

## 1.23 Gas Calculations

Gas calculations at A-level are done in two different ways although both link the volumes of a gas to the amount in moles of the gas. The same amount in moles of any gas will have the same volume under the same conditions of temperature and pressure.

### Molar Gas Volume (OCR +EDEXCEL only)

The simplest way of working out the volume of gas is using the molar gas volume of a gas. This works on the principle that 1 mole of any gas will have the same volume under the same conditions of temperature and pressure. Most often this is given as 24dm<sup>3</sup> which is the volume of a gas at room pressure (100kPa) and room temperature 25°C. Different temperatures and pressures will have different molar gas volumes.

$$\text{Gas Volume (dm}^3\text{)} = \text{amount} \times 24$$

This equation gives the volume of a gas at room pressure (100kPa) and room temperature 25°C.

### Ideal Gas Equation

The ideal gas equation is a slightly more difficult way of working out a gas volume. It can calculate the volume for any temperature and pressure though so is a more useful method.

More detail is given on page 2 about how the ideal gas equation is derived and its limitations.

The biggest problems students have with this equation is choosing and converting to the correct units, so pay close attention to the units.

$$PV = nRT$$

Unit of Pressure (P): Pa  
Unit of Volume (V): m<sup>3</sup>  
Unit of Temp (T): K  
n = moles  
R = 8.31 JK<sup>-1</sup>mol<sup>-1</sup>

#### Converting temperature

$$^{\circ}\text{C} \rightarrow \text{K add 273}$$

**Example 1:** What is the mass of Cl<sub>2</sub> gas that has a pressure of 100kPa, temperature 293K, volume 500cm<sup>3</sup>.  
(R = 8.31 JK<sup>-1</sup>mol<sup>-1</sup>)

$$\begin{aligned} \text{moles} &= PV/RT \leftarrow \begin{array}{l} 100 \text{ kPa} = 100\,000 \text{ Pa} \\ 500 \text{ cm}^3 = 0.0005 \text{ m}^3 \end{array} \\ &= 100\,000 \times 0.0005 / (8.31 \times 293) \\ &= 0.0205 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Mass} &= \text{moles} \times M_r \\ &= 0.0205 \times (35.5 \times 2) \\ &= 1.46 \text{ g} \end{aligned}$$

**Example 2:** 0.15g of a volatile liquid was injected into a sealed gas syringe. The gas syringe was placed in an oven at 70°C at a pressure of 100kPa and a volume of 80cm<sup>3</sup> was measured. What is the M<sub>r</sub> of the volatile liquid ? (R = 8.31 JK<sup>-1</sup>mol<sup>-1</sup>)

$$\begin{aligned} \text{moles} &= PV/RT \leftarrow \begin{array}{l} 100 \text{ kPa} = 100\,000 \text{ Pa} \\ 80 \text{ cm}^3 = 0.00008 \text{ m}^3 \end{array} \\ &= 100\,000 \times 0.00008 / (8.31 \times 343) \\ &= 0.00281 \text{ mol} \end{aligned}$$

$$\begin{aligned} M_r &= \text{mass}/\text{moles} \\ &= 0.15 / 0.00281 \\ &= 53.4 \text{ g mol}^{-1} \end{aligned}$$

## Derivation of the Ideal Gas Equation

The ideal gas equation is a combination of three important gas laws.

**Boyle's Law** - The volume of a fixed amount of gas at a given temperature is inversely proportional to its pressure.

$$V \propto \frac{1}{P} \quad \boxed{P_1V_1 = P_2V_2} \quad \text{at constant } T \text{ \& } n$$

**Charles's Law** - If a fixed amount of a gas is held at constant pressure, the volume of the gas is directly proportional to the temperature of the gas.

$$\boxed{V \propto T} \quad \text{So for a system that changes } T \text{ and } V -$$
$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}} \quad \text{at constant } P \text{ \& } n$$

**Avogadro's Law** - Equal volumes of gas at constant temperature and pressure have equal numbers of molecules.

<b>Boyle's Law</b> $\boxed{V \propto \frac{1}{P}}$	<b>Charles's Law</b> $\boxed{V \propto T}$	<b>Avogadro's Law</b> $\boxed{V \propto n}$
Expressed Mathematically -		
$\boxed{V \propto \frac{Tn}{P}}$		
$\boxed{V = R \left( \frac{nT}{P} \right) \Rightarrow PV = nRT}$		

An ideal gas has the following properties:

1. An ideal gas is considered to be a "point mass". A point mass is a particle so small that its mass is very nearly zero. This means an ideal gas particle has almost no volume.
2. There are no attractive or repulsive forces involved during collisions. Also, the kinetic energy of the gas molecules remains constant since the intermolecular forces are lacking.

