Master

Kirkcaldy High School

Higher Chemistry

Unit 3 – part 1

Getting the Most from Reactants

Name: ____________________

Class: ____________________

Teacher: ____________________

Assessment Page

Homework

Homework title	Date	Mark/Total Mark
		/

Notes/comments

Check tests

Test title	Date	Mark/Total Mark
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Notes/comments

Date: _______________

Measuring Efficiency of Chemical Reactions

Overarching question(s) for this topic

- How can we measure and calculate the efficiency of chemical reactions to maximise profit and reduce environmental harm?

Industrial Processes

Industrial processes are designed to maximise profit and minimise the impact on the environment.

Key Factors in Industrial Process Design:

1. **Feedstock availability, sustainability, and cost** – The process relies on raw materials that are easily sourced, renewable (to ensure long-term supply), and cost-effective. Expensive or scarce feedstocks increase production costs and complexity.

    Raw materials are unprocessed natural resources used to make finished products.

    Examples include iron ore for steel and crude **oil** for fuels and plastics.

2. **Recycling opportunities** – Processes that allow materials to be reused or recycled reduce both **waste** and production costs, making the process more efficient and environmentally friendly.

3. **Energy requirements** – Reducing energy consumption is crucial for cutting operational costs and lowering the environmental footprint. Energy-efficient designs are more sustainable and competitive.

4. **Marketability of by-products** – By-products that can be sold or used in other processes add value and reduce waste, improving the overall profitability of the operation.

5. **Product yield** – Maximising the amount of useful product from the raw materials improves efficiency and cost-effectiveness, ensuring that demand can be met with minimal waste.

Environmental Considerations:

1. **Minimizing waste** – Waste reduction is a primary goal, as it cuts disposal costs and environmental impact.

2. **Avoiding toxic substances** – Processes are designed to prevent the use or production of harmful chemicals to protect both human health and ecosystems.

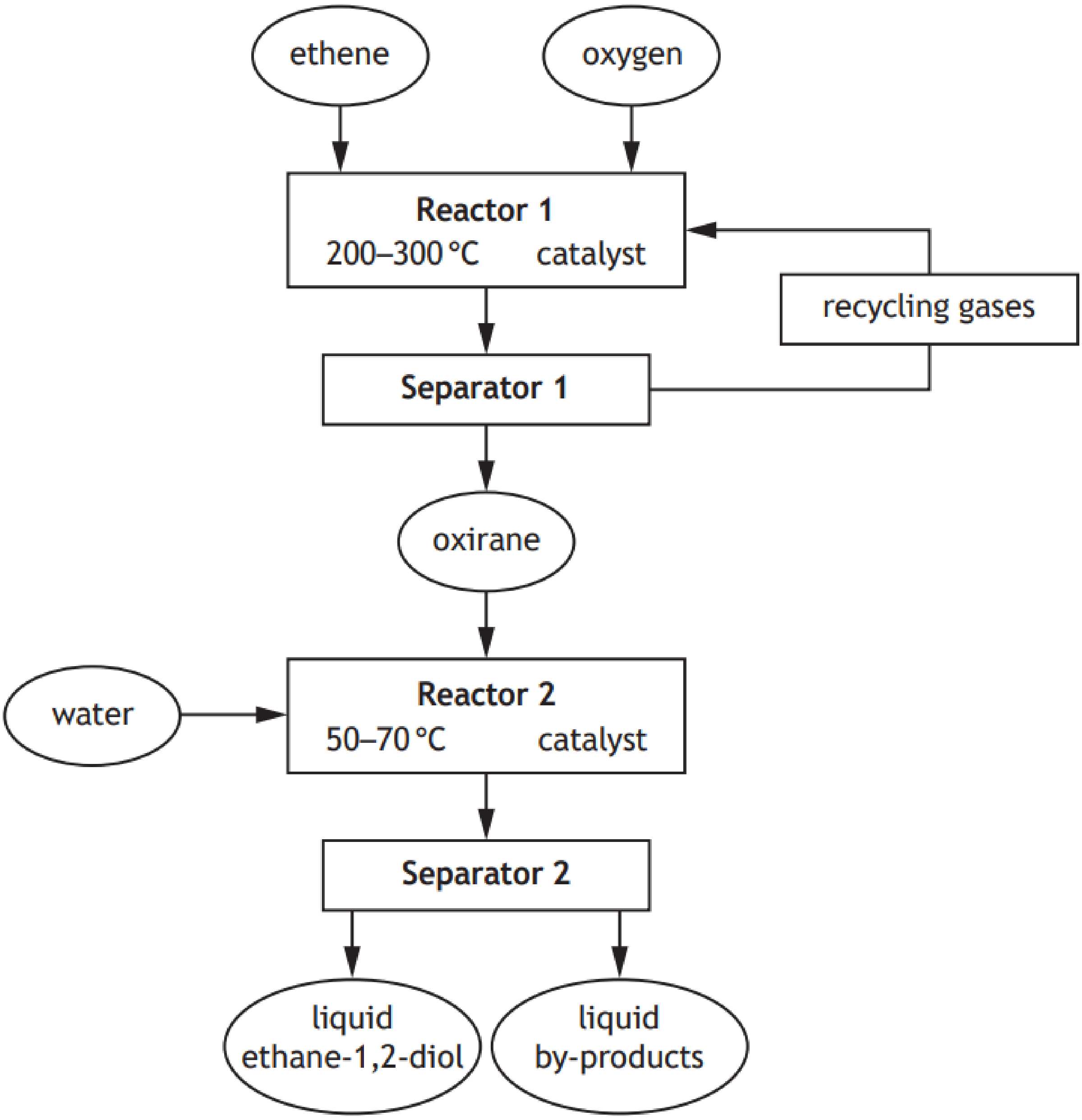
3. **Designing biodegradable products** – When applicable, products are created to naturally break down over time, reducing long-term environmental harm. This is particularly relevant for single-use products or packaging.

