# Chemistry Paper 3 Notes

Grass

July 2022

# Intro

By Grass and me. Creating the hyperef'd table of contents is pain.

# Contents

# General Experiment notes

1. Always read the full instructions for the whole experiment before starting
2. Use appropriate apparatus $\Rightarrow$ Transfer with measuring cylinder if they ask for a certain volume of a solution (NEVER use beaker for measurement)
3. Reading the volume of liquid: If meniscus curved upwards, read the volume frm the bottom of the meniscus. If it curves downwards, read it frm the top of the meniscus.

# Titration

## 3.1 Tables

Normal Titration Table:

| Titration Number | 1 | 2 | 3 |
|---|---|---|---|
| Final Burette Reading / $cm^3$ | | | |
| Initial Burette Reading / $cm^3$ | | | |
| Volume of R / $cm^3$ | | | |
| Best Titration Results ( ✓) | | | |

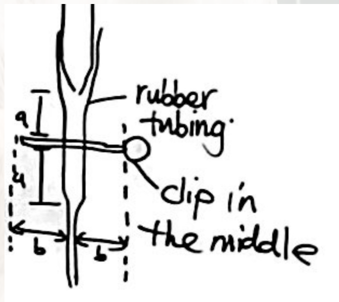
Thermometric Titration Table

| Final Burette Reading / $cm^3$ | Initial Burette Reading / $cm^3$ | Volume of R / $cm^3$ | Highest Temperature / °$C$ | Total Temperature Change / °$C$ |
|---|---|---|---|---|
| 0.00 | 0.00 | 0.00 | 30.0 | 0.0 |
| 5.00 | 0.00 | 5.00 | 32.5 | +2.5 |
| 10.00 | 0.00 | 10.00 | 35.0 | +5.0 |
| ⋮ | ⋮ | ⋮ | ⋮ | ⋮ |
| $F_n$ | 0.00 $I_n$ | $F_n$ : 0.00 | $H_n$ | $H_n$ : +30.0 |

**Burette**

**Note / AFI:**

1. Wash the filter funnel with the solution to be placed in the burette
2. Place the filter funnel on the mouth of the burette when filling it up
3. Remove the filter funnel after, and for all burette readings
4. Make sure the rubber tubing is filled with the solution. i.e. Ensure the tip of the burette contains no air
5. Middle of burette clip on the middle of the rubber tubing




6. Burette Readings in 2 d.p.!
7. Burette must be staightened
8. Can use formula to calculate concentration of acid A o/r base B:
   Let 1 be acid A, 2 be base B

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$
$$M_2 = \frac{M_1V_1n_2}{V_2n_1}$$

   Where $M_i$ is the molar concentration of $i$, $V_i$ the volume of $i$, and $n_i$ the the mole ratio (in a reaction between the two substances)
9. T2CA

(a) T - Table: 1m
(b) 2 - 2 d.p.: 1m
(c) C - Consistency: 1m
(d) A - Accuracy: 2m

## 3.2 Pipette

**Note / AFI:**

1. Do not purposely 'tap' remaining liquid out!
2. Holding of pipette; Hold on to the top portion when inserting the pipette filler (NOT the bulb!)

## 3.3 Burette + Pipette

1. If solution is spilled on the sides of the conical flask, use the deionsed water to wash it off to ensure maximum accuracy.
2. Read the burette and pipette readings at eye level.
3. Ensure they are no bubbles when filling up the burette and pipette

## 3.4 Accuracy of Instruments

| | |
|---|---|
| Burette | $\pm 0.05$ ($cm^3$) |
| Pipette | 20.00 / 25.00 $cm^3$ |
| Thermometer | $\pm 0.5$ (°C) |
| Electronic Balance | 2 d.p. (g) |
| Measuring Cylinder | $\pm 0.5$ ($cm^3$) |
| Stopwatch | Nearest $s$ |
| Gas Syringe | Nearest $cm^3$ |

# Dessert Qns

1. A student repeated the experiment and accidentally used a $20.0\text{cm}^3$ pipette to measure solution $Q$ into the flask in each titration.
   The student thought he had used a $25.0\text{cm}^3$ pipette
   Describe and explain the effect that this would have on the percentage by mass of iodine in the solution calculated by the student.

   *Ans* : Using $20.0\text{cm}^3$ of $Q$ instead of $25.0\text{cm}^3$ will result in lower volume of $R$ needed [1]. This will result in a lower concentration of $I_2$ calculated in $Q$ and hence a lower percentage by mass of $I_2$ in $Q$. [1]

2. Explain why it is necessary to stir the mixture in the beaker before measuring the highest temperature reached by the mixture.

   *Ans* : It is to ensure that all the metal carbonate $X$ has reacted completely with the respective acids. [1]

3. State one **key** source of error in the results obtained in Experiments 1 to 3 (temperature experiment). Suggest **one** improvement you could make to the experiment to reduce this error.[1]

   *Ans* : Beaker can be insulated with poor heat conductor / cover with lid nested in beaker [2]

[1] Experimental Error, NOT human error

# Temperature Related Experiments

1. Change in temperature / enthalpy must have "+" and "-" signs

# Q.A. Test

## 6.1 Note / AFI:

1. Always rmb to use test tube or Q.A. test – The question might not explicitly state this
2. Hold boiling tube at an angle and reduce the flame height
3. Open the bunsen burner air hole halfway if you have problems lighting it
4. Adding excess bench reagent ⇒ Pour some test sample away and continue adding the reagent (if the test tube is getting too full)
5. Collect solid samples with spatula (always use the small scoop)
6. Even if there is nothing seen/heard/smelled for a particular step, the observation should still be written, as "No visible Change".
7. When writing observations rmb to put **all** the information down. [1]
8. Some observations required are implied, but not explicitly stated. [2]

## 6.2 Best/Modal Answers for Observations:

1. Nothing happens ⇒ No visible change
2. Add aqueous ammonia slowly and stir with a glass rod until no further changes are seen ⇒ [3]A light blue precipitate was formed, which dissolved in excess to form a deep blue solution

[1]E.g.: What is the color and smell of the gas; write colorless and odorless if nothing specific seen or smelled

[2]E.g.: Place the boiling tube in the test-tube rack to allow its contents to settle.
⇒ After waiting, rmb to add in the observation for this; like "Blue precipitate settled at the bottom of the test tube.

