



UNIVERSITY OF CAMBRIDGE INTERNATIONAL EXAMINATIONS
General Certificate of Education Ordinary Level

CANDIDATE
NAME

CENTRE
NUMBER

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CHEMISTRY

5070/04

Paper 4 Alternative to Practical

May/June 2009

1 hour

Candidates answer on the Question Paper.

No Additional Materials are required.

READ THESE INSTRUCTIONS FIRST

Write your Centre number, candidate number and name on all the work you hand in.
Write in dark blue or black pen.
You may use a soft pencil for any diagrams, graphs or rough working.
Do not use staples, paper clips, highlighters, glue or correction fluid.
DO NOT WRITE IN ANY BARCODES.

Answer **all** questions.
Write your answers in the spaces provided in the Question Paper.

The number of marks is given in brackets [] at the end of each question or part question.
At the end of the examination, fasten all your work securely together.

For Examiner's Use

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This document consists of **15** printed pages and **1** blank page.



- 1 A student did the following experiment to find the formula of magnesium oxide.

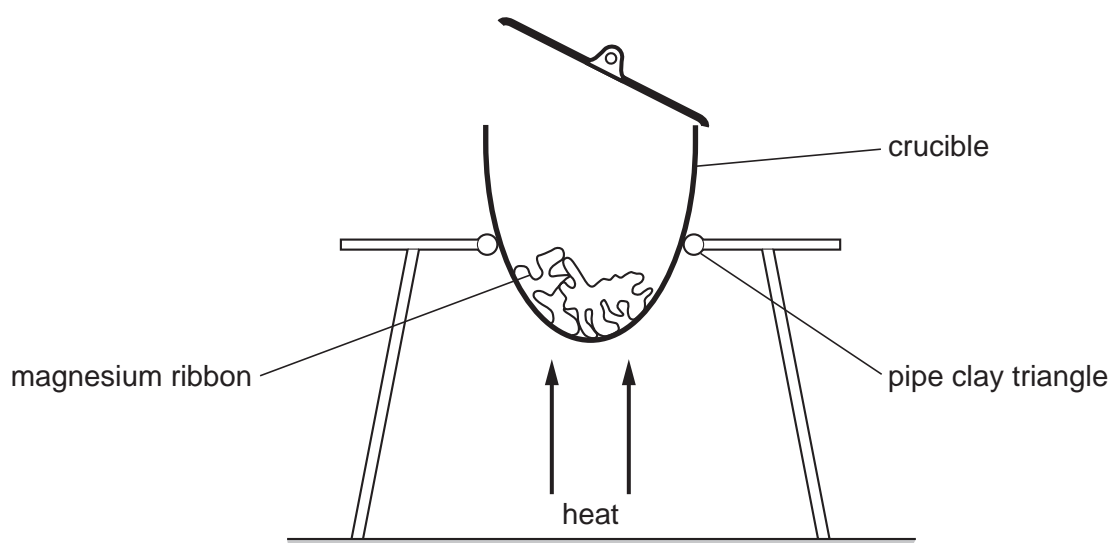
A 10 cm length of magnesium ribbon was loosely coiled and placed in a previously weighed crucible which was then reweighed.

$$\begin{array}{rcl} \text{mass of crucible + magnesium} & = & 13.08 \text{ g} \\ \text{mass of crucible} & = & 12.72 \text{ g} \end{array}$$

- (a) Calculate the mass of magnesium.

..... g [1]

The crucible was placed on a pipe clay triangle and heated strongly for several minutes. During the heating the crucible lid was lifted and replaced several times. The magnesium was converted into magnesium oxide.



- (b) (i) Why was it necessary for a lid to be placed on the crucible during heating?

..... [1]

- (ii) Why was the lid lifted during heating?

..... [1]

- (c) Describe the appearance of

- (i) magnesium,

..... [1]

- (ii) magnesium oxide.

..... [1]

After cooling, the crucible was weighed. It was then reheated, cooled and reweighed.

final mass of crucible + magnesium oxide = 13.32 g

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(d) Why was the crucible re-heated?

..... [1]

(e) (i) Calculate the mass of magnesium oxide.

..... g [1]

(ii) Calculate the mass of oxygen that reacted with the magnesium.