Industrial Process Flow Charts

Ethane-1,2-diol can be made from ethene.

The flow chart of an industrial process to produce ethane-1,2-diol is shown.

Annotate the flow chart to show how profit was maximised, add suggestions on how they could improve this further.



Percentage Yield

For a particular set of reaction conditions, the percentage yield provides a measure of the degree to which the **limiting reactant** is converted into the **desired product**.

The percentage yield can be calculated using the following equation (also on page 4):

$$\%\,Yield = \frac{actual\ yield}{theoretical\ Yield} \times 100$$

The **'theoretical yield'** is the quantity of desired product obtained, assuming full conversion of the limiting reactant, as calculated from the **balanced equation**.

The **'actual yield'** is the **quantity** of the **desired product** formed under the reaction conditions.

Your teacher will demonstrate the method for calculating the % yield below:

A student used 2·5 g of ethanol and a slight excess of ethanoic acid to produce 2·9 g of ethyl ethanoate.

ethanol + ethanoic acid ⇌ ethyl ethanoate + water

mass of one mole = 46·0 g	mass of one mole = 88·0 g

(One mole of ethanol reacts with one mole of ethanoic acid to produce one mole of ethyl ethanoate.)

Calculate the percentage yield of ethyl ethanoate.

Complex example

Ammonia is manufactured from hydrogen and nitrogen by the Haber Process.

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3$$

During a reaction, 80 kg of ammonia is produced from 60 kg of hydrogen, calculate the percentage yield.