[3]No need write in the front like: "Upon adding aqueous ammonia / silver nitrate / whatever reagents"

# Experimental Planning

## 7.1 Things to have

(in order which they should be presented)

1. Diagram
2. Approach: (Short summary of what the experiment is going to be)[1]
3. Steps: 1.
4. 2.
5. $\vdots$
6. $n$.
7. Conclusion (Look at the qns to know what to write here, i.e results that can be conculded after the expt)[2]

## 7.2 Notes for Steps

1. Be specific. Don't leave it up to interpretation. i.e. Write down the mass / volume of reagent to be added.

---

[1] E.g.:

(a) Approach: Comparing the effectiveness of the 3 metal oxide catalysts by measuring the volume of oxygen produced after a fixed time

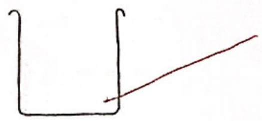
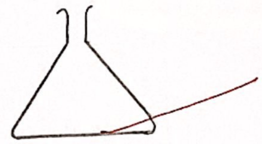
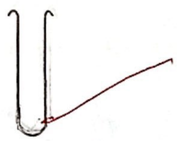
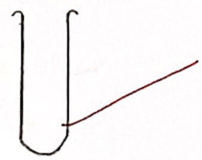
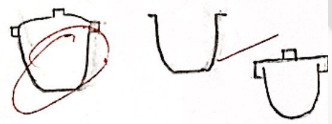
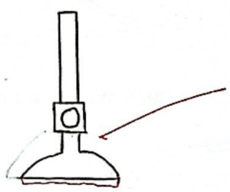
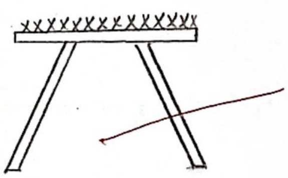
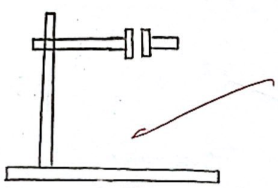
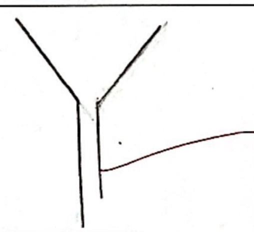
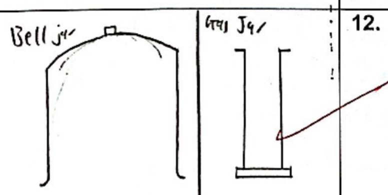
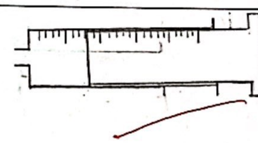
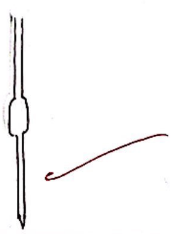
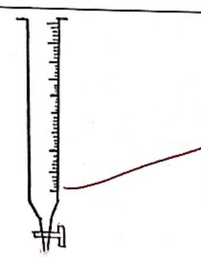
(b) Approach: Measure mass of contents before and after heating using electronic balance

[2] E.g.:

(a) The metal oxide that takes the least amount of time to produce $50cm^3$ of oxygen gas is the most effective catalyst, while the one that takes the longest time is the least effective catalyst.

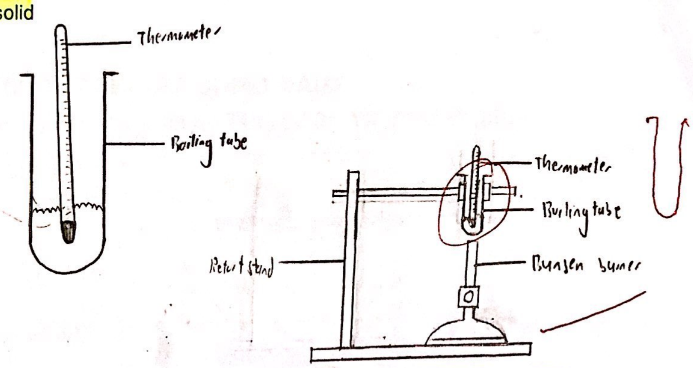
(b) Use the formula $\frac{\text{mass before heating}-\text{mass after heating}}{\text{mass of hydrated crystals}} \times 100\%$ to find the percentage mass of water in copper (II) sulfate crystals

## 7.3 Apparatus Drawing

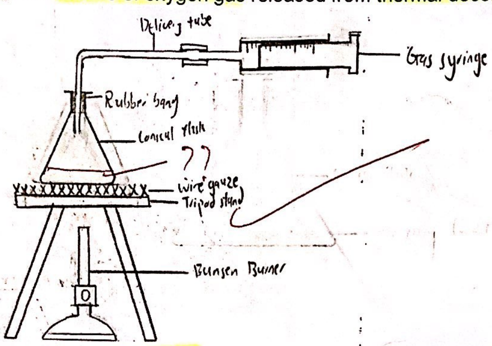
| | | | | | |
|---|---|---|---|---|---|
| 1. | Beaker |  | 2. | Conical flask |  |
| 3. | Test tube |  | 4. | Boiling tube |  |
| 5. | Evaporating dish |  | 6. | Crucible |  |
| 7. | Bunsen burner |  | 8. | Tripod stand (with wire gauze) |  |
| 9. | Retort stand |  | 10. | Filter funnel |  |
| 11. | Gas jar |  | 12. | Gas syringe |  |
| 13. | Pipette |  | 14. | Burette |  |

**15.** Draw an experimental setup for the following investigations.

**(a)** Measure the melting point of a solid




**(b)** Determine the volume of oxygen gas released from thermal decomposition of mercury(II) oxide.




**(c)** Determine the speed of reaction by measuring the decrease in mass of a reaction mixture.

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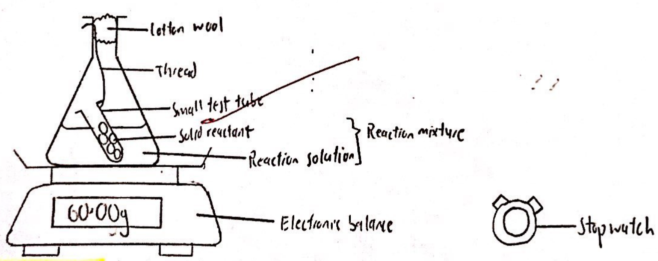
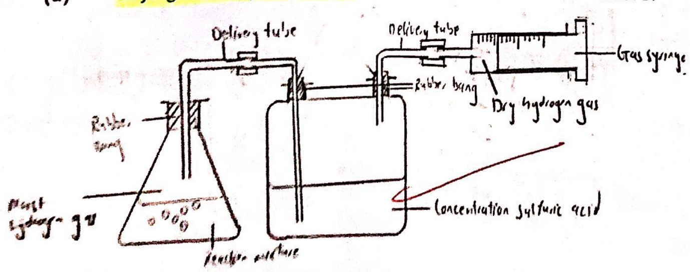



