The Gaseous State Discussion Questions

1 Which option has the largest volume at 25 °C and 101 kPa?

A 9 g of CH_4 **B** 12 g of H_2O

C 15 g of C_2H_6 **D** 20 g of CO_2

At 25 °C, H_2O is in liquid state and occupy smaller volume than gas
Amount of CH_4 = 0.563 mol Amount of C_2H_6 = 0.5 mol
Amount of CO_2 = 0.455 mol
From the Ideal Gas Equation: pV = nRT, we have V ∝ n.
Hence CH_4 would have the largest volume

2 2.5 dm^3 of methane, CH_4, at a pressure of 1.50 atm and 4.5 dm^3 of ethane, C_2H_6, at a pressure of 1.00 atm were introduced into a 3.0 dm^3 vessel at constant temperature.

What is the final pressure of the gas mixture?

A 0.825 atm **B** 2.50 atm **C** 2.75 atm **D** 3.57 atm

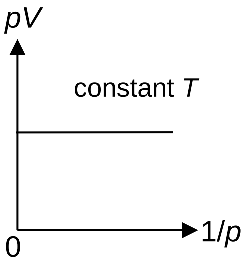
For methane:
$p_1 V_1 = p_2 V_2$
$(1.50)(2.5) = (p_{methane})(3.0)$
$p_{methane}$ = <u>1.25 atm</u>

For ethane:
$p_1 V_1 = p_2 V_2$
$(1.0)(4.5) = (p_{ethane})(3.0)$
p_{ethane} = <u>1.5 atm</u>

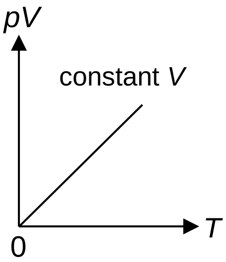
$p_{total} = p_{methane} + p_{ethane}$
= 1.25 + 1.5
= <u>2.75 atm</u>

3 Which graph is not a correct description of the behaviour of a fixed mass of an ideal gas?
(Note: *T* is measured in Kelvin scale.)

A



B



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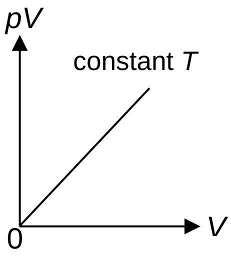

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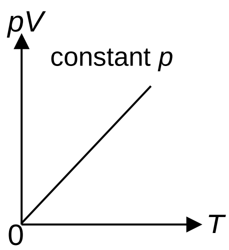


From the Ideal Gas Equation: pV = nRT, we have pV ∝ T (when n is constant).
With increasing T, the value of pV should increase proportionally. Hence graph B and D are correct.
So the value of pV remains constant with constant T, hence a horizontal line for graph A is correct.
This also means a horizontal line for graph C should be expected instead.

4 A vessel contained a gas mixture of nitrogen and oxygen at 30 °C and a pressure of 200 kPa. After all the oxygen has reacted with zinc, the total pressure in the vessel was 150 kPa at 30 °C.

$$2\,Zn\,(s) + O_2\,(g) \rightarrow 2\,ZnO\,(s)$$

What was the molar ratio of nitrogen to oxygen in the mixture initially?

A 1 : 3 **B** 2 : 3 **C** 3 : 1 **D** 3 : 2

After the reaction has occurred, only N_2 is left in the vessel. So $p_{nitrogen}$ = 150 kPa
$p_{total} = p_{nitrogen} + p_{oxygen}$ = 200 kPa
So p_{oxygen} = 50 kPa
Since $p \propto \eta$ at constant temperature ,

$$\frac{n_{nitrogen}}{n_{oxygen}} = \frac{p_{nitrogen}}{p_{oxygen}} = \frac{150}{50} = \frac{3}{1}$$

5 A container contains 1 dm^3 of gas **X** and 1 dm^3 of gas **Y** at s.t.p. Both **X** and **Y** behave as ideal gases and the relative molecular mass of **X** is 4 times that of **Y**.

Which statement is true?

A The amount of gas **X** is 4 times that of **Y**.

B The partial pressure of each gas is 50 kPa.

C The total pressure of the system will decrease when 1 mole of Ne is pumped into the container.

D The volume of the container doubles when the temperature of the system increases from 20 °C to 40 °C under constant pressure.

Statement A is false. At s.t.p, amount of gas ∝ volume of gas, both gases have the same amount of gas particles.

Statement C is false. An increase in the amount of gases at constant volume, should result in an increase in total pressure.