..... g [1]

(f) Using your answers to **(a)** and **(e)(ii)**, calculate the formula of magnesium oxide.

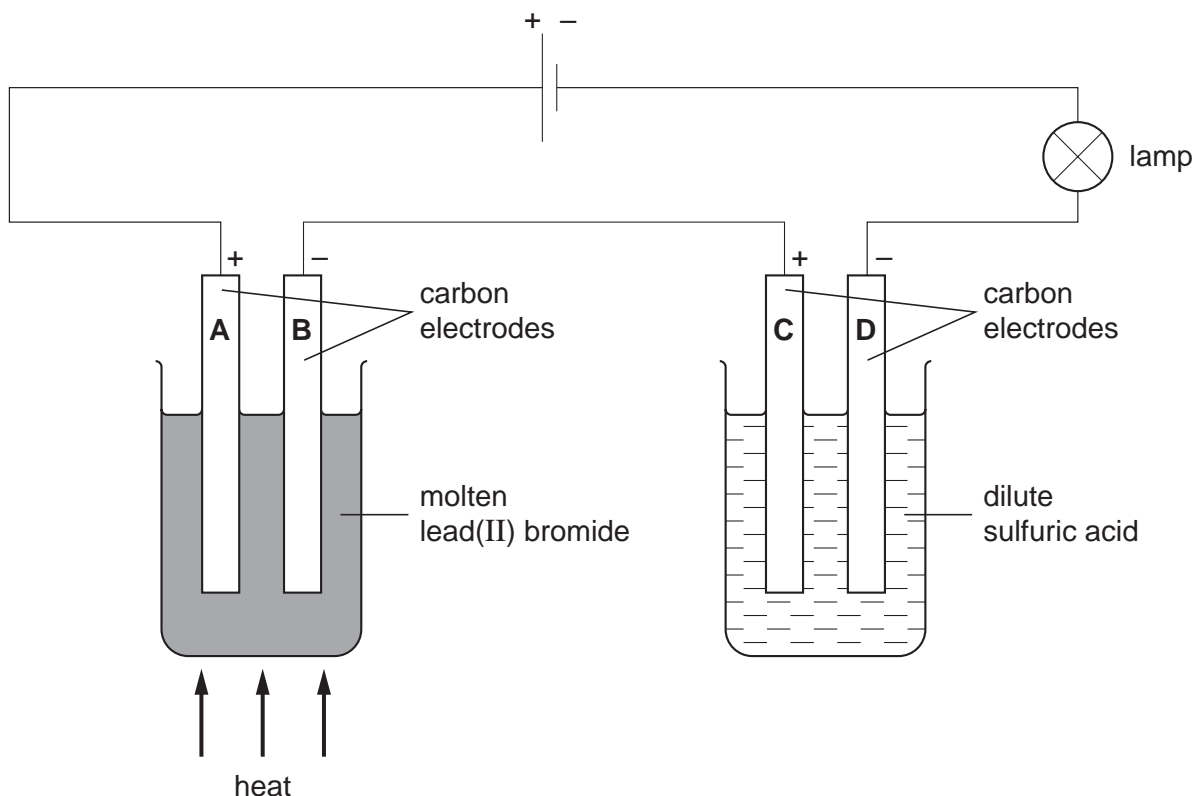
[A_r : Mg, 24; O, 16]

..... [2]

[Total: 10]

- 2 A student electrolysed molten lead(II) bromide and dilute sulfuric acid in the apparatus shown below.

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- (a) What was produced at the electrodes **A**, **B**, **C** and **D**?

A

B

C

D [4]

- (b) What, if anything, happened to the brightness of the lamp when the following changes were made to the experiment?

- (i) Water was added to the dilute sulfuric acid.

..... [1]

- (ii) The heating was stopped and the apparatus was left to cool. Explain your answer.

.....

..... [2]

[Total: 7]

In questions 3 to 7 inclusive, place a tick (✓) in the box against the best answer.