**(d)** Drying moist hydrogen gas produced from a reaction mixture.



# Examples

## 8.1 Titration

### 8.1.1 Timed Assignment 2022: Paper 3 Practical

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2 A dibasic acid, $H_2X$ is commonly used as a reducing agent and in disinfectants.

No. 4

You are going to determine the volume of the dibasic acid required to neutralise a fixed volume of sodium hydroxide solution and hence, deduce the molecular formula of the acid.

**Read all the instructions below carefully before starting the experiment in Question 2.**

**Instructions**

You are going to carry out a titration experiment.

**P** is 0.0940 mol/dm$^3$ aqueous sodium hydroxide.
**Q** is aqueous solution of a dibasic acid, $H_2X$.

Burette Q

**(a)** Put **Q** into the burette.

Pipette P (25 cm$^3$)

Pipette 25.0 cm$^3$ of **P** into a conical flask.

Titrate **P** with **Q** using the indicator, methyl orange. Stop the titration when the indicator turns orange. This is the end-point of the titration.

Record your titration results in the space provided, repeating the titration as many times as you consider necessary to achieve consistent results.

**Results**

2 d.p.!!!

| Titration Number | 1 | 2 | 3 |
|---|---|---|---|
| Final burette reading / cm$^3$ | 24.80 | 24.90 | 24.90 |
| Initial burette reading / cm$^3$ | 0.00 | 0.00 | 0.00 |
| Volume of Q / cm$^3$ | 24.80 | 24.90 | 24.90 |
| Best titration results (✓) | | ✓ | ✓ |

Mr Leo's average: 28.30

① Methyl orange indicator → Add just 1 or 2 drops

② QA Test → Used beaker instead of test tube
NEED to use test tube

T – 1
2 – 1
C – 1
A – 0

③ Holding of pipette →
Hold on to the top portion when inserting the pipette filler (NOT the bulb!)

3

[5]

④ Read instructions in advance
(whole set of instructions)

⑤ Use appropriate apparatus → Transfer with measuring cylinder if they ask for certain volume of a solution.
(NEVER use beaker for measurement)

**(b)** From your titration results, obtain an average volume of **Q** to be used in your calculations. Show clearly how you obtained this volume.

☆ MUST show working!

$\frac{24.40 + 24.90}{2} = 24.90\ cm^3$

average volume of **Q** ....24.90 cm³.... [1]

**(c)** 1 mole of the acid $H_2X$ reacts with 2 moles of sodium hydroxide. Using your results, calculate the concentration, in mol/dm³, of $H_2X$ in **Q**.

~~Volume~~ mol of P (NaOH)

$= 0.0940 \times \frac{25}{1000}$

$= 0.00235$ mol

molar concentration, volume

no. of mol of 1: $M_1V_1$

$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$

mole ratio

$M_2 = \frac{M_1 V_1 n_2}{V_2 n_1}$

NaOH : $H_2X$

2 : 1

0.00235 mol → 0.001175 mol

~~tan~~ conc. of $H_2X$ in Q

$= \frac{0.001175}{\left(\frac{24.90}{1000}\right)}$

$= \frac{47}{1196}$ mol/dm³

$= 0.0393$ mol/dm³ (3 s.f.)

concentration of $H_2X$ in **Q** ....0.0393.... mol/dm³ [2]

ecf

**(d)** Assuming that the concentration, in g/dm³, of $H_2X$ is 3.53 g/dm³, use your answer in **(c)** to calculate the relative molecular mass of $H_2X$.

$M_r$ of $H_2X$

$= \frac{3.53}{\left(\frac{47}{1196}\right)}$

$= \cancel{89.8}$ 90 (correct to nearest whole no.)

relative molecular mass of $H_2X$ ....90.... [1]

ecf

7

**(e)** The expected molecular formula of $H_2X$ is determined to be either $H_2SO_3$ or $H_2SO_4$. Based on your answer in **(d)**, identify the molecular formula of $H_2X$.

sulfurous acid

$M_r$ of $H_2SO_3$

$= 2(1) + 32 + 3(16)$

$= 82$

$M_r$ of $H_2SO_4$

$= 2(1) + 32 + 4(16)$

$= 98$

$90 - 82 = 8$

$98 - 90 = 8$

Your proving does not confirm if it's $H_2SO_3$ and not $H_2SO_4$. Lucky guess!!

molecular formula of $H_2X$ ....$H_2SO_3$.... [3]

3

**(f)** A student repeated the experiment and accidentally used a solution of sodium carbonate, $Na_2CO_3$, instead of sodium hydroxide in the titration. The solution of sodium carbonate is of the same concentration as **P**.

Would the volume of **Q** required be higher or lower than the titre values obtained in **(b)**? Explain why.

Volume of Q required would be higher.

Number of moles of Q required would be higher as 1 mole of $Na_2CO_3$ reacts with 1 mole of $H_2X$ in Q.

[2]

0

[Total: 14]

$H_2X + Na_2CO_3 \rightarrow Na_2X + H_2O +$

N : H

3

### 8.1.2 Iodine Antiseptic

Name: Yu Shao Hong ( 34 ) Class: 4-I Date: ________ Mark: ______ /23

2 Iodine is widely used as an antiseptic.

You are going to determine the concentration of iodine in an antiseptic solution by titration with aqueous sodium thiosulfate, $Na_2S_2O_3$, using starch solution as an indicator.

$$2Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + 2NaI$$

**Read all the instructions below carefully before starting the experiment in Question 2.**

**Instructions**

You are going to carry out a titration experiment.

