Raffles Institution
Year 5 H2 Chemistry 2025

Tutorial 4 – The Gaseous State
(Suggested Answers for Discussion Questions)

Discussion Questions

11 Mass of CO_2 in 500 cm^3 of saturated solution = 0.500 × 1.50 = 0.750 g
Mass of CO_2 released = 2.00 – 0.750 = 1.25 g
Amount of CO_2 released = 1.25/molar mass = 1.25/44.0 = 0.0284 mol
Volume of CO_2 released = 0.0284 × 24
= 0.682 dm^3
= **<u>682 cm^3</u>**

12 **(a)** For a fixed amount of ideal gas at constant T, pV remains constant as p increases.

(b) Using the ideal gas equation, $pV = nRT$

$$V = \frac{(0.40)(8.31)(300)}{12.0 \times 10^5}$$
$= 8.31 \times 10^{-4}\ m^3$
$= 0.831\ dm^3$

(c)

pressure, p / Pa	volume, V / dm^3	pressure x volume, pV / Pa dm^3
5.0×10^5	1.924	9.62×10^5
10.0×10^5	0.926	9.26×10^5
15.0×10^5	0.592	8.88×10^5

When p = 12.0×10^5 Pa,
estimated pV = $9.26 \times 10^5 - \frac{2}{5}(9.26 \times 10^5 - 8.88 \times 10^5)$
= 9.11×10^5 Pa dm^3

When estimated 9.11×10^5 Pa dm^3,
V = 9.11×10^5 Pa $dm^3 \div 12.0 \times 10^5$ Pa
= 0.759 dm^3

(d) Value of V in **(c)** is smaller than that in **(b)**. This implies that the gas occupies a volume smaller than what it would were it ideal. As the pressure exerted on the system is moderately high, <u>the polar HC*l* gas molecules come closer together and the intermolecular attractive forces (id-id and pd-pd) between the HC*l* gas molecules become significant</u>, causing the volume to be smaller than what it would were it ideal.

13 (a) 2 main assumptions are:

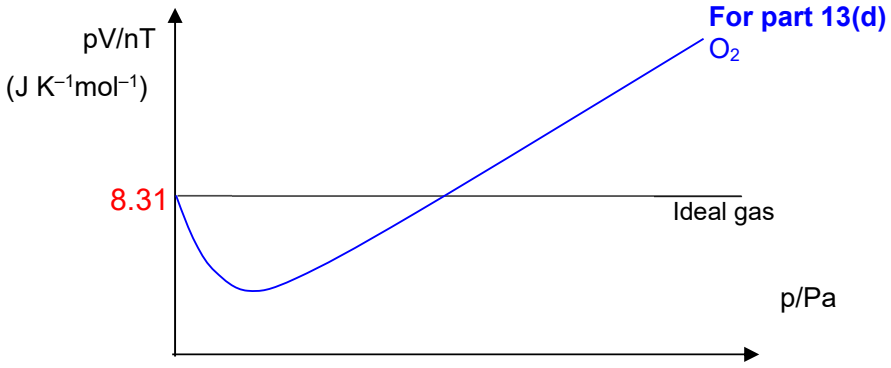
- The gas particles <u>exert negligible intermolecular attractive forces</u> on one another.
- The <u>volume of gas particles is negligible compared to the volume of container</u>.

Reject: volume of gas is negligible (missing 'particles')

Other assumptions:

- The <u>gas particles are in constant random motion</u>.
- <u>Collision between gas particles</u> are <u>perfectly elastic</u>.
- The <u>average kinetic energy of gas particles</u> is <u>proportional to absolute temperature</u>.

(b) $\frac{pV}{nT} = R = 8.31$ J K^{-1}mol^{-1}



(c) Under <u>low pressure</u> and <u>high temperature</u>

- At <u>low pressure</u>, the gas particles are very <u>far apart</u>. Hence the <u>volume of the gas particles</u> can be considered <u>negligible</u> compared to the volume of the container. In addition, the <u>intermolecular attractive forces</u> between widely spaced gas particles are also <u>negligible.</u>

- At <u>high temperature</u>, the gas particles possess <u>sufficiently high kinetic energy</u> to overcome the intermolecular attractive forces. Hence, the <u>intermolecular attractive forces</u> can be considered <u>negligible</u>.

(d) At moderately high pressure:

- O_2 molecules come closer together and the <u>intermolecular attractive forces between O_2 molecules become significant</u>. This causes the gas to occupy a volume smaller than that of an ideal gas. This results in a <u>smaller V</u> and pV/nT < 8.31 (negative deviation).

At very high pressure:

- The <u>molecular size</u> of O_2 gas particles is <u>significant</u>, resulting in the gas to occupy a <u>larger V</u> than expected and and pV/nT > 8.31 (positive deviation).
- The O_2 molecules are so close together that the repulsive forces between their electron clouds become significant.
- Thus, the <u>volume of the gas particles is **not negligible** as compared to the volume of the container.</u>

Notes for consideration
In explaining deviation of a particular real gas versus ideal gas, it is still useful to look back at the two assumptions of ideal gas behavior (i.e. negligible IMF and negligible particle size) in (a) and how these two assumptions are violated when we look at a real gas.