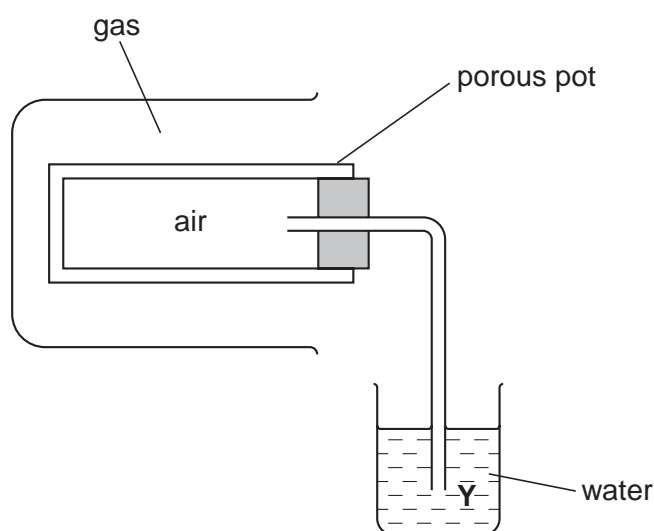
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3 A student recorded the results of four experiments with ethanol. Which result was **not** correct?

- (a) On reaction with acidified potassium dichromate(VI), the colour of the solution changed from orange to green.
- (b) The ethanol dissolved in water.
- (c) On combustion a gas was produced that popped in a flame.
- (d) On reaction with ethanoic acid, the ester ethyl ethanoate was produced.

[Total: 1]

4 The apparatus shown in the diagram was used to compare the rate of diffusion of a gas with the rate of diffusion of air.



A beaker containing the gas was placed over the porous pot.

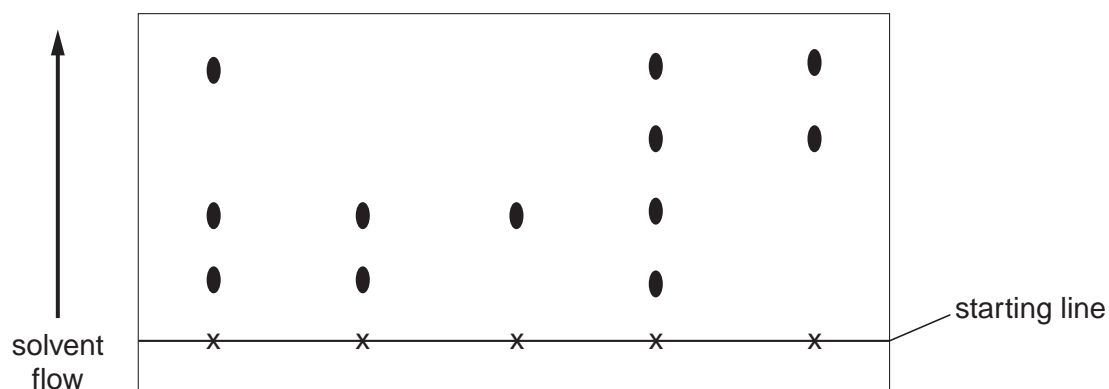
For which of the following gases did bubbles emerge at Y?

- (a) carbon dioxide [M_r : 44]
- (b) hydrogen [M_r : 2]
- (c) oxygen [M_r : 32]
- (d) sulfur dioxide [M_r : 64]

[Total: 1]

- 5 The diagram shows a chromatogram which was prepared using spots of five different inks.

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What is the total number of different dyes present in the five inks?

- (a) 3
- (b) 4
- (c) 5
- (d) 12

[Total: 1]

- 6 Some metals react with dilute hydrochloric acid to give hydrogen.