Questions (in your jotter – answers on the next page)

1. Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

   $$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

   If 24 g of magnesium chloride is produced from 15 g of magnesium, calculate the percentage yield. (Mg GFM = 24.3 g/mol, $MgCl_2$ GFM = 95.3 g)

2. Copper reacts with sulfur to form copper(I) sulfide.

   $$2Cu(s) + S(s) \rightarrow Cu_2S(s)$$

   If 100 g of copper(I) sulfide is produced from 90 g of copper, calculate the percentage yield. (Cu GFM = 63.5 g/mol,Cu_2S GFM = 159.1 g/mol)

3. Calcium carbonate decomposes to form calcium oxide and carbon dioxide.

   $$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

   If 32 g of calcium oxide is produced from 60 g of calcium carbonate, calculate the percentage yield. ($CaCO_3$ GFM = 100.1 g/mol, CaO GFM = 56.1 g/mol)

4. Zinc reacts with hydrochloric acid to produce zinc chloride and hydrogen gas.

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

If 50 g of zinc chloride is produced from 30 g of zinc, calculate the percentage yield. (Zn GFM = 65.4 g/mol, $ZnCl_2$ GFM = 136.4 g/mol)

5. Sulfur dioxide reacts with oxygen to form sulfur trioxide.

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

If 72 g of sulfur trioxide is produced from 60 g of sulfur dioxide, calculate the percentage yield. (SO_2 GFM = 64.1 g/mol, SO_3 GFM = 80.1 g/mol)

6. Aluminium reacts with oxygen to form aluminium oxide.

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

If 83 g of aluminium oxide is produced from 50 g of aluminium, calculate the percentage yield. (Al GFM = 27 g/mol, Al_2O_3 GFM = 102 g/mol)

7. Phosphorus reacts with chlorine to produce phosphorus pentachloride.

$$P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s)$$

If 250 g of phosphorus pentachloride is produced from 100 g of phosphorus, calculate the percentage yield. (P_4 GFM = 124 g/mol, PCl_5 GFM = 208.5 g/mol)

8. Iron reacts with oxygen to produce iron(III) oxide (rust).

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

If 3.5 kg of iron(III) oxide is produced from 2.5 kg of iron, calculate the percentage yield. (Fe GFM = 55.8 g/mol, Fe_2O_3 GFM = 159.6 g/mol)

9. Sulfur reacts with fluorine to form sulfur hexafluoride.

$$S(s) + 3F_2(g) \rightarrow SF_6(g)$$

If 4 kg of sulfur hexafluoride is produced from 2 kg of sulfur, calculate the percentage yield. (S GFM = 32.1 g/mol, SF_6 GFM = 146.1 g/mol)

Q	A
1	40.8%
2	88.7%
3	95.2%
4	79.9%
5	96.0%
6	87.9%
7	37.20%
8	97.90%
9	43.90%

Calculations given % Yield

Your teacher will demonstrate how to perform this calculation:

In the early 1900s, phenol was produced by the following reaction.

$$C_6H_6 + H_2SO_4 + 2NaOH \rightarrow C_6H_5OH + Na_2SO_3 + 2H_2O$$

benzene	phenol
mass of one mole = 78·0 g	mass of one mole = 94·0 g

Calculate the mass of phenol, in kg, produced from 117 kg of benzene if the percentage yield is 90%.

Hydrogen gas reacts with nitrogen to form ammonia in the Haber process.

$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

The mass of ammonia produced in this reaction was 160 kg, with a percentage yield of 80%, calculate the mass of hydrogen used.

Questions (in your jotter)

1. Hydrogen gas reacts with nitrogen to form ammonia in the Haber process.

   $$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

   During a reaction 30 g of hydrogen is used in the reaction, and the percentage yield of ammonia is 80%, calculate the mass of ammonia produced. (H_2 GFM = 2 g/mol, NH_3 GFM = 17 g/mol)

2. Ethanol is dehydrated to produce ethene.

   $$C_2H_5OH(l) \rightarrow C_2H_4(g) + H_2O(g)$$

   During a reaction 100 g of ethanol is used and the percentage yield of ethene is 75%, calculate the mass of ethene produced.

3. Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

   $$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

   If 200 g of magnesium is used, and the percentage yield of magnesium chloride is 90%, what is the actual yield of magnesium chloride in grams?

4. Calcium carbonate decomposes to form calcium oxide and carbon dioxide.

   $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

   If the actual yield of calcium oxide is 425 g, and the percentage yield is 85%, what mass of calcium carbonate was used?