Here, O_2 gas deviates from ideal gas behavior because 1) it experiences significant intermolecular attractive forces (id-id) and 2) molecular size of O_2 are not negligible – both points violate the ideal gas assumptions. The second point on repulsive forces will help to bring out the trend in positive deviation.

14 (a) (i) At the intercept of the pV axis (low pressure), both gases approach ideal gas behaviour.
$\Rightarrow$ pV = nRT = (1)(8.31)(298) = **2.48 x 10^3 J (or Pa m^3)** (to 3 sf)

(ii) CO_2 shows greater negative deviation at moderately high pressure and greater positive deviation at very high pressure.

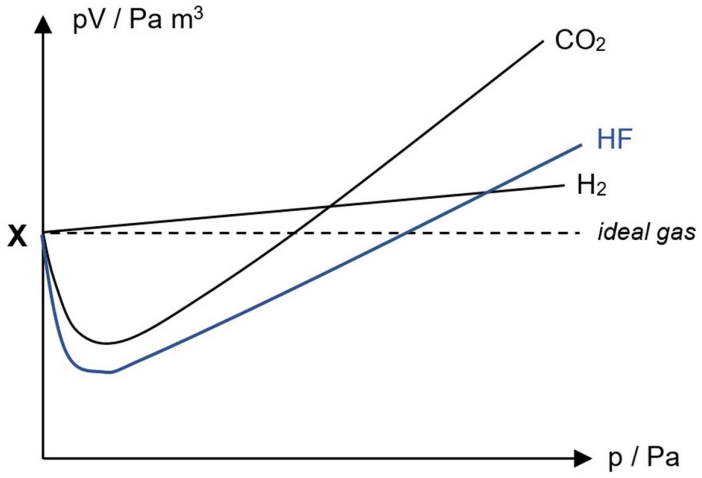
- For greater negative deviation:

CO_2 has a greater electron cloud/ number of electrons compared to H_2. Thus, CO_2 forms stronger instantaneous dipole – induced dipole interactions as its electron cloud is more polarisable.

- For greater positive deviation:

The molecular size of CO_2 molecules is significantly larger than that of H_2. The size of electron cloud for CO_2 is bigger than that of H_2. This results in stronger repulsive forces between CO_2 molecules compared to H_2 molecules.

(b)



At moderately high pressures, the pV of HF has a greater negative deviation from ideality than CO_2 due to stronger hydrogen bonds between HF molecules as compared to the weaker id-id interactions between CO_2 molecules.

Comments for tutors:
You may wish to discuss positive deviation of HF. Usually assessment questions will have clearer mark allocation and specify what needs to be discussed. In the case of HF, likely negative deviation is the focus due to strong hydrogen bonds.

At very high pressure, the pV of HF has a positive deviation which is lesser than CO_2 but greater than H_2. This is because HF has a smaller molecular size than CO_2 but larger molecular size than H_2.

15 **(a)** Amount of oxygen = $120/(16.0 \times 2) = 3.75$ mol
Amount of hydrogen = $5.00/(1.0 \times 2) = 2.50$ mol
Mole fraction of oxygen, x_{O2} = $3.75/(3.75 + 2.50) = 0.600$

Let the final total pressure be p_T.
$p_{O2} = x_{O2}\, p_T$, $3.00 = 0.600 \times p_T \Rightarrow p_T =$ **5.00 atm**

(b) On exploding the mixture: $2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$
2.50 3.75
H_2 is limiting and all H_2 will be used up.

Amount of oxygen that remains after the reaction = 3.75 - (2.50/2) = 2.50 mol

Assume that the pressure and volume of water is negligible (water formed condenses upon cooling) $\Rightarrow$ the final pressure is only due to 2.50 mol of O_2.

Under the same conditions, 3.75 mol of oxygen has a pressure of 3.00 atm
$\Rightarrow$ 2.50 mol of oxygen would have a pressure = 2.50/3.75 x 3.00 = **2.00 atm**

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

Initial total pressure = p + p = 2p (Dalton's Law of partial pressure)
Final total pressure = 0.70 x 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

Since V is constant (sealed vessel) and assuming constant T and ideal gas behaviour:
$pV = nRT \Rightarrow n \propto p$
Hence, amount of Xe reacted $\propto$ (p – 0.8p) = 0.2p
amount of F_2 reacted $\propto$ (p – 0.6p) = 0.4p
i.e. $\frac{n_{Xe}}{n_{F_2}} = \frac{p_{Xe}}{p_{F_2}} = \frac{0.2p}{0.4p} = \frac{1}{2}$ or 1 mol Xe reacts with 2 mol F_2.

$\Rightarrow$ **$Xe\,(g) + 2F_2\,(g) \rightarrow XeF_4\,(s)$**