These assumptions do not actually hold true for many gases. They work best for the Noble gases. Heavier gases and ones with significant intermolecular forces will deviate from ideal gas behaviour

## Changing the Conditions of a gas

Questions may involve the same amount of gas under different conditions.

### Example 3

40 cm<sup>3</sup> of oxygen and 60 cm<sup>3</sup> of carbon dioxide, each at 298 K and 100 kPa, were placed into an evacuated flask of volume 0.50 dm<sup>3</sup>. What is the pressure of the gas mixture in the flask at 298 K?

There are two approaches to solving this

1. Work out moles of gas using ideal gas equation then put back into ideal gas equation with new conditions
2. Or combine the equation  $n = PV/RT$  as on right

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Can do this as moles of gas do not change

As Temperature is the same can make the above equation  $P_1 V_1 = P_2 V_2$

$$\begin{aligned} P_2 &= P_1 V_1 / V_2 \\ &= 100000 \times 1 \times 10^{-4} / 5 \times 10^{-4} \\ &= 20000 \text{ Pa} \end{aligned}$$

## Questions using the Ideal Gas Equation

(In all questions use the Ideal Gas Equation and use the gas constant  $R = 8.31 \text{ JK}^{-1}\text{mol}^{-1}$ )

- 1.1) Calculate the volume of 1.50 mol of carbon dioxide gas at 298K and 100 kPa.
- 1.2) A sample of ammonia gas occupied a volume of  $0.0278\text{m}^3$  at  $28^\circ\text{C}$  and 103 kPa. Calculate the number of moles of ammonia in the sample.
- 1.3) A gas cylinder, of volume  $5.00 \times 10^{-3} \text{ m}^3$ , contains 348 g of argon gas. Calculate the pressure of the argon gas in the cylinder at a temperature of  $20^\circ\text{C}$
- 1.4) A 0.153g gaseous sample of ammonia occupied a volume of  $2.91 \times 10^{-4} \text{ m}^3$  at a temperature  $T$  and a pressure of 100 kPa. Calculate the number of moles of ammonia present and deduce the value of the temperature  $T$ .
- 1.5) A sample of ethanol vapour,  $\text{C}_2\text{H}_5\text{OH}$  ( $M_r = 46.0$ ), was maintained at a pressure of 100 kPa and at a temperature of 358K. Calculate the volume, in  $\text{cm}^3$ , that 2.78 g of ethanol vapour would occupy under these conditions.
- 1.6) When a sample of liquid, **A**, of mass 0.306 g was vaporised, the vapour was found to occupy a volume of  $1.76 \times 10^{-4} \text{ m}^3$  at a pressure of 110 kPa and a temperature of  $200^\circ\text{C}$ . Calculate the number of moles of **A** in the sample and hence deduce the relative molecular mass of **A**.
- 1.7) Compound **B** is an oxide of sulphur. At 423 K, a gaseous sample of **B**, of mass 0.354 g, occupied a volume of  $148 \text{ cm}^3$  at a pressure of 105 kPa. Calculate the number of moles of **B** in the sample, and hence calculate the relative molecular mass of **B**.
- 1.8) In an experiment, 0.612 mol of hydrogen gas was produced when a sample of magnesium reacted with dilute nitric acid. Calculate the volume that this gas would occupy at 298 K and 98.0 kPa. Include units in your final answer.

## Calculations involving masses, solutions and gases

Commonly in questions converting between quantities of substances reacting we will use more than just mass data. We might have the volume and concentration of a solution, or the volume of a gas. We need to adapt our existing method or reacting masses to include other quantities. Any of the equations below can be used to convert quantities into moles.