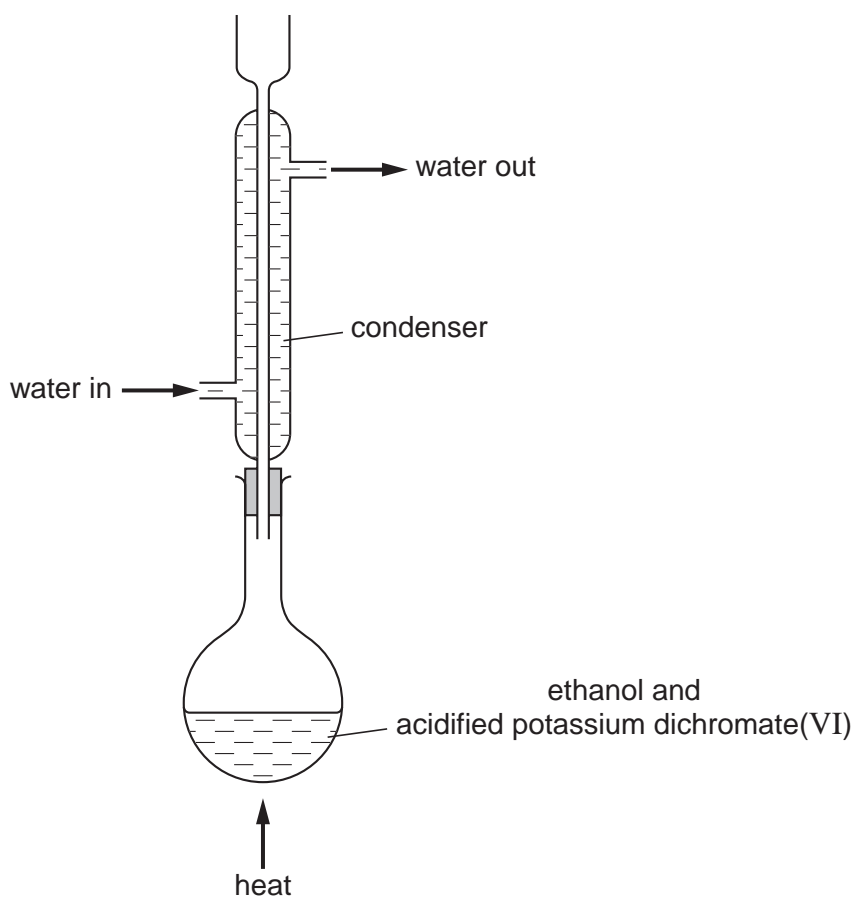
Which of the following metals will **NOT** react with dilute hydrochloric acid?

- (a) copper
- (b) iron
- (c) magnesium
- (d) zinc

[Total: 1]

7 Acidified potassium dichromate(VI) was used to oxidise ethanol to ethanoic acid.

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The purpose of the condenser was

- (a) to prevent the conversion of ethanoic acid back to ethanol.
- (b) to prevent condensation of the oxidising agent.
- (c) to prevent escape of the unreacted alcohol.
- (d) to prevent reaction between ethanol and ethanoic acid.

[Total: 1]

8 A student was given a bottle containing small pieces of scrap iron.

He was asked to find the purity of this sample of iron.

A small quantity of the iron was placed in a previously weighed container which was then reweighed.

$$\begin{array}{l} \text{mass of container + iron} = 6.16 \text{ g} \\ \text{mass of container} = 4.72 \text{ g} \end{array}$$

(a) Calculate the mass of iron.

..... g [1]

The iron was placed in a conical flask and excess dilute sulfuric acid was added to react completely with the iron. The flask was warmed and the iron reacted with the sulfuric acid to produce aqueous Fe^{2+} ions.

A gas was produced during this reaction.

(b) (i) Name this gas.

..... [1]

(ii) Give a test for this gas.

..... [1]

(iii) State the equation for the reaction between iron and sulfuric acid.

..... [1]

When all the iron had reacted the resulting solution was cooled and made up to 250 cm^3 with distilled water in a graduated flask. This was solution **S**.

Using a pipette, 25.0 cm^3 of **S** was transferred into a conical flask.

Solution **T** was 0.0200 mol/dm^3 potassium manganate(VII).

A burette was filled with **T**.

T was run into the conical flask until an end-point was reached.

Aqueous potassium manganate(VII) is purple.

Aqueous Fe^{2+} is pale green.