5. Ethanol is dehydrated to produce ethene.

   $C_2H_5OH(l) \rightarrow C_2H_4(g) + H_2O(g)$

   If the actual yield of ethene is 75 g, and the percentage yield is 75%, what mass of ethanol was used?

6. Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

   $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

   If the actual yield of magnesium chloride is 180 g, and the percentage yield is 90%, what mass of magnesium was used?

Calculating the cost of reactant given % yield

Given costs for the reactants, a percentage yield can be used to calculate the cost of reactant(s) required to produce a given mass of product.

Your teacher will demonstrate how to perform these calculations:

A student prepared a sample of methyl cinnamate from cinnamic acid and methanol.

cinnamic acid + methanol → methyl cinnamate + water

mass of one mole = 148 g	mass of one mole = 162 g

The student obtained 3·7 g of methyl cinnamate from 6·5 g of cinnamic acid. Calculate the percentage yield.

The student wanted to scale up the experiment to make 100 g of methyl cinnamate. Cinnamic acid costs £35·00 per 250 g. Calculate the cost of cinnamic acid needed to produce 100 g of methyl cinnamate.

Atom Economy

The atom economy measures the **proportion** of the total mass (atoms) of all starting materials converted into the desired product in the balanced equation.

It is a measure of the efficiency of the reaction based on the number of products formed.

By definition reaction that **only** produces the **desired product** and no waste products has **100% atom economy**.

Reactions with a **high** proportion of **waste products** have a **low** % atom economy.

The following equation is used to calculate the % atom economy.

$$\%\ atom\ economy = \frac{mass\ of\ desired\ product}{total\ mass\ of\ reactants} \times 100$$

Your teacher will demonstrate how to perform the calculation below:

Carbon can combine with oxygen to make carbon monoxide, CO.
Carbon monoxide is used in the production of iron from iron(III) oxide.

$$\begin{array}{ccccccc} Fe_2O_3(s) & + & 3CO(g) & \rightarrow & 2Fe(\ell) & + & 3CO_2(g) \\ GFM = 159{\cdot}6\ g & & GFM = 28{\cdot}0\ g & & GFM = 55{\cdot}8\ g & & GFM = 44{\cdot}0\ g \end{array}$$

Calculate the atom economy for the production of iron.

Methane reacts with steam to produce hydrogen.

$$\begin{array}{ccccccc} CH_4(g) & + & H_2O(g) & \rightarrow & CO(g) & + & 3H_2(g) \\ GFM = 16\ g & & GFM = 18\ g & & GFM = 28\ g & & GFM = 2\ g \end{array}$$

Calculate the atom economy for the formation of hydrogen.

An equation for the formation of silicon nitride is shown.

$3SiCl_4$	+	$16NH_3$	$\rightarrow$	Si_3N_4	+	$12NH_4Cl$
mass of one mole = 170·1 g		mass of one mole = 17·0 g		mass of one mole = 140·3 g		mass of one mole = 53·5 g

Calculate the atom economy for the formation of silicon nitride.

$TiCl_4$	+	$2Mg$	$\rightarrow$	$2MgCl_2$	+	Ti
mass of one mole = 189·9 g		mass of one mole = 24·3 g		mass of one mole = 95·3 g		mass of one mole = 47·9 g

The atom economy for the production of titanium in the above equation is equal to

A $\dfrac{47{\cdot}9}{189{\cdot}9 + 24{\cdot}3} \times 100$

B $\dfrac{47{\cdot}9}{189{\cdot}9 + (2 \times 24{\cdot}3)} \times 100$

C $\dfrac{95{\cdot}3 + 47{\cdot}9}{189{\cdot}9 + 24{\cdot}3} \times 100$

D $\dfrac{(2 \times 47{\cdot}9)}{189{\cdot}9 + 24{\cdot}3} \times 100$

Iron can be produced from iron(III) oxide.

$2Fe_2O_3(s)$	+	$3C(s)$	$\rightarrow$	$4Fe(s)$	+	$3CO_2(g)$
GFM = 159·6 g		GFM = 12·0 g		GFM = 55·8 g		GFM = 44·0 g

The atom economy for the production of iron is

A 69·9%

B 62·8%

C 58·2%

D 32·5%.