Statement D is false. Under constant pressure,

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

$$\frac{V_2}{40+273} = \frac{V_1}{20+273}$$

$$V_2 = \frac{313}{293} V_1$$

Statement B is true. With same gas volume for A and B at s.t.p, there is equal amount of gas A and B.
Since $p_{total} = p_A + p_B$ = 1 bar (10^5 Pa) @ s.t.p and $p \propto \eta$,
$p_A = p_B$ = ½ (100 kPa) = 50 kPa

6 When magnesium powder reacts with dilute hydrochloric acid, hydrogen gas is evolved.

How would the volume of the hydrogen gas be affected by temperature and pressure?

A It increases with increase in temperature and is independent of pressure.

B It decreases with increase in temperature and is independent of pressure.

C It increases with increase in both temperature and pressure.

D It increases with increase in temperature and decreases with increase in pressure.

From the Ideal Gas Equation: $pV = \eta RT$,
we have $V \propto \frac{1}{p}$ and $V \propto T$ (for a given amount of gas)

7 Which expression gives the pressure exerted by 1.6×10^{-3} mol of N_2 in a container of volume 3.0 dm^3 at 273 °C?

A $\dfrac{1.6 \times 10^{-3} \times 8.31 \times 273}{3.0 \times 10^{-6}}$ Pa

B $\dfrac{1.6 \times 10^{-3} \times 8.31 \times (273 + 273)}{3.0 \times 10^{-6}}$ Pa

C $\dfrac{1.6 \times 10^{-3} \times 8.31 \times 273}{3.0 \times 10^{-3}}$ Pa

D $\dfrac{1.6 \times 10^{-3} \times 8.31 \times (273 + 273)}{3.0 \times 10^{-3}}$ Pa

From the Ideal Gas Equation: $pV = \eta RT$,
We have $p = \frac{\eta RT}{V}$ and the units of V and T must be converted to SI units.

8 Flask **A** contains 1 dm^3 of helium at 2 kPa pressure and flask **B** contains 2 dm^3 of neon at 1 kPa pressure.

If the flasks are connected at constant temperature, what is the final pressure (in kPa)?

A $1\frac{1}{3}$ **B** $1\frac{1}{2}$ **C** $1\frac{2}{3}$ **D** 2

For He:
$p_1 V_1 = p_2 V_2$
$(2)(1) = (p_{He})(3)$
p_{He} = 2/3 kPa

For Ne:
$p_1 V_1 = p_2 V_2$
$(1)(2) = (p_{Ne})(3)$
p_{Ne} = 2/3 kPa

$p_{total} = p_{He} + p_{Ne}$
= 2/3 + 2/3
= 1 1/3 kPa

9 Which set of changes will have the greatest effect on the density of a fixed mass of an ideal gas?

	Pressure	Temperature / K
A	double	halve
B	double	double
C	halve	halve
D	halve	constant

From the Ideal Gas Equation: pV = ηRT,
We have $\rho = \frac{pM}{RT}$, thus when pressure is doubled and temperature is halved, the density of the gas will quadruple.

10 Which statements are correct?

1 The density of an ideal gas at constant pressure is inversely proportional to the temperature.

2 One of the assumptions of the kinetic theory of gases is that gaseous particles are separated by distances which are large compared with their dimensions.

3 Real gases behave most nearly as an ideal gas at high pressures and low temperatures.

A 1, 2 and 3 are correct

B 1 and 2 only are correct

C 2 and 3 only are correct

D 1 only is correct

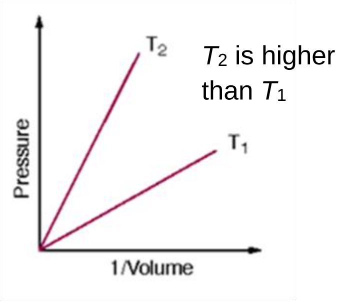
From the Ideal Gas Equation, we have $\rho = \frac{pM}{RT}$. So statement 1 is true.

An ideal gas has gaseous particles with **negligible volumes** compared to the volume of container. Thus the gas particles are far apart. So statement 2 is true.

Real gases behave least as an ideal gas at high pressures and low temperatures.
At high pressure, the gas particles are very close together that there are significant intermolecular forces of attractions & significant volume occupied by the gas particles.
At low temperature, the gas particles have low KE to overcome intermolecular forces of attractions between them. Hence Statement 3 is wrong.

11 Which graphs show the ideal behavior of a gas?

1



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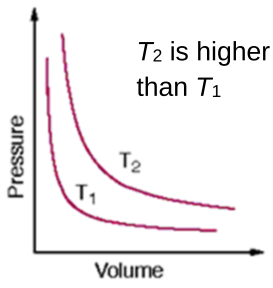

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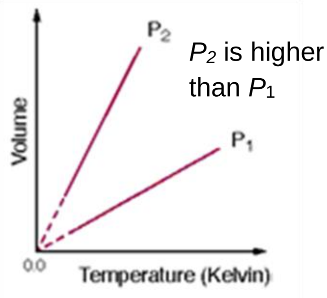


A 1, 2 and 3 are correct **B** 1 and 2 only are correct

C 2 and 3 only are correct **D** 1 only is correct

From the Ideal Gas Equation: $pV = \eta RT$, (for a given amount of gas)
Graph 1 and 2 show the correct pressure – volume relationship, $p \propto \frac{1}{V}$.
At a particular volume, $p \propto T$. This is correctly shown in both graphs 1 and 2.

