03 The Gaseous State

Set B **Answer key**

A Multiple choice

3 MARKS

1 **D** 800 kPa [1]

At constant T, apply $p_1V_1 = p_2V_2$ to each gas.

$$p_{O_2} = \frac{200 \times 5.0}{2.5} = 400 \text{ kPa}$$

$$p_{N_2} = \frac{500 \times 2.0}{2.5} = 400 \text{ kPa}$$

Total pressure = 400 + 400 = 800 kPa

2 **C** NO_2 [1]

$pV = nRT$, where V, R and T are the same for all four gases, so $p \propto n$.

The greatest M_r gives the smallest number of moles for the same mass, and hence the lowest pressure.

gas	H_2	CH_4	NO_2	Ar
M_r or A_r	2.0	16.0	46.0	39.9
$n = 5.00 \div M_r$	2.50	0.313	0.109	0.125

3 **B** helium [1]

An ideal gas should have negligible intermolecular forces of attraction.

Ammonia has hydrogen bonding, which has the strongest intermolecular forces of attraction.

Between helium, krypton and nitrogen, helium has the smallest electron cloud, so the weakest instantaneous dipole–induced dipole interactions.

B Structured questions

33 MARKS

4 **(a)** High temperature and low pressure. [1]

(b)

$$V = \frac{(5.2)(8.31)(15 + 273)}{101\,300} = 0.123 \text{ m}^3$$

$$= \mathbf{123 \text{ dm}^3}$$ [1]

(c) Hydrogen bonding in ammonia causes the molecules to be closer to one another, so the gas occupies a volume smaller than expected. [1]

5 **(a)**

Molecules of an ideal gas have zero (negligible) volume. [1]

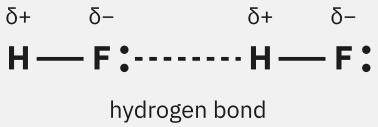
There are negligible (no) intermolecular forces of attraction between ideal gas molecules. [1]

Collisions between ideal gas molecules are perfectly elastic (no loss of energy during collision). [1]

(b)(i)

Hydrogen bond drawn from a lone pair on F of one molecule to the H of the other, and labelled.

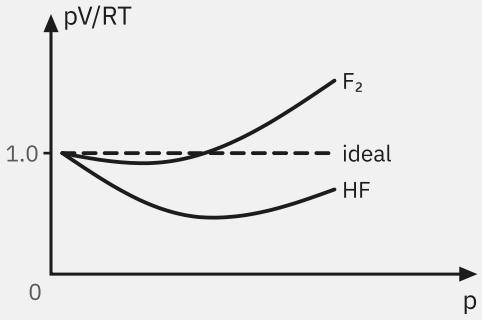
Lone pair and dipoles (δ+ on H, δ− on F) shown. [1]



(ii)

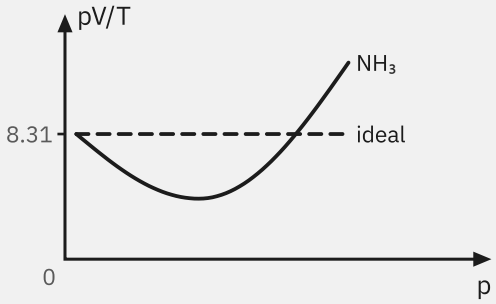
Correct sketch: HF dips further below 1.0 than fluorine. [1]

HF deviates more from ideal behaviour because it has stronger intermolecular hydrogen bonding between its molecules, compared to F_2, which has only instantaneous dipole–induced dipole interactions between its molecules. [1]



6 **(a)**

Horizontal line at $pV/T = nR = 8.31\ J\ K^{-1}$ for 1 mol. [1]



(b)

NH_3 graph below the ideal line at moderate pressure. [1]

At moderately high pressure, NH_3 molecules come closer together and the intermolecular attractive forces (hydrogen bonds) between NH_3 molecules become significant. This causes the gas to occupy a volume smaller than that of an ideal gas. [1]

(c)