### 1. For pure solids, liquids and gases

$$\text{amount} = \frac{\text{mass}}{\text{MolarMass}}$$

Unit of Mass: grams  
Unit of amount : mol

### 2. For Gases

$$\text{Gas Volume (dm}^3\text{)} = \text{amount} \times 24$$

This equation give the volume of a gas at room pressure (1atm) and room temperature 25°C.

Or use the ideal gas equation to work out gas volumes at other temperatures and pressures

$$PV = nRT$$

### 3. For solutions

$$\text{Concentration} = \frac{\text{amount}}{\text{volume}}$$

Unit of concentration: mol dm<sup>-3</sup> or M  
Unit of Volume: dm<sup>3</sup>

### General method

#### Step 1:

Use one of the above equations to convert any given quantity into amount in mol  
Mass → amount  
Volume of gas → amount  
Conc and vol of solution → amount

#### Step 2:

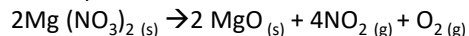
Use balanced equation to convert amount in mol of initial substance into amount in mol of second substance

#### Step 3

Convert amount, in mol, of second substance into quantity question asked for using relevant equation  
e.g. amount, Mr → mass  
Amount gas → vol gas  
amount, vol solution → conc

#### Example 4 : Using Ideal Gas

What is the total volume of gas produced in dm<sup>3</sup> at 333K and 100kPa when 0.651 g of magnesium nitrate decomposes when heated?



Step 1: work out moles of magnesium nitrate

$$\begin{aligned}\text{Moles} &= \text{mass} / \text{Mr} \\ &= 0.651 / 148.3 \\ &= 0.00439 \text{ mol}\end{aligned}$$

Step 2: use balanced equation to give moles of gas produced

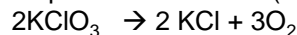
2 moles Mg(NO<sub>3</sub>)<sub>2</sub> : 4NO<sub>2</sub>(g) + O<sub>2</sub>(g) ie 5moles of gas  
So 0.00439 Mg(NO<sub>3</sub>)<sub>2</sub> : 0.01098( 0.00439 x 5/2) moles gas

Step 3: work out volume of gas

$$\begin{aligned}\text{Volume} &= nRT/P \\ &= (0.01098 \times 8.31 \times 333) / 100000 \\ &= 0.000304\text{m}^3 \\ &= 0.303\text{dm}^3\end{aligned}$$

#### Example 5: Using molar gas volume

What volume in cm<sup>3</sup> of oxygen gas would be produced from the decomposition of 0.532 g of potassium chlorate(V)?



Step 1: work out amount, in mol, of potassium chlorate(V)?

$$\begin{aligned}\text{amount} &= \text{mass} / \text{Mr} \\ &= 0.532 / 122.6 \\ &= 0.00434 \text{ mol}\end{aligned}$$

Step 2: use balanced equation to give amount in mol of O<sub>2</sub>

2 moles KClO<sub>3</sub> : 3 moles O<sub>2</sub>  
So 0.00434 HNO<sub>3</sub> : 0.00651moles O<sub>2</sub>

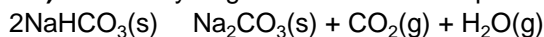
Step 3: work out volume of O<sub>2</sub>

$$\begin{aligned}\text{Gas Volume (dm}^3\text{)} &= \text{amount} \times 24 \\ &= 0.00651 \times 24 \\ &= 0.156 \text{ dm}^3 \\ &= 156 \text{ cm}^3\end{aligned}$$

## Questions combining masses, equations and ideal gas volumes

(In all questions use the Ideal Gas Equation and use the gas constant  $R = 8.31 \text{ JK}^{-1}\text{mol}^{-1}$ )

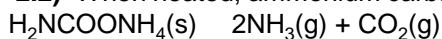
**2.1)** Sodium hydrogencarbonate decomposes on heating as shown in the equation below.



A sample of  $\text{NaHCO}_3$  was heated until completely decomposed. The  $\text{CO}_2$  formed in the reaction occupied a volume of  $472 \text{ cm}^3$  at  $1.00 \times 10^5 \text{ Pa}$  and  $298 \text{ K}$ .

Calculate the number of moles of  $\text{CO}_2$  formed in this decomposition and the mass of  $\text{NaHCO}_3$  that was decomposed.