**Q** is an antiseptic solution containing iodine.
**R** is 0.160 mol/dm³ aqueous sodium thiosulfate.

**(a) (i)** Put **R** into the burette.

Pipette 25.0 cm³ of **Q** into a flask.

Add **R** from the burette until the red-brown colour fades to pale yellow, **then** add a few drops of the starch solution. This will give a dark blue solution.

Continue adding **R** slowly from the burette until one drop of **R** causes the blue colour to disappear, leaving a colourless solution.

Record your titration results in the space provided, repeating the titration as many times as you consider necessary to achieve consistent results.

**Results**

| Titration number | 1 | 2 | 3 |
|---|---|---|---|
| Final burette reading / cm³ | 23.20 | 46.50 | 32.50 |
| Initial burette reading / cm³ | 0.00 | 24.20 | 10.00 |
| Volume of R / cm³ | 23.20 | 22.30 | 22.50 |
| Best titration results (✓) | | ✓ | ✓ |

[5]

(ii) From your titration results, obtain an average volume of **R** to be used in your calculations. Show clearly how you obtained this volume.

$\frac{22.30+22.50}{2} = 22.40$

average volume of **R** ...... 22.40 cm$^3$ ......

(b) (i) **R** is 0.160 mol/dm$^3$ aqueous sodium thiosulfate.

Using your results, calculate the concentration, in mol/dm$^3$, of iodine in **Q**.

mol of R
$= 0.160 \times \frac{22.40}{1000}$
$= 0.003584$ mol

$Na_2S_2O_3 : I_2$
$2 : 1$
0.003584 mol $\rightarrow$ 0.001792 mol

Concentration of $I_2$ in Q
$= \frac{0.001792}{\left(\frac{25}{1000}\right)}$
$= 0.07168$ mol/dm$^3$
$= 0.0717$ mol/dm$^3$ (3 s.f.)

concentration of iodine in **Q** ...... 0.0717 (3 s.f.) ...... mol

(ii) Using your answer from **(i)**, calculate the mass of iodine in 1 dm$^3$ of **Q**. [$A_r$: I, 127

mol of Q
$= 0.07168 \times 1$
$= 0.07168$ mol

mass of $I_2$ in Q
$= 0.07168 \times 2(127)$
$= 18.20672$ g
$= 18.2$ g (3 s.f.)

mass of iodine in 1 dm$^3$ of **Q** ...... 18.2 g (3 s.f.) ......

(iii) Given that 1 cm$^3$ of **Q** has a mass of 1.00 g, calculate the percentage by mass of the solution.

~~Conc. of Q in g/dm$^3$~~
~~$= \frac{1}{\left(\frac{1}{1000}\right)}$~~
~~$= 1000$ g/dm$^3$~~

mass of 1 dm$^3$ of Q
$= 1 \times 1000$
$= 1000$ g

% by mass of $I_2$ in Q
$= \frac{18.20672}{1000} \times 100\%$
$= 1.820672\%$
$= 1.82\%$ (3 s.f.)

percentage by mass of iodine in **Q** ...... 1.82 % (3 s.f.) ......

**(c)** A student repeated the experiment and accidently used a 20.0 $cm^3$ pipette to measure **solution Q** into the flask in each titration.

The student thought he had used a 25.0 $cm^3$ pipette.

Describe and explain the effect this would have on the percentage by mass of iodine in the solution calculated by the student.

The student would calculate a lower concentration of iodine in Q as the mol of R, and hence Q, would decrease proportionately when using a 20.0 $cm^3$ pipette instead, but the student will calculate it using a higher volume of 25.0 $cm^3$ instead of the correct volume of 20.0 $cm^3$. So the ~~percentage by~~ mass of iodine calculated, and hence the percentage by mass of iodine, which is $\frac{\text{mass of iodine}}{\text{mass of sample}} \times 100\%$ will decrease. [2]

[Total: 14

B: Using 20.0 $cm^3$ of Q instead of 25.0 $cm^3$ will result in lower volume of R needed [1]
This will result in lower concentration of $I_2$ calculated in Q and hence a lower percentage by mass of $I_2$ in Q. [1]

### 8.1.3 Experiment on Volumetric Analysis 4 April 2022

Redox Titration

Aqueous hydrogen peroxide, $H_2O_2$, can be prepared in the laboratory by adding dilute sulfuric acid to barium peroxide, $BaO_2$, and filtering the resulting mixture to remove insoluble barium sulfate.

$$BaO_2\,(s) + H_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + H_2O_2\,(aq)$$

conc. of P ?

P: Contains excess $H_2SO_4\,(aq)$

V of P = 10.0 $dm^3$

**P** is an acidified solution of hydrogen peroxide prepared in this way. You are to determine its concentration and the mass of barium peroxide required to prepare 10.0 $dm^3$ of **P**.

The concentration of hydrogen peroxide can be determined by titrating it with potassium manganate(VII). No indicator is necessary since the products of the reaction are almost colourless and one drop of potassium manganate(VII) in excess produces an easily seen pale pink colour.

$M_{Barium\ Peroxide}$ ?

**Q** is 0.0200 mol/$dm^3$ potassium manganate(VII), $KMnO_4$.

(Pip) P => Acidified $H_2O_2$ ( conc. ? )

(Bur) Q => 0.02 mol/$dm^3$ $KMnO_4$ at 25.0 $cm^3$

**(a)** Fill the burette with solution **Q**. You may find it easier to read the top of the meniscus as the colour of **Q** is so intense.

Pipette a 25.0 $cm^3$ (or 20.0$cm^3$) portion of **P** into a flask and titrate with **Q**. At first, the purple colour disappears rapidly. As the titration proceeds, this disappearance is less rapid. At the end-point, one drop of **Q** produces a pink colour that does not disappear on swirling.

Colourless => Purple pink

Record your titration results in the space provided, repeating the titration as many times as you consider necessary to achieve consistent results.

**Results**

| Titration No. | 1 | 2 | 3 |
|---|---|---|---|
| Initial burette reading / $cm^3$ | 0.00 | 11.20 | 16.80 |
| Final burette reading / $cm^3$ | 26.40 | 37.20 | 43.00 |
| Volume of Q / $cm^3$ | 26.40 | 26.00 | 26.10 26.20 |
| Best titration results (✓) | | ✓ | ✓ |

T – 0

Z – 1

C – 1

A – 0

**(b)** From your titration results, obtain an average volume of **Q** to be used in your calculations.

Show clearly how you obtained this volume.