Limiting and Excess

During any given reaction with two (or more) reactants, there will one reactant that will **fully react**, named the **limiting reactant**, and the other reactant will not fully react, known as the **excess reactant**.

By calculating the moles of each reactant, the limitin and excess reactant can be determined.

Your teacher will show you how to perform this calculation:

Magnesium carbonate reacts with nitric acid.

$$MgCO_3(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(\ell) + CO_2(g)$$

0·05 mol of magnesium carbonate was added to a solution containing 0·06 mol of nitric acid.

Which of the following statements is true?

A 0·05 mol of carbon dioxide is produced

B 0·06 mol of magnesium nitrate is produced

C Magnesium carbonate is in excess by 0·02 mol

D Nitric acid is in excess by 0·01 mol

$CaCO_3(s)$	+	$2HNO_3(aq)$	$\rightarrow$	$Ca(NO_3)_2\,(aq)$	+	$CO_2(g)$	+	$H_2O(\ell)$
Mass of 1 mol = 100 g				Mass of 1 mol = 164 g				

2·00 g of calcium carbonate ($CaCO_3$) was reacted with 200 cm^3 of 0·1 mol l^{-1} nitric acid (HNO_3).

Take the volume of 1 mole of carbon dioxide to be 24 litres.

In the reaction

A $CaCO_3$ is the limiting reactant

B an excess of 0·1 mol of nitric acid remains at the end of the reaction

C 1·64 g of calcium nitrate is produced by the reaction

D 480 cm^3 of carbon dioxide is produced by the reaction.

0·40 g of sodium sulfite, Na_2SO_3, is reacted with 50 cm^3 of dilute hydrochloric acid, concentration 1·0 mol l^{-1}.

$$Na_2SO_3(s) + 2HCl(aq) \rightarrow H_2O(\ell) + 2NaCl(aq) + SO_2(g)$$

mass of
one mole
= 126·1 g

Show, by calculation, that sodium sulfite is the limiting reactant.

A preservative added to some fizzy drinks is made by reacting sorbic acid and potassium hydroxide.

In an experiment, 7 g of sorbic acid, $C_6H_8O_2$, is reacted with 250 cm^3 of potassium hydroxide solution, concentration 0.5 mol l^{-1}.

$$C_6H_8O_2(s) + KOH(aq) \rightarrow H_2O(\ell) + C_6H_7O_2K(aq)$$

$GFM = 112$ g

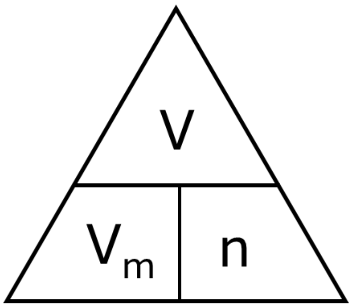
Show, by calculation, that sorbic acid is the limiting reactant.

Molar Volume

The number of of **moles** of a **gas** can be found by using its **volume** and **molar volume**.

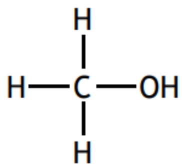
The **molar volume** (V_m) of a gas is the **volume one mole of gas** takes up at a given **temperature** and **pressure**. This is measured in **L/mol** and is often around 24 L/mol.

A calculation triangle can be used for this and calculatations from balanced equations can be performed.



Your teacher will demonstrate how to perform this calculation:

Methanol can be used as a fuel, in a variety of different ways.



(a) An increasingly common use for methanol is as an additive in petrol.

Methanol has been tested as an additive in petrol at 118 g per litre of fuel.

Calculate the volume of carbon dioxide, in litres, that would be released by combustion of 118 g of methanol.

$$2CH_3OH(\ell) \quad + \quad 3O_2(g) \quad \rightarrow \quad 2CO_2(g) \quad + \quad 4H_2O(\ell)$$

(Take the molar volume of carbon dioxide to be 24 litres mol^{-1}).