**2.2)** When heated, ammonium carbamate,  $\text{H}_2\text{NCOONH}_4$ , decomposes as shown below.

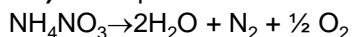


In a closed container, a  $6.50 \text{ g}$  sample of ammonium carbamate was heated. The solid decomposed completely into ammonia and carbon dioxide at  $463 \text{ K}$  and  $99.7 \text{ kPa}$ .

(i) Calculate the number of moles of ammonium carbamate used and the total number of moles of gas produced.

(ii) Calculate the total volume of gas produced at  $463 \text{ K}$  and  $99.7 \text{ kPa}$ . Include units in your answer.

**2.3)** A sample of ammonium nitrate decomposed on heating as shown in the equation below.

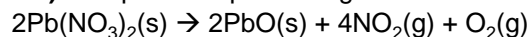


On cooling the resulting gases to  $298 \text{ K}$ , the volume of nitrogen and oxygen together was found to be  $0.0550 \text{ m}^3$  at a pressure of  $98.0 \text{ kPa}$ .

(i) Calculate the total number of moles of nitrogen and oxygen formed.

(ii) Deduce the number of moles of ammonium nitrate decomposed and hence calculate the mass of ammonium nitrate in the sample.

**2.4)** An equation representing the thermal decomposition of lead(II) nitrate is shown below.



A sample of lead(II) nitrate was heated until the decomposition was complete. At a temperature of  $450 \text{ K}$  and a pressure of  $100 \text{ kPa}$ , the total volume of the gaseous mixture produced was found to be  $1.35 \times 10^{-4} \text{ m}^3$ .

(i) Calculate the total number of moles of gas produced in this decomposition.

(ii) Deduce the number of moles, and the mass, of  $\text{NO}_2$  present in this gaseous mixture.

(iii) Calculate the original mass of the lead nitrate that was decomposed

**2.5)** Potassium nitrate,  $\text{KNO}_3$ , decomposes on strong heating, forming oxygen and solid  $\text{KNO}_2$  as the only products.

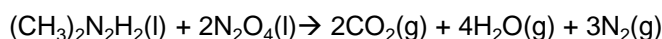
A  $1.00 \text{ g}$  sample of  $\text{KNO}_3$  was heated strongly until fully decomposed

(i) Calculate the number of moles of  $\text{KNO}_3$  in the  $1.00 \text{ g}$  sample.

(ii) Calculate the number of molecules of  $\text{O}_2$  produced (The Avogadro constant  $L = 6.022 \times 10^{23}$ )

(iii) Calculate the volume of oxygen gas produced in  $\text{cm}^3$  at  $298 \text{ K}$  and  $100 \text{ kPa}$

**2.6)** UDMH,  $(\text{CH}_3)_2\text{N}_2\text{H}_2$ , can be used in combination with dinitrogen tetroxide,  $\text{N}_2\text{O}_4$  as a fuel in thruster rockets. A thruster rocket contained a total mass of propellant (UDMH and,  $\text{N}_2\text{O}_4$ , together) of  $240 \text{ kg}$  mixed in a 1:2 molar ratio. The equation below shows the reaction



(i) Calculate the mass of UDMH in the propellant mixture.

(ii) Calculate the total number of moles of gas produced from the complete reaction of this mass of UDMH with dinitrogen tetroxide,

(iii) Calculate the total volume of product gases formed from the reaction of this mass of UDMH with dinitrogen tetroxide at a temperature of  $-9.0 \text{ }^\circ\text{C}$  and a pressure of  $550 \text{ Pa}$ .

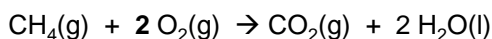
## Reacting Volumes of Gas

Equal volumes of any gases measured under the same conditions of temperature and pressure contain equal numbers of molecules (or atoms if the gas is monatomic)

1 mole of any gas at room pressure (1atm) and room temperature 25°C will have the volume of 24dm<sup>3</sup>

Volumes of gases reacting in a balanced equation can be calculated by simple ratio

**Example 6** If one burnt 500 cm<sup>3</sup> of methane at 1atm and 300K what volume of Oxygen would be needed and what volume of CO<sub>2</sub> would be given off under the same conditions?