27 °C = 300 K and 227 °C = 500 K

When compressed at constant T to half V, apply $p_1V_1 = p_2V_2$

$4.00 \times 5.00 = p_2 \times 2.50$

$p_2 = 8.00$ atm [1]

When heated from 300 K to 500 K at constant V, apply $\frac{p_2}{T_2} = \frac{p_3}{T_3}$

$$\frac{8.00}{300} = \frac{p_3}{500}$$

$p_3 =$ **13.3 atm** [1]

7 **(a)**

Volume of O_2 consumed per minute = $\frac{21 - 14}{100} \times 500 \times 12$

$= 420\ cm^3 = 4.20 \times 10^{-4}\ m^3$ [1]

$pV = nRT$

$(101\ 325)(4.20 \times 10^{-4}) = n(8.31)(37 + 273)$ [1]

$n =$ **0.0165 mol** $(1.652 \times 10^{-2}\ mol)$ [1]

(b)

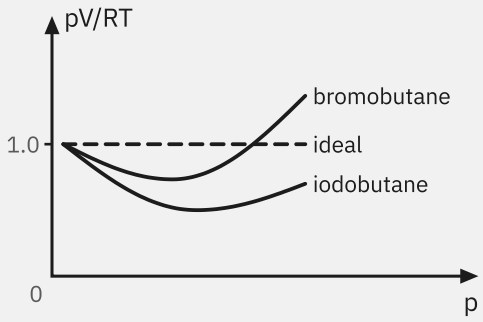
From reaction 1, 6 mol of O_2 releases 2818 kJ of energy.

Energy released per minute = $\frac{2818}{6} \times 0.01652$

= **7.76 kJ** [1]

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Correct sketch: bromobutane dips below 1.0, but less than iodobutane. [1]



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Any two of: [2]

There are significant intermolecular forces of attraction (hydrogen bonding) between molecules, which are not negligible.

or The particles of N_2H_4 take up space relative to the volume of the gas, which is not negligible.

or Collisions are inelastic.

10 **(a)(i)**

Down the group, the atomic size increases. Hence, the volume of the gas particles becomes less negligible (more significant) compared to the volume of the container. [1]

Also, down the group, the number of electrons in the atoms increases and the electron cloud becomes bigger and more polarisable, leading to stronger instantaneous dipole–induced dipole interactions. Hence, the intermolecular forces of attraction become less negligible (more significant). [1]

(ii)

$n = \frac{pV}{RT} = \frac{8.00 \times 10^4 \times 1.00 \times 10^{-3}}{8.31 \times 300} = 0.03209$ mol [1]

$M_r = \frac{2.79}{0.03209} =$ **86.9** [1]

(b)

$A_r = \frac{80(2.4) + 82(11.7) + 83(11.6) + 84(56.9) + 86(17.4)}{100}$

= **83.9** [1]

(c)

$$n = \frac{1.00 \times 10^5 \times 1.00 \times 10^{-3}}{8.31 \times 293} = 0.04107 \text{ mol}$$ [1]

$$A_r = \frac{3.40}{0.04107} = \mathbf{82.8}$$ [1]

The sample is not likely to have the same isotopic composition as that in (b), as its average A_r is different from the average A_r calculated in (b). [1]

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Since Xe and F_2 are mixed in equal volumes, the partial pressure of both gases is the same. Let it be p.

Initial total pressure = p + p = 2p

Final total pressure = 0.70 × 2p = 1.4p

Final partial pressure of Xe = $\frac{4}{7}(1.4p) = 0.8p$, final partial pressure of F_2 = 1.4p − 0.8p = 0.6p [1]

V and T are constant, so pV = nRT gives $p \propto n$.

Amount of Xe reacted $\propto$ p − 0.8p = 0.2p

Amount of F_2 reacted $\propto$ p − 0.6p = 0.4p, so 1 mol of Xe reacts with 2 mol of F_2. [1]

$$Xe(g) + 2F_2(g) \rightarrow XeF_4(s)$$ [1]

EXAM TIP At constant V and T, pressure is proportional to amount, so partial pressures can be used in place of moles.