$\frac{26.00 + 26.10}{2} = 26.05 \text{ cm}^3$

ecf.

average volume of **Q** ....26.05 cm³.........

**(c)** **Q** is 0.0200 mol/dm³ potassium manganate(VII).

**Five** moles of hydrogen peroxide react with **two** moles of potassium manganate(VII)

Using your results from **(b)**, calculate the concentration, in mol/dm³, of the hydrogen peroxide in **P**.

mol of Q

$= \frac{26.05}{1000} \times 0.0200$

$= 0.000521 \text{ mol}$

$H_2O_2 : KMnO_4$

$5 : 2$

$0.0013025 \text{ mol} \leftarrow 0.000521 \text{ mol}$

Conc. of $H_2O_2$ in P

$= \frac{0.0013025}{\left(\frac{25}{1000}\right)}$

$= 0.0521 \text{ mol/dm}^3$

$1 = H_2O_2,\ 2 = KMnO_4$

$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$

$\frac{M_1 \times 25}{5} = \frac{0.02 \times 27.50}{2}$

$M_1 = 0.0521 \text{ mol/dm}^3$

2

Concentration of hydrogen peroxide in **P** is ....0.0521........ mol/dm³. [2]

**(d)** Using your answer in **(c)**, calculate the mass of barium peroxide needed to prepare 10.0 dm³ of solution **P**.
[$A_r$: O, 16; Ba, 137]

mol of $H_2O_2$ in 10.0 dm³ of P

$= 0.0521 \times 10.0$

$= 0.521 \text{ mol}$

$BaO_2 : H_2O_2$

$1 : 1$

$0.521 \text{ mol} \leftarrow 0.521 \text{ mol}$

Mass of $BaO_2$

$= 0.521 \times [137 + 2(16)]$

$= 88.049 \text{ g}$

(88.0g)

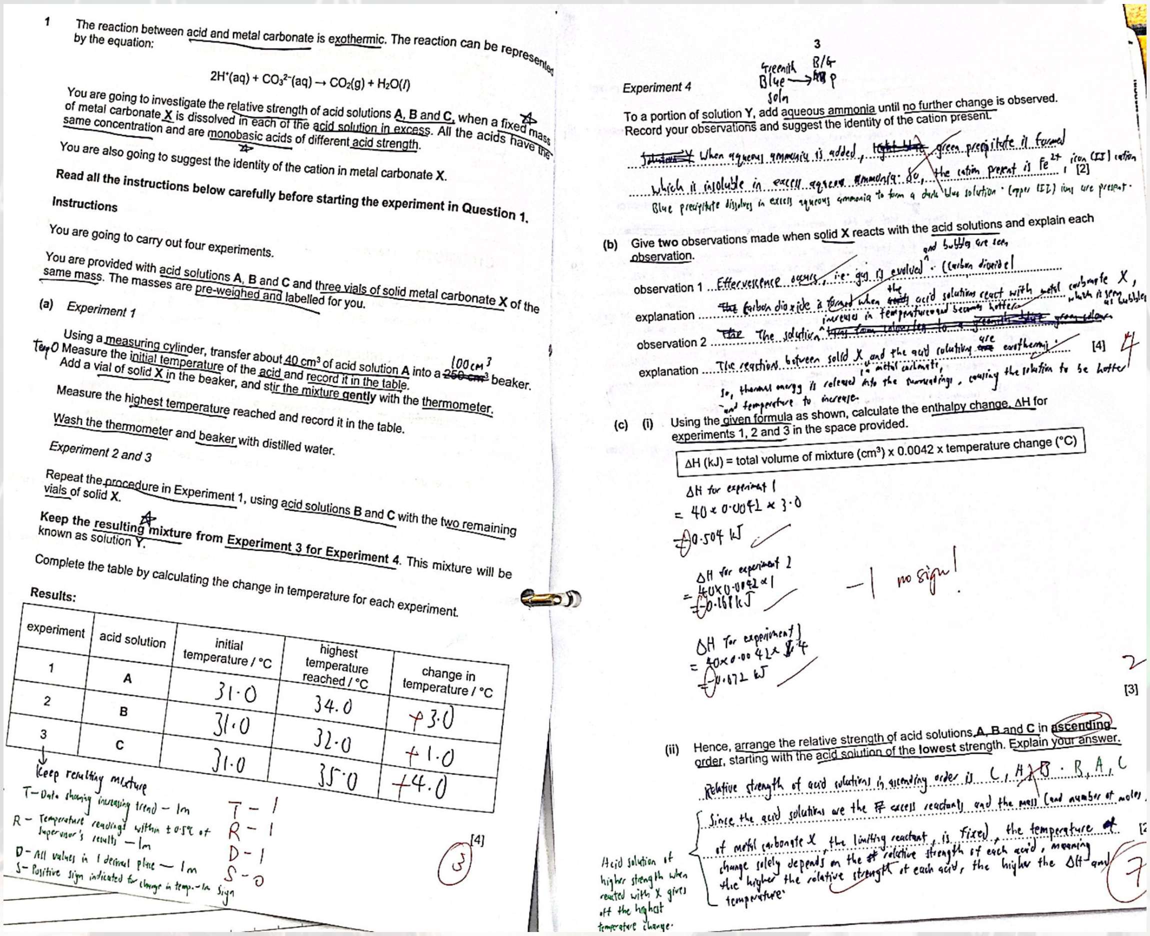
Mass of barium peroxide required is ....88.0........ g. [2]

1

3sf.

## 8.2 Short Experiment

### 8.2.1 Timed Assignment 2022: Paper 3 Practical



1 The reaction between acid and metal carbonate is exothermic. The reaction can be represented by the equation:

$$2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$$

You are going to investigate the relative strength of acid solutions **A**, **B** and **C**, when a fixed mass of metal carbonate **X** is dissolved in each of the acid solution in excess. All the acids have the same concentration and are monobasic acids of different acid strength.

You are also going to suggest the identity of the cation in metal carbonate **X**.

**Read all the instructions below carefully before starting the experiment in Question 1.**

**Instructions**

You are going to carry out four experiments.