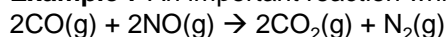


1 mole      2 mole      1 mole

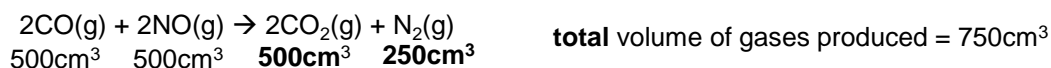
500cm<sup>3</sup>    1dm<sup>3</sup>      500cm<sup>3</sup>

Simply multiply  
gas volume x2

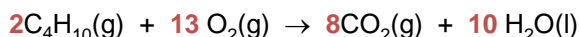
**Example 7** An important reaction which occurs in the catalytic converter of a car is



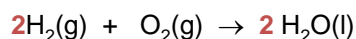
In this reaction, when 500 cm<sup>3</sup> of CO reacts with 500 cm<sup>3</sup> of NO at 650 °C and at 1 atm. Calculate the **total** volume of gases produced at the same temperature and pressure ?



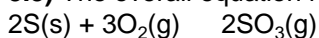
**3.1)** If one burnt 1 dm<sup>3</sup> of butane, what volume of Oxygen would be needed and what volume of CO<sub>2</sub> would be given off?



**3.2)** What volume of **air** is needed for the complete combustion of 500 cm<sup>3</sup> of hydrogen? (Assume air contains 20% oxygen)

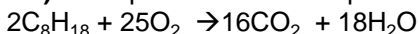


**3.3)** The overall equation for the reaction between sulfur and oxygen to form sulfur trioxide is shown below.



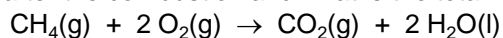
2 dm<sup>3</sup> of O<sub>2</sub>(g) reacted completely with excess sulfur. What volume, in dm<sup>3</sup>, of sulfur trioxide would form?

**3.4)** The equation for the complete combustion of octane is



The volume of 1 mol of any gas (measured at room temperature and pressure) is 24 dm<sup>3</sup>. Calculate the volume of oxygen (measured at room temperature and pressure) required for the complete combustion of 5 mol of octane.

**3.5)** 90 cm<sup>3</sup> of methane was reacted with 450 cm<sup>3</sup> of oxygen. How much of each gas is present after the combustion and what is the total volume of all gases at the end?



**3.6)** 20cm<sup>3</sup> of an unknown hydrocarbon needed 120 cm<sup>3</sup> of oxygen for complete combustion. 80cm<sup>3</sup> of CO<sub>2</sub> was produced. All volumes were measured at rtp. Find the formula of the hydrocarbon.

## Answers

1.1  $V = 0.0371\text{m}^3$

1.2  $n = 1.14\text{ mol}$

1.3  $P = 4250\text{kpa}$

1.4  $T = 389\text{K}$

1.5  $V = 1800\text{cm}^3$

1.6  $M_r = 62.1$

1.7  $M_r = 80.1$

1.8  $V = 0.0155\text{m}^3$

2.1 moles  $\text{CO}_2 = 0.01906$  mass  $\text{NaHCO}_3 = 3.2\text{g}$

2.2 (i) moles of ammonium carbamate = 0.083 moles gas = 0.25

(ii)  $V = 9.65 \times 10^{-3}\text{ m}^3$

2.3 (i) 2.177 moles total gas

(ii) Moles ammonium nitrate 1.45mol mass = 116g

2.4 (i) moles = 0.00361mol

(ii) Moles  $\text{NO}_2 = 0.002888$  Mass  $\text{NO}_2 = 0.133\text{g}$

(iii) Mass lead nitrate 0.478g

2.5 (i)  $9.89 \times 10^{-3}\text{ mol}$

(ii)  $2.98 \times 10^{21}$

(ii)  $122\text{ cm}^3$

2.6 (i) mass UDMH 59kg

(ii) Moles gas 8850 mol

(iii) Volume gas  $35300\text{m}^3$

3.1  $6.5\text{ dm}^3\text{ O}_2$   $4\text{ dm}^3\text{ CO}_2$

3.2  $1.25\text{ dm}^3$

3.3  $1.33\text{ dm}^3$

3.4  $1500\text{ dm}^3$

3.5  $270\text{cm}^3\text{ O}_2$   $90\text{cm}^3\text{ CO}_2$  total  $360\text{cm}^3$

3.6  $\text{C}_4\text{H}_8$