PROSPERITY ACADEMY

A2 CHEMISTRY 9701

Crash Course

RUHAB IQBAL

COMPLETE P5 PREPARATION

EVEYRTHING YOU NEED TO KNOW!

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Linear law:- Compare to the form $y = mx + c$

a) $$\log y = A - \frac{\Delta H_{fusion}}{2.30 \times R T_m}$$

y axis :- $\log y$
x axis :- $1/T_m$

$$\underset{\underset{y}{\downarrow}}{\log y} = \underset{\underset{m}{\downarrow}}{\frac{-\Delta H_{fusion}}{2.30R}} \underset{\underset{x}{\downarrow}}{\left(\frac{1}{T_m}\right)} + \underset{\underset{c}{\downarrow}}{A}$$

b) $$\log\left(1/t_{av}\right) = n \log v + c$$

y axis :- $\log\left(1/t_{av}\right)$
x axis :- $\log v$

$$\underset{\underset{y}{\downarrow}}{\log\left(1/t_{av}\right)} = \underset{\underset{m}{\downarrow}}{n} \underset{\underset{x}{\downarrow}}{\log v} + \underset{\underset{c}{\downarrow}}{c}$$

c) $$E = E^{\ominus} + \frac{0.059}{3} \log\left[M_n^{z+}\right]$$

y axis = E
x axis = $\log\left[M_n^{z+}\right]$

$$\underset{\underset{y}{\downarrow}}{E} = \underset{\underset{m}{\downarrow}}{\frac{0.059}{3}} \underset{\underset{x}{\downarrow}}{\left(\log\left[M_n^{z+}\right]\right)} + \underset{\underset{c}{\downarrow}}{E^{\ominus}}$$

d) $$\log\left(\alpha - \alpha_{final}\right) = A - \frac{Kt}{2.30}$$

y axis :- $\log\left(\alpha - \alpha_{final}\right)$
x axis :- t

$$\underset{\underset{y}{\downarrow}}{\log\left(\alpha - \alpha_{final}\right)} = \underset{\underset{m}{\downarrow}}{\frac{-K}{2.30}} \underset{\underset{x}{\downarrow}}{(t)} + \underset{\underset{c}{\downarrow}}{A}$$