Graph 3 shows the correct volume – temperature relationship, $V \propto T$.
At a particular temperature, Vol $\propto \frac{1}{p}$. This is incorrectly shown, gradient should be smaller at P_1

12 Two moles of oxygen and one mole of argon are contained in a cylinder with a volume of 10.0 dm^3 at 298 K.

Calculate the total pressure and partial pressure of oxygen.

p_{total} $(10.0 \times 10^{-3}) = (2+1)(8.31)(298)$
p_{total} = <u>742 kPa</u>
$p_{O_2} = (\frac{2}{3})(742) =$ <u>495 kPa</u>

13 A gas canister, used in camping stoves, contains mainly butane under sufficient pressure to cause it to liquefy partially.

(a) State two assumptions of the kinetic theory as applied to ideal gases and use these to explain whether you might expect the gas in the canister to behave as an ideal gas.

Assumptions:
1. A gas is composed of tiny particles that have negligible volumes(size) compared to the volume of the container.
2. There are negligible intermolecular forces of attraction between the gas particles.

Since butane is stored under very high pressure (partially liquefied butane), the butane particles are very close together, experiencing strong forces of attraction and occupy significant volume. So both assumptions are <u>not valid</u> and the gas will not behave ideally.

(b) The canister was connected to a gas syringe and the valve opened slightly to allow some of the gas into the syringe. It was found that 0.200 g of the gas took up a volume of 96.0 cm^3 at a temperature of 20.0 °C and a pressure of 1.01×10^5 Pa.

Calculate the average relative molecular mass of the gas mixture.

$pV = \frac{m}{M}RT$

$(1.01 \times 10^5)(96.0 \times 10^{-6}) = (\frac{0.200}{M})(8.31)(20 + 273)$

$M = 50.2 \text{ g mol}^{-1}$

Thus, the relative molecular mass, $\underline{M_r = 50.2}$ (no units)

(c) Based on safety considerations, explain why the gas canister with residual butane should not be disposed of by burning.

The residual **gas** in canister **expands when temperature increases**.
This rapid increase in gas pressure in canister will cause the **canister to explode**.

14 Cylinders of pressurized carbon dioxide are used to produce carbonated drinks. One such cylinder has an internal volume of 2.5 dm^3, and contains 2.3 kg of carbon dioxide.

(a) Calculate the pressure (in Pa) the carbon dioxide gas would exert inside the cylinder at 25 °C.

Amount of $CO_2 = \frac{2300}{44.0} = 52.27$ mol

$p(2.5 \times 10^{-3}) = (52.27)(8.31)(298)$

$p = \underline{5.18 \times 10^7 \text{ Pa}}$

(b) In fact, the pressure inside the cylinder is 2.2×10^7 Pa under these conditions. Explain why this differs from the pressure you calculated in (a).

Under such high pressure, CO_2 will not behave as an ideal gas. The **intermolecular forces of attraction** between the **gas molecules** holds them **closer together** and **reduces the force of collision** of gas molecules against the walls of container

(c) A 500 cm^3 can of cola has 2.0 g of carbon dioxide dissolved in it under pressure. When the can is opened, carbon dioxide is released to the atmosphere immediately till it goes flat. A saturated solution of carbon dioxide is then left in the can with a concentration of 1.5 g dm^{-3} at a pressure of 1 atm and 25 °C.

Calculate the volume of carbon dioxide that is released to the atmosphere when it goes flat. [Assume carbon dioxide behaves as an ideal gas under these conditions.]