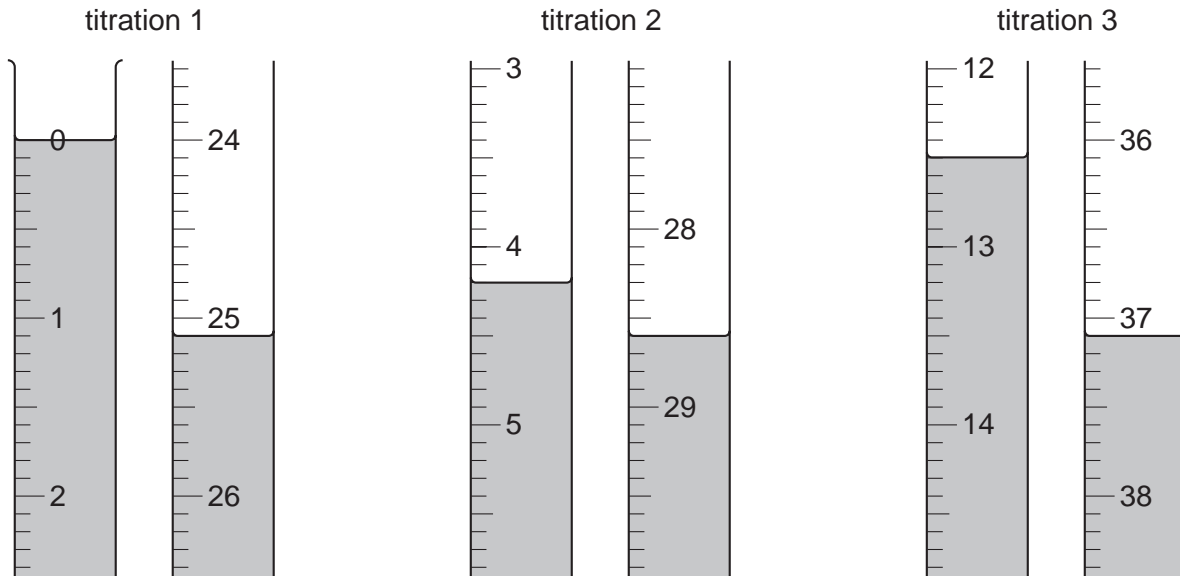
(c) What colour is the solution in the flask at the end-point?

Explain your answer.

.....
 [2]

Three titrations were done. The diagrams below show parts of the burette with the liquid levels at the beginning and end of each titration.

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(d) Use the diagrams to complete the results table.

titration number	1	2	3
final burette reading/cm ³			
initial burette reading/cm ³			
volume of T used/cm ³			
best titration results (✓)			

Summary

Tick (✓) the best titration results.

Using these results, the average volume of T was cm³.

[4]

- (e) Calculate the number of moles of potassium manganate(VII) in the average volume of **T** required in (d).

T was 0.0200 mol/dm^3 potassium manganate(VII).

..... moles [1]

- (f) Five moles of Fe^{2+} react with one mole of potassium manganate(VII).

Calculate the number of moles of Fe^{2+} in 25.0 cm^3 of **S**.

..... moles [1]

- (g) Calculate the number of moles of Fe^{2+} in 250 cm^3 of **S**.

..... moles [1]

- (h) Calculate the mass of Fe^{2+} in 250 cm^3 of **S**.

[A_r : Fe, 56]

..... g [1]

- (i) Using your answers to (a) and (h), calculate the percentage purity of the sample of iron.

..... % [1]

[Total: 15]

- 9 The following table shows the tests a student did on compound **L** and the conclusions made from the observations.

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Complete the table by describing these observations including the name of any gas evolved.

test	observation	conclusion
<p>(a) (i) L was dissolved in hydrochloric acid.</p> <p>(ii) The gas which was given off was tested.</p>		<p>L may contain the CO_3^{2-} ion.</p> <p>The presence of the CO_3^{2-} ion in L is confirmed.</p>
<p>(b) The colour of the solution from (a)(i) was noted. This solution was divided into two parts for tests (c) and (d).</p>		<p>L is a compound of a transition metal.</p>
<p>(c) (i) To the first part aqueous sodium hydroxide was added until a change was seen.</p> <p>(ii) An excess of aqueous sodium hydroxide was added to the mixture from (i).</p>		<p>L may contain Cu^{2+} ions.</p>
<p>(d) (i) To the second part aqueous ammonia was added until a change was seen.</p> <p>(ii) An excess of aqueous ammonia was added to the mixture from (i).</p>		<p>The presence of Cu^{2+} ions in L is confirmed.</p>