Phosphine (PH_3) is used as an insecticide in the storage of grain.

Phosphine can be produced by the reaction of water with aluminium phosphide

$$AlP(s) + 3H_2O(\ell) \longrightarrow PH_3(g) + Al(OH)_3(aq)$$

2·9 kg of aluminium phosphide were used in a phosphine generator.

Calculate the volume of phosphine gas, in litres, that would have been produced.

(Take the volume of 1 mole of phosphine to be 24 litres).

Hydrogen peroxide decomposes to form water and oxygen.

$$2H_2O_2(\ell) \rightarrow 2H_2O(\ell) + O_2(g)$$

A dishwasher tablet produces 0·051 g of hydrogen peroxide (mass of one mole = 34 g).

Calculate the volume of oxygen that would be produced when 0·051 g of hydrogen peroxide decomposes.

Take the volume of 1 mole of oxygen gas to be 24 litres.

A scientist prepared a sample of lithium nitride by reacting 0·9 litres of nitrogen gas with 0·5 g of lithium.

$$6Li(s) + N_2(g) \rightarrow 2Li_3N(s)$$

$$GFM = 6{\cdot}9\ g$$

Determine, by calculation, which of the reactants was in excess.

Take the volume of 1 mole of nitrogen gas to be 24 litres.

(Clearly show your working for the calculation.)

Comparing moles/volume

At a given temperature and pressure we assume that the molar volume of all gases are equal. Therefore, if the moles of gas are equal the volume is equal.

This means we can also use **molar ratio coefficients** with the **volumes themselves** without converting to moles. This can also be used in the context of **limiting** and **excess** gases. If they ask for the **total volume of gases** after the reaction is complete, the **excess gas** must be **added** to the total volume.

Which of the following gas samples has the same volume as 16.0 g of oxygen?

(All volumes are measured at the same temperature and pressure)

A 21.0 g of carbon monoxide

B 44.0 g of carbon dioxide

C 46.0 g of nitrogen dioxide

D 46.0 g of dinitrogen tetroxide

Under the same conditions of temperature and pressure, which of the following gases would occupy the largest volume?

A 0·20 g of hydrogen

B 0·44 g of carbon dioxide

C 0·60 g of neon

D 0·80 g of argon

Which of the following gases has the same volume as 128·2 g of sulphur dioxide?

A 2·0 g hydrogen

B 8·0 g helium

C 32·0 g oxygen

D 80·8 g of neon.

With limiting and excess

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(\ell)$$

ethyne

What volume of gas would be produced by the complete combustion of 100 cm^3 of ethyne gas?

All volumes were measured at atmospheric pressure and room temperature.

A 200 cm^3

B 300 cm^3

C 400 cm^3

D 800 cm^3

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

How many litres of nitrogen dioxide gas could theoretically be obtained in the reaction of 1 litre of nitrogen monoxide gas with 2 litres of oxygen gas?

(All volumes are measured under the same conditions of temperature and pressure.)

A 1

B 2

C 3

D 4

20 cm^3 of butane is burned in 150 cm^3 of oxygen.

$$C_4H_{10}(g) + 6\tfrac{1}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$$

What is the total volume of gas present after complete combustion of the butane?

A 80 cm^3

B 100 cm^3

C 180 cm^3

D 200 cm^3

Past Papers

S & B 1st 20	**2015**	**2016**	**2017**	**2018**	**2019**	**2021**	**2022**	**2023**
MC		13	6,14	12,13,14	13-15,17,18	19,20	19	23
S2	3,7a,12aii	3b, 7ciii, 8bii	5aii,9bii	2cii, 3cii-iii, 7b, 11c, 12b	1ai, 6aiii,6bii, 8b-c, 10b-c	2b, 4ci, 4eiiiA, 6bii, 7a, 8bii, 9dii	2a, 2bi, 7c, 8b	2bii, 5ai, 6aii-iii, 6ciiB, 10c

TEAMS: Check Test – Unit 1: Key Area 3a