You are provided with acid solutions **A**, **B** and **C** and three vials of solid metal carbonate **X** of the same mass. The masses are pre-weighed and labelled for you.

(a) *Experiment 1*

Using a measuring cylinder, transfer about 40 $cm^3$ of acid solution **A** into a ~~250 $cm^3$~~ 100 $cm^3$ beaker.
Temp 0 Measure the initial temperature of the acid and record it in the table.
Add a vial of solid **X** in the beaker, and stir the mixture **gently** with the thermometer.

Measure the highest temperature reached and record it in the table.

Wash the thermometer and beaker with distilled water.

*Experiment 2 and 3*

Repeat the procedure in Experiment 1, using acid solutions **B** and **C** with the two remaining vials of solid **X**.

**Keep the resulting mixture from Experiment 3 for Experiment 4.** This mixture will be known as solution **Y**.

Complete the table by calculating the change in temperature for each experiment.

**Results:**

| experiment | acid solution | initial temperature / °C | highest temperature reached / °C | change in temperature / °C |
|---|---|---|---|---|
| 1 | A | 31.0 | 34.0 | +3.0 |
| 2 | B | 31.0 | 32.0 | +1.0 |
| 3 | C | 31.0 | 35.0 | +4.0 |

Keep resulting mixture

T – Data showing increasing trend – 1m
R – Temperature readings within ±0.5°C of supervisor's results – 1m
D – All values in 1 decimal place – 1m
S – Positive sign indicated for change in temp. – 1m

T – 1
R – 1
D – 1
S – 0 Sign

[4] 3

3

*Experiment 4*

Greenish Blue soln → B/G ppt

To a portion of solution **Y**, add aqueous ammonia until no further change is observed.
Record your observations and suggest the identity of the cation present.

~~Solution Y~~ When aqueous ammonia is added, ~~light blue~~ green precipitate is formed which is insoluble in excess aqueous ammonia. So, the cation present is $Fe^{2+}$, iron (II) cation. [2]

Blue precipitate dissolves in excess aqueous ammonia to form a dark blue solution. Copper (II) ions are present.

(b) Give **two** observations made when solid **X** reacts with the acid solutions and explain each observation.

observation 1 Effervescence occurs, i.e. gas is evolved and bubbles are seen. (carbon dioxide)

explanation ~~The~~ Carbon dioxide is formed when the ~~acids~~ acid solution react with metal carbonate X, which is seen as bubbles.

observation 2 ~~The~~ The solution increases in temperature and becomes hotter. ~~turns from colourless to a greenish blue green colour~~

explanation The reactions between solid X and the acid solutions ~~are~~ are exothermic in metal carbonate, so, thermal energy is released into the surroundings, causing the solution to be hotter and temperature to increase. [4] 4

(c) (i) Using the given formula as shown, calculate the enthalpy change, ΔH for experiments 1, 2 and 3 in the space provided.

$$\boxed{\Delta H\ (kJ) = \text{total volume of mixture } (cm^3) \times 0.0042 \times \text{temperature change } (^\circ C)}$$

ΔH for experiment 1
$= 40 \times 0.0042 \times 3.0$
$= -0.504$ kJ

ΔH for experiment 2
$= 40 \times 0.0042 \times 1$
$= -0.168$ kJ

–1 no sign!

ΔH for experiment 3
$= 40 \times 0.0042 \times 4$
$= -0.672$ kJ

2

[3]

(ii) Hence, arrange the relative strength of acid solutions **A**, **B** and **C** in ascending order, starting with the acid solution of the **lowest** strength. Explain your answer.

Relative strength of acid solutions in ascending order is ~~C, A, B~~. B, A, C

Since the acid solutions are the ~~#~~ excess reactants and the mass (and number of moles) of metal carbonate X, the limiting reactant, is fixed, the temperature ~~of~~ change solely depends on the ~~#~~ relative strength of each acid, meaning the higher the relative strength of each acid, the higher the ΔH and temperature. [2]

Acid solution of higher strength when reacted with X gives off the highest temperature change.

7

**(d)** Explain why it is necessary to stir the mixture in the beaker before measuring the highest temperature reached by the mixture.

It is to ensure that all the metal carbonate X has reacted with the respective acids [1]

(Also increases speed of reaction between X and the acids)

**(e)** State **one** key source of error in the results obtained in Experiments 1 to 3.
Suggest **one** improvement you could make to the experiment to reduce this error.

error: thermal energy produced exothermically is lost to the surroundings easily ~~as the top of the beaker is~~

=> (Expt error, not human error)

improvement: Cover the beaker quickly after the reaction has completed and or wrap the beaker with some insulating materials, like a cloth. [2]

~~To ensure that both acid~~

Beaker can be insulated with poor heat conductor / cover with lid

use styrofoam cup nested in beaker

**(f)** If the same experiments were repeated using an excess of metal oxide instead of metal carbonate, the temperature change would be positive. Suggest why this is so.

metal oxides are basic oxides. So, when they react with the acid solutions, they neutralise each other, ~~met~~ where $\Delta H < 0$, releasing ~~giving~~ thermal energy into the surroundings, thereby creating a positive temperature change. [2]

2

$2HCl + CaO \longrightarrow CaCl_2 + H_2O$ [Total: 20]

(4)

### 8.2.2 Combined Chemistry Q.A. Test 2020

8

**2** You are provided with:

- two samples of solid **P**
- a sample of solution **Q**
- two samples of solution **R**, labelled $\mathbf{R_1}$ and $\mathbf{R_2}$
- a sample of solution **I** (screened methyl orange indicator).

**Q** is a salt and solutions $\mathbf{R_1}$ and $\mathbf{R_2}$ have different concentrations.

Carry out the following tests. You should test and identify any gases evolved. Carefully record your observations.