Conclusion: the formula of compound **L** is

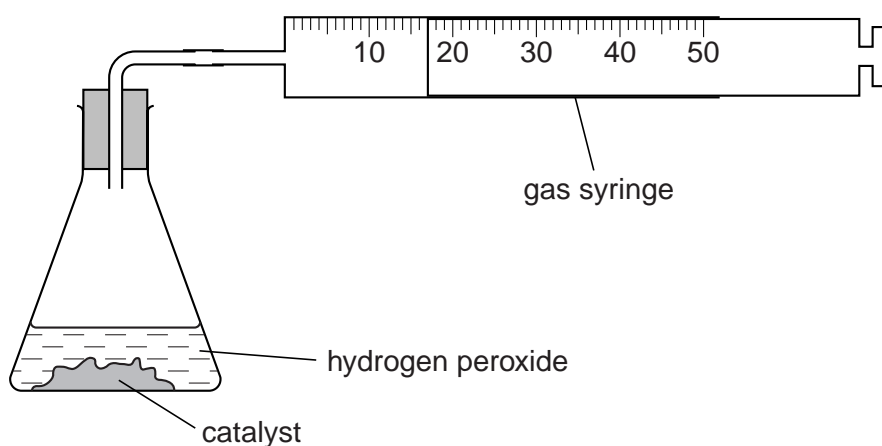
[Total: 9]

- 10 Hydrogen peroxide, H_2O_2 , decomposes slowly at room temperature to form water and oxygen.



A student investigated how the rate of decomposition changed by using two catalysts, manganese(IV) oxide and copper.

The volume of oxygen produced was measured at intervals using the apparatus shown below.



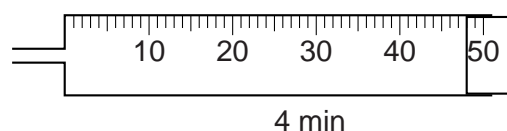
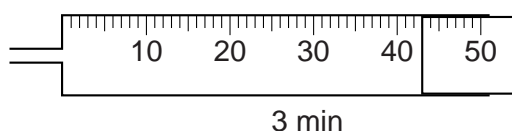
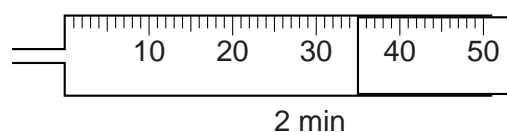
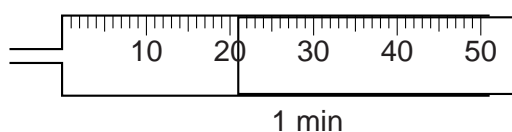
The student carried out two experiments using the same volume of hydrogen peroxide but with the same mass of a different catalyst in each experiment.

Experiment 1 uses manganese(IV) oxide as the catalyst.

Experiment 2 uses copper as the catalyst.

- (a) The results for experiment 1 and some of the results for experiment 2 are shown in the table.

Use the diagrams to complete the results for experiment 2.