Mass of CO_2 escaped = 2.0 – 1.5(0.500) = 1.25 g

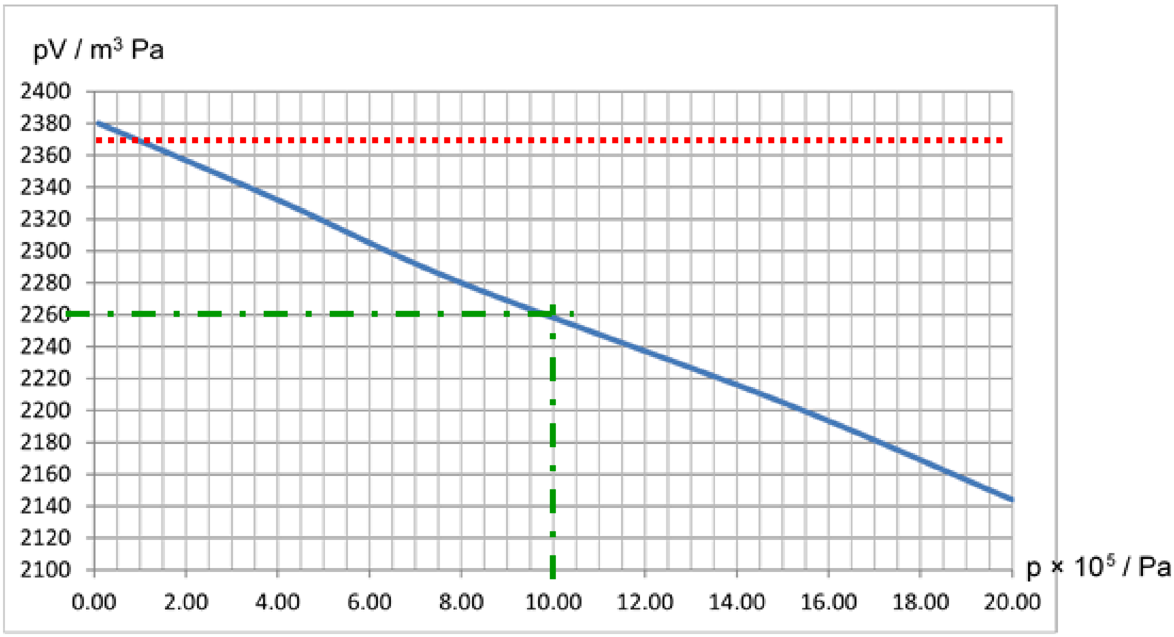
Amount of CO_2 escaped $= \frac{1.25}{44.0} = 0.02841$mol

$(101325)V = (0.02841)(8.31)(298)$

$V = 6.94 \times 10^{-4}$ m^3 (or 694 cm^3)

Do not accept: $V = 0.02841 \times 24.0 = 682$ cm^3 as r.t.p T is 20°C

15 The volume of 1 mole of carbon dioxide was measured at various pressures but at a constant temperature of 285 K. The results obtained were plotted as shown below.



(a) Sketch, on the above axes, a graph to show how the *pV* value should change with pressure for 1 mole of an ideal gas at the same temperature.

From ideal gas equation,
pV = 1 × 8.31 × 285 = 2369
Draw a horizontal line near pV = 2370

(b) Use the graph to calculate the volume of 1 mole of carbon dioxide at a pressure of 10.0 × 10^5 Pa. Calculate the volume at 285 K that the ideal gas equation predicts for this pressure and comment on the difference between the two values.

From the graph when p = 10^6 Pa,
pV = 2260 ⇒ V_{real} = 2.26 × 10^{-3} m^3
From ideal gas equation,
(10.0 × 10^5)V = 1 × 8.31 × 285 ⇒ V_{ideal}= 2.37 × 10^{-3} m^3

The significant intermolecular forces of attractions draw the CO_2 gas molecules closer together and decreases the volume occupied by the gas. Hence V_{real} < V_{ideal}.

16 0.2 mol of radon is stored in a 20.0 dm^3 flask with 0.8 mol of fluorine at 127 °C.

(a) Calculate the total pressure in the flask.

From the Ideal Gas Equation: $pV = \eta RT$,
$p_T (20 \times 10^{-3}) = (0.200 + 0.800)(8.31)(127 + 273)$
$p_T = 166$ kPa

(b) Calculate the partial pressure of radon and fluorine in the flask.

partial pressure of radon = $\frac{0.2}{1.00}(166.2)$ = 33.2 kPa
partial pressure of fluorine = 166.2 – 33.24 = 133 kPa

(c) Radon and xenon are in the same group. A scientist hypothesises that radon has similar chemical reaction as xenon. Xenon can react with fluorine according to the following reaction.

$$Xe(g) + F_2(g) \longrightarrow XeF_2(g)$$

The mixture of radon and fluorine in the flask was sparked at 127 °C. What is the final pressure in the flask if the temperature is maintained at 127 °C?

	Rn	+	F_2	→	RnF_2
initial amt /mol	0.2		0.8		0
final amt /mol	0		0.8 - 0.2 = 0.6		0.2

Total amount of gases = 0.6 + 0.2 = 0.8 mol
$p_T(20 \times 10^{-3}) = 0.8(8.31)(400)$
$p_T = 133$ kPa

Answer keys for MCQ:

1	2	3	4	5	6	7	8	9	10	11
A	C	C	C	B	D	D	A	A	B	B

12) 495 kPa	13b) $M_r = 50.2$	14a) 5.18×10^7 Pa	14c) 6.94×10^{-4} m^3
15b) $2.37 \times 10^{-3} m^3$	16a) 166 kPa	16b) $P_{Rn} = 33.2$ kPa $P_{F2} = 133$ kPa	16c) 133 kPa