The volumes given below are approximate and should be estimated rather than measured unless instructed otherwise.

| | test | observations |
|---|---|---|
| **(a)** | Add about 5cm depth of limewater into a clean test-tube.<br><br>Remove the stopper from one of the boiling tubes containing solid **P**.<br><br>Heat this boiling tube containing solid **P** gently for about one minute.<br><br>While you are heating the boiling tube containing solid **P**, withdraw several samples of gas from the boiling tube using a dropper pipette.<br><br>Do **not** allow the dropper pipette to touch the boiling tube.<br><br>Bubble each sample of gas into the limewater.<br>P: $CO_3^{2-}$ ✓ | The samples of gas formed a white precipitate with the limewater.<br><br>B: No visible change (during heating)<br><br>Colourless and odourless gas evolved, produced a white precipitate in limewater.<br><br>[2] |
| **(b)** | Add about 1cm depth of solution **Q** into the other boiling tube containing solid **P**. Shake the mixture.<br><br>Place the boiling tube in the test-tube rack to allow **its contents to settle.**<br><br>P: $Cu^{2+}$ ✓ | When Q was added to P, a light blue precipitate was formed.<br><br>B: Blue precipitate produced<br>Bubbling occur<br>Colourless and odourless gas produce a white precipitate when bubbled through limewater.<br><br>Blue precipitate settled at bottom of test-tube<br><br>[2] |
| | **While you are waiting for the contents of the boiling tube to settle, continue with (c).** | |

For Examiner's Use

Smell
No $CO_2$ $O_2/H_2$
Yes NH OR Cl OR SO

* 0012511521209 *

9

| | test | observations |
|---|---|---|
| (c) | Add about 1cm depth of solution **Q** into a clean test-tube.<br><br>To this test-tube, add about 1cm depth of dilute nitric acid.<br><br>To the same test-tube, add about 1cm depth of barium nitrate solution.<br><br>Q : $SO_4^{2-}$ ✗ | When barium nitrate was added, the solution remained blue. ✓<br><br>OR: No visible reaction<br><br>[1] |
| (d) | Add about 1cm depth of solution **Q** into a clean test-tube.<br><br>To this test-tube add about 1cm depth of dilute nitric acid.<br><br>To the same test-tube, add about 1cm depth of silver nitrate solution.<br><br>Q : $Cl^-$ ✓ | ~~When silver nitrate was added~~, a white precipitate was formed. ✓ |
| | Then add aqueous ammonia slowly and stir with a glass rod until no further changes are seen.<br><br>Q : $Cu^{2+}$ ✓ | ~~Upon adding aqueous ammonia~~, a light blue precipitate was formed, which dissolved in excess aqueous ammonia to form a deep blue solution ✓<br><br>[3] |
| (e) | Add about 1cm depth of solution **Q** into a clean test-tube.<br><br>To this test-tube, add aqueous sodium hydroxide slowly with shaking, until no further changes are seen. | ~~Upon adding aqueous sodium hydroxide~~, a light blue precipitate was formed which is insoluble in excess aqueous sodium hydroxide. ✓<br><br>[1] |

**(f) (i)** Using your observations, deduce the identity of salt **Q**.

For the cation and for the anion in salt **Q**, justify your deductions using evidence from your observations.

salt **Q** is Copper (II) chloride

Use either observation from testing NaOH or $NH_3$

justification for the cation Bf: In (c) Upon adding aqueous ammonia, a light blue precipitate formed which dissolved in excess aqueous ammonia to form a colourless solution. And when aqueous sodium hydroxide was added, a light blue precipitate formed which is insoluble in excess sodium hydroxide.

justification for the anion Bf: In (d) When silver nitrate was added after dilute nitric acid, a white precipitate was formed, which must be $Cl^-$.

[2]

**(ii)** Suggest one conclusion about **P** that can be drawn from your observations. Support your conclusion with evidence from your observations.

conclusion P contains carbonate anions.

evidence Bf: In (a), (b) When P was heated, the gas evolved reacted with limewater to form a white precipitate, meaning ~~it was likely a carbonate salt which decomposed to give the carbon dioxide gas that reacted with limewater.~~

[1]

**(g)** In this question, you are going to compare the concentrations of two solutions of a basic compound, $\mathbf{R_1}$ and $\mathbf{R_2}$.

| | | |
|---|---|---|
| **(i)** | Using a measuring cylinder, transfer 2.0 $cm^3$ of solution $\mathbf{R_1}$ into the test-tube labelled $\mathbf{R_1}$.<br><br>Using a different measuring cylinder, transfer 2.0 $cm^3$ of solution $\mathbf{R_2}$ into the test-tube labelled $\mathbf{R_2}$.<br><br>Add about 6 drops of solution **I** (screened methyl orange indicator) to both test-tubes. | |
| | Using a dropper pipette, add dilute nitric acid, drop by drop with shaking, to $\mathbf{R_1}$ until the indicator changes from green to purple.<br><br>Count and record the number of drops of acid added to solution $\mathbf{R_1}$. | number of drops of acid added to $\mathbf{R_1}$<br><br>= 9 |
| | Using **the same** dropper pipette, add dilute nitric acid, drop by drop with shaking, to $\mathbf{R_2}$ until the indicator changes from green to purple.<br><br>Count and record the number of drops of acid added to solution $\mathbf{R_2}$. | number of drops of acid added to $\mathbf{R_2}$<br><br>= 6<br><br>[2] |

**(ii)** Use your results from (g)(i) to deduce which of the solutions $\mathbf{R_1}$ or $\mathbf{R_2}$ is more concentrated. Explain your answer.

The more concentrated solution of **R** is solution $R_1$.

explanation Given the same volume of $R_1$ and $R_2$, $R_1$ took more nitric acid to reach the endpoint of the titration, suggesting there was a greater number of moles of the basic compound in $R_1$, and hence making $R_1$ more concentrated than $R_2$.

[1]

OR: Larger volume of nitric acid is needed to neutralise the more concentrated $R_1$.

[Total: 15]

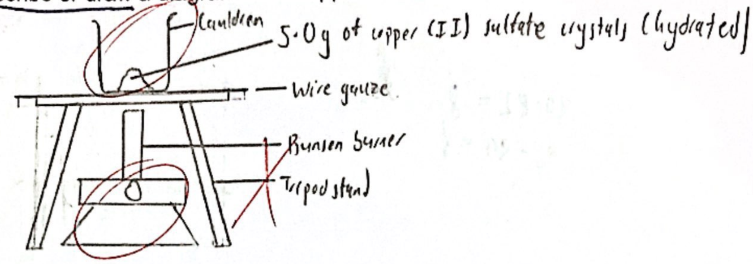
## 8.3 Experimental Planning

### 8.3.1 Timed Assignment 2022: Paper 3 Practical

3 Copper(II) sulfate crystals contain water of crystallisation which may be removed by heating.

(a) Given a sample of copper(II) sulfate crystals and the usual apparatus found in a school laboratory, plan an experiment to find the percentage by mass of water in copper(II) sulfate crystals.