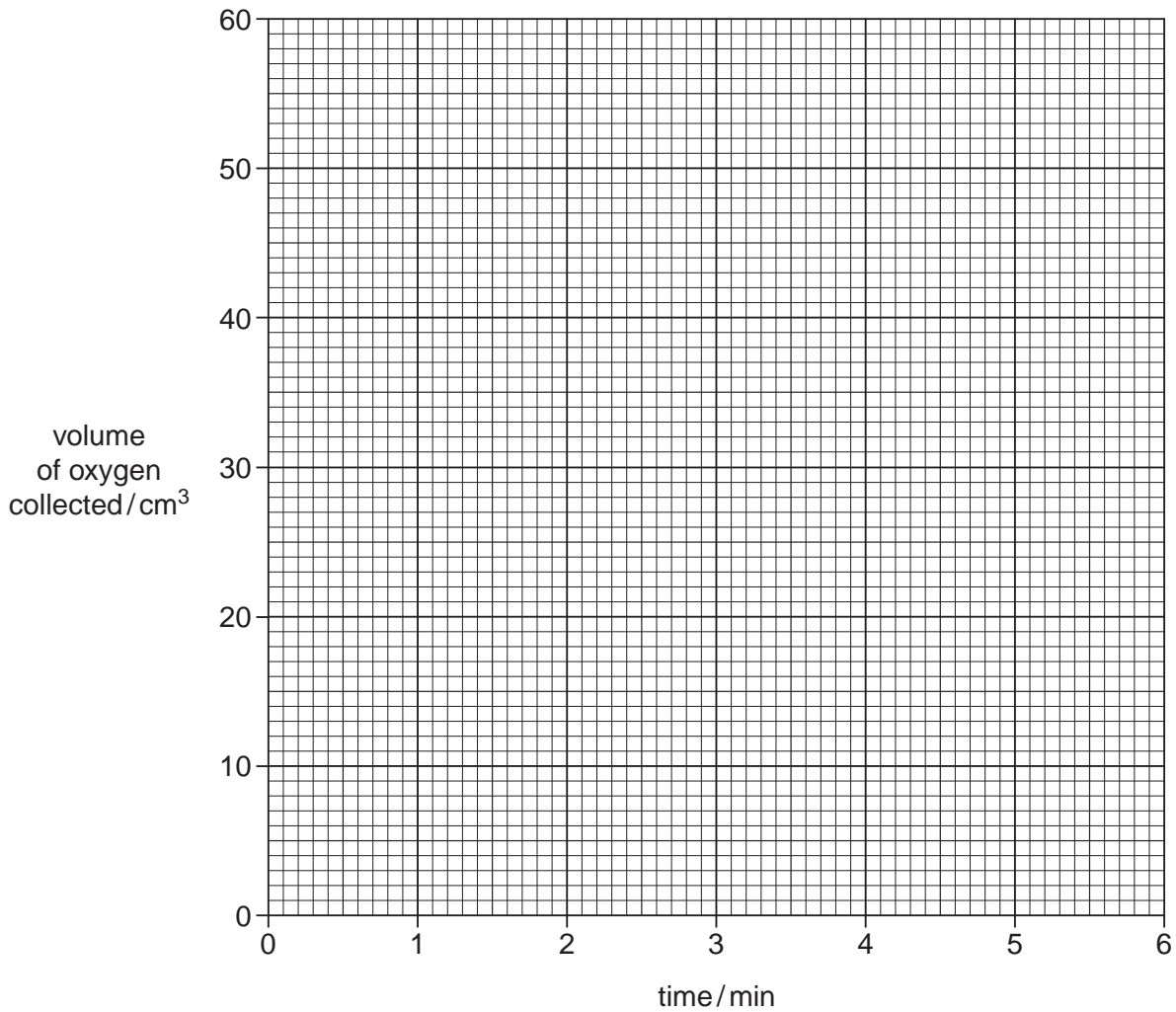
time/min	1	2	3	4	5	6
volume of oxygen collected in experiment 1/cm ³	9	17	24	29	32	35
volume of oxygen collected in experiment 2/cm ³					50	50

[2]

(b) Plot the results from experiments 1 and 2 on the grid below and draw a smooth curve through each set of points.

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Label the curves 1 and 2.



[3]

(c) Which of the two experiments first reached completion?

Explain your answer.

..... [1]

- (d) Use your graph to estimate the time taken in experiment 1 (using manganese(IV) oxide) to double the volume of oxygen produced from 15 cm³ to 30 cm³. Record your answers in the table below.

	experiment 1
time taken to produce 30 cm ³ /min	
time taken to produce 15 cm ³ /min	
time taken to double the volume from 15 cm ³ to 30 cm ³ /min	

[2]

The rate of a reaction may be calculated using the formula:

$$\text{rate of reaction} = \frac{\text{volume of oxygen produced/cm}^3}{\text{time taken/min}}$$

- (e) Using the two graphs and the above formula calculate the rate of each reaction after the first 2.5 minutes.

- rate of reaction using manganese(IV) oxide (experiment 1)

..... cm³/min

- rate of reaction using copper (experiment 2)

..... cm³/min

[3]

- (f) Using your answers to (e) suggest which is the better catalyst, manganese(IV) oxide or copper?

Explain your answer.

..... [1]

- (g) At the end of experiment 2 the copper was removed from the solution by filtration. It was dried and weighed. How does this mass of copper compare with the mass of copper used at the start of the experiment?

Explain your answer.

..... [1]

- (h) Suggest how the rate of decomposition in either experiment could be further increased.

..... [1]

[Total: 14]

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