1]

$$C_xH_y + (x + \frac{y}{4})O_2 \rightarrow xCO_2 + \frac{y}{2}H_2O$$

Volume of CO_2 produced = 65.0 − 35.0 = 30.0 cm^3

Volume of O_2 reacted = 80.0 − 35.0 = 45.0 cm^3 [1]

$\frac{n_A}{n_{CO_2}} = \frac{10.0}{30.0} = \frac{1}{x}$, so x = 3

$\frac{n_A}{n_{O_2}} = \frac{10.0}{45.0} = \frac{1}{3 + y/4}$, so y = 6

Molecular formula of A: $\mathbf{C_3H_6}$ [1]

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Increasing deviation: $\mathbf{He < Ne < O_2 < C_2H_6 < NH_3 < H_2O}$ [1]

H_2O and NH_3 deviate more than the rest: they have stronger hydrogen bonding, compared with the weaker instantaneous dipole–induced dipole interactions in the others. H_2O has more extensive and stronger hydrogen bonding than NH_3. [1]

$He < Ne < O_2 < C_2H_6$: as the number of electrons increases, the electron cloud is larger and more polarisable, so the instantaneous dipole–induced dipole interactions are stronger. [1]

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(a)

At zero pressure, all three gases behave ideally, so their pV values meet at the same point (pV = nRT is the same constant). [1]

(b)

For an ideal gas, pV stays constant as p increases.

Ne: as pressure increases (volume decreases), the volume of the Ne atoms becomes significant compared with the total volume of the container. The pressure exerted is greater than for an ideal gas, so pV is greater than ideal. [1]

O_2 and NH_3: as pressure increases, the molecules are pushed closer together, so the intermolecular forces of attraction become significant. The molecules collide with the walls less forcefully, so the pressure is lower than ideal and pV decreases. [1]

NH_3 deviates more than O_2 because of the stronger hydrogen bonding between NH_3 molecules, compared with the weaker instantaneous dipole–induced dipole interactions between O_2 molecules. [1]

NOTE Ne has only very weak instantaneous dipole–induced dipole interactions (a very small electron cloud), so the volume of its atoms has the larger effect.

15 **(a)**

$pV = nRT$

$(6.00 \times 10^5)(670 \times 10^{-6}) = n(8.31)(273 + 20)$ [1]

$n =$ **0.165 mol** [1]

(b)

$$\frac{p_1}{T_1} = \frac{p_2}{T_2}$$

$$\frac{6.00}{20 + 273} = \frac{\text{p at 10 000 m}}{5 + 273}$$ [1]

p at 10 000 m = **5.69 bar** = 5.69×10^5 Pa [1]

(c)

It is not necessary to deflate the bicycle tyre.

The maximum difference between the internal and external pressure the tyre can endure without bursting = 8 – 1.01 = 6.99 bar [1]

At 10 000 m, the difference between the internal and external pressure of the tyre = 5.693 – 0.28 = 5.41 bar, which is within what the tyre can endure without bursting. [1]

EXAM TIP The tyre bursts because of the pressure difference, so compare differences, not the internal pressures alone.