You should include the measurements you would take and show how you would use your results to calculate the percentage by mass of water in copper(II) sulfate crystals. You should describe or draw a diagram of the apparatus that may be used to remove the water.




Diagram

Measure mass of content before and after heating using electronic balance.

Heat to ensure that all water has been removed/ heat to dryness/ heat to constant mass / heat until solid turns white.

mass before heating - mass after heating / mass of hydrated crystals × 100% = % water

Approach: Measure and record the change in mass after heating hydrated copper (II) sulfate crystals till they turn anhydrous. Then calculate the percentage by mass of water in hydrated $CuSO_4$

Steps: 1. ~~Place~~ Set up the experiment as shown in the diagram, but with an electronic balance where the mass of the empty cauldren crucible and the cauldren crucible with 5g of $CuSO_4$ crystals are measured with the electronic balance ~~where 5.0g of $CuSO_4$ crystals is measured with the electronic balance and placed in the~~

2. Use the bunsen burner to heat the hydrated $CuSO_4$ crystals till they become anhydrous.

3. Measure the mass of cauldren crucible with the anhydrous $CuSO_4$ left using the electronic balance.

4. Calculate change in mass (g) of the $CuSO_4$: Mass of cauldren with 5g of $CuSO_4$ crystals − mass of cauldren with anhydrous $CuSO_4$ left. [4]

5. Using the formula $\frac{\text{change in mass of } CuSO_4 \text{ (step 4)}}{5.0g} \times 100\%$, find the percentage by mass of water in the copper (II) sulfate crystals.

3

(b) The formula for copper(II) sulfate crystals is $CuSO_4.yH_2O$ where **y** is the number of moles of water of crystallisation in one mole of crystals. A student does an experiment and finds that **y** = 4. The correct value of **y** for her sample is 5.

Suggest an error in her experiment that would result in this difference. Explain how this error would lead to a lower value of **y**.

You can assume that all her measurements were read and recorded correctly and that her calculation was correct.

The hydrated copper (II) sulfate crystals did not fully turn anhydrous, and some water of crystallisation was left after heating in the crystals. Therefore, a higher mass of supposedly anhydrous $CuSO_4$ crystals is measured, resulting in a lower change in ~~vol~~ mass of the crystals, which is the mass of water in the crystals, OR ⇒ miscalculated. So, the percentage by mass of water in the hydrated $CuSO_4$ crystals also ~~decreased~~ decreases, resulting in a lower number of moles of water, and therefore a lower mole ratio of water, and y calculated. [2]

2

[Total: 6]

5

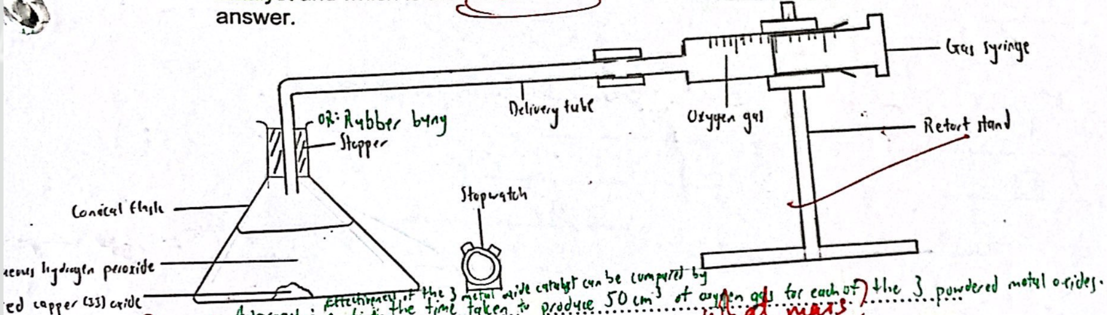
### 8.3.2 Hydrogen Peroxide decomposition, 4 April 2022

(f) Hydrogen peroxide decomposed easily, as shown in the equation below. [4]

$$2H_2O_2\,(aq) \rightarrow 2H_2O\,(l) + O_2\,(g)$$

The reaction may be catalysed by several different metal oxides including copper(II) oxide, iron (III) oxide and manganese (IV) oxide. Given a supply of aqueous hydrogen peroxide, the powdered metal oxides and the apparatus normally found in a school chemistry laboratory, describe an experiment which would enable you to decide which of the three metal oxides is the most effective catalyst and which is the ~~least effective~~. Include a labelled diagram in your answer.

2




approach?

Approach: Effectiveness of the 3 metal oxide catalyst can be compared by finding the time taken to produce 50 $cm^3$ of oxygen gas for each of the 3 powdered metal oxides.

what mass?

1. Set up the experiment as shown below, but add the powdered copper (II) oxide at the last. Upon adding the powdered copper (II) oxide, immediately seal the conical flask with the stopper and start the stopwatch. what vol. of $H_2O_2$?

2. Stop the stopwatch when 50 $cm^3$ of oxygen gas has been collected.

3. Repeat the experiment twice, using powdered iron (III) oxide ~~and~~ or maganese (IV) oxide each once in place of powdered copper (II) oxide.

Approach write at the top

4. Compare the time taken to produce 50$cm^3$ of oxygen gas for each of the 3 metal oxides. The metal oxide that takes the least amount of time to do so is the most effective catalyst. least effective?

** End of Practical **

2

Approach: Effectiveness of the 3 metal oxide catalysts can be compared by measuring the volume of oxygen gas produced after a fixed time.

1. Set up the experiment as shown above, adding 1.0g of powdered copper (II) oxide at the last.
2. Upon adding the 1.0g of powdered copper (II) oxide, immediately seal the conical flask using the stopper and start the stopwatch.
3. Stop the stopwatch when 50 $cm^3$ of oxygen gas has been collected.
4. Repeat the experiment twice, using powdered iron (III) oxide or manganese (IV) oxide each once in place of powdered copper (II) oxide.
5. The metal oxide that takes the least amount of time to produce 50 $cm^3$ of oxygen gas is the most effective catalyst, while the one that takes the longest time is the least effective catalyst.