e) $$\log\left(\frac{D}{m}\right) = A + b \log[X]$$

y axis :- $\log\left(\frac{D}{m}\right)$

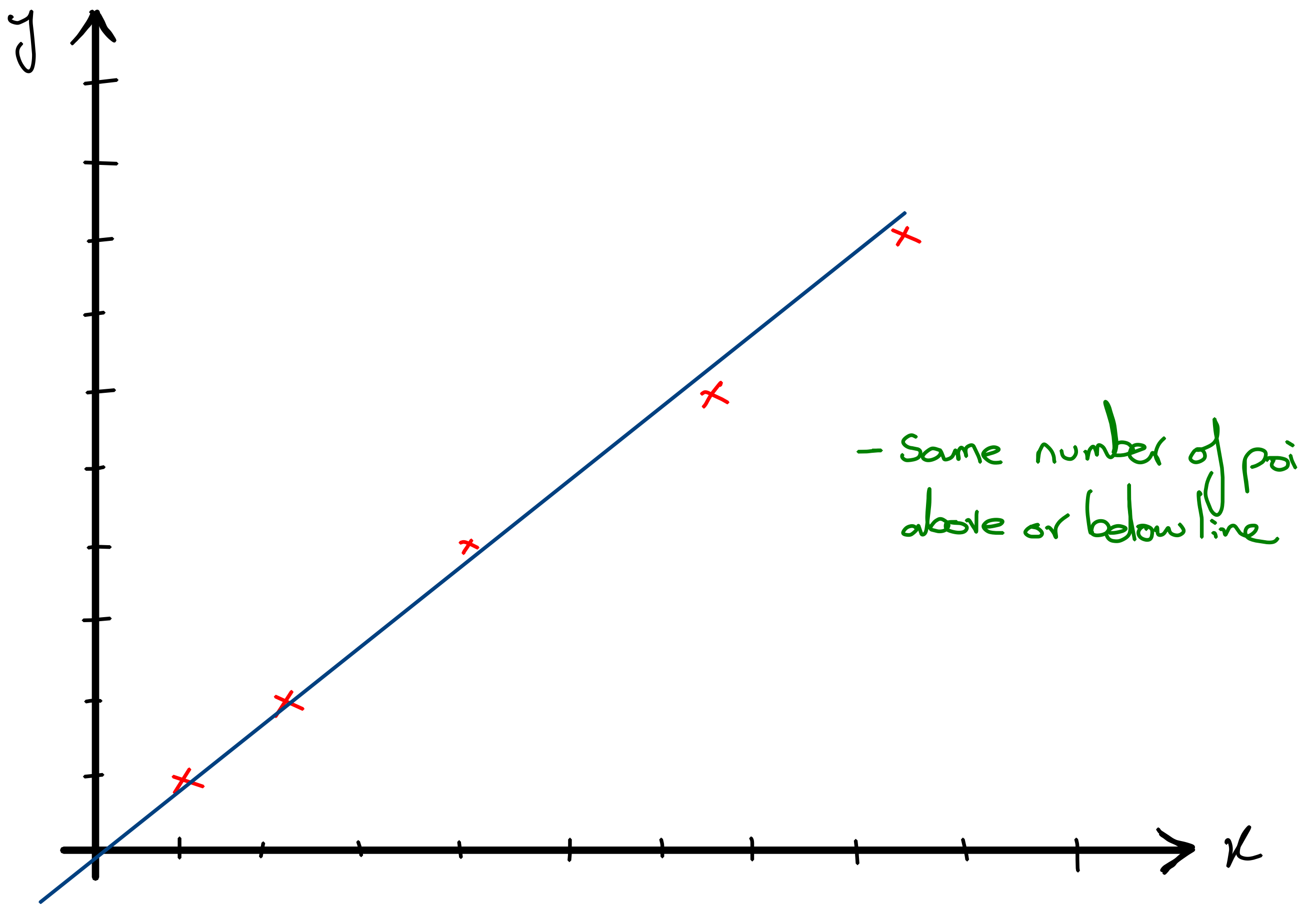
x axis :- $\log[X]$

$$\underset{\underset{y}{\downarrow}}{\log\left(D/m\right)} = \underset{\underset{m}{\downarrow}}{b} \underset{\underset{x}{\downarrow}}{\log[X]} + \underset{\underset{c}{\downarrow}}{A}$$

* logs have no units

Graphing :-

- Use x to mark points
- If asked to draw line of best fit ⟶ it could either be curved line or straight line (depends on points)
- If asked to draw straight line of best fit then draw only straight line
- If asked to draw curve of best fit then draw curve.



- Same number of points above or below line

Q. Do you consider results to be reliable?

Ans. Yes all points are on or very close to line

Percentage error :-

absolute error

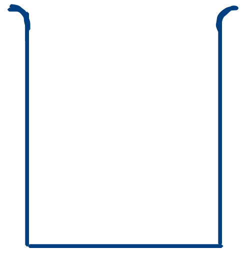
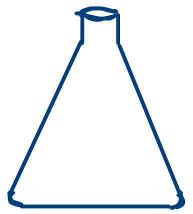
$$\frac{\Delta x}{x} \times 100 = \% \text{ error}$$

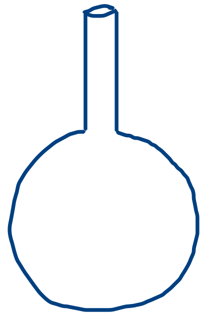
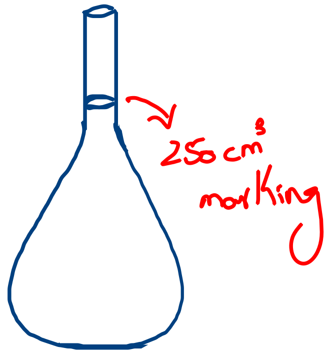
↳ principal value

To calculate % errors:-

- If only 1 reading is taken to find out the value = $\frac{\text{error in instrument}}{\text{measurement}} \times 100$

- If 2 readings are taken and measurement is difference = $\frac{2 \times \text{error in instrument}}{\text{measurement}} \times 100$

Apparatus & Diagram	Uses	Additional details/ Limitations:-
1) Beaker 	- For storing solutions - mild reactions - dissolving - heating/water bath	- Don't use for any measurements - Don't use for vigorous reactions, there is a chance of spilling - cannot be used to make exact solutions (Volumetric flask is used for this purpose) - Cannot be used for collecting gas as stoppers are not available.
2) Test tube 	- to collect and test small samples of substances - boiling tube is used for heating especially decomposition - collecting gases	- not suitable for storing large volumes - donot tell volume
3) Conical flask 	- suitable for reactions as:- 1) no spillage as small opening 2) can easily shake 3) stoppers available so gases dont escape - collecting gases - used for titrations	- donot tell volume

Apparatus & Diagram	Uses	Additional details/ Limitations:-
4) Round bottom flask 	- Used for heating	- donot tell volume - is not used for gas collection Don't confuse with volumetric flask
5) Volumetric flask  Other volumes such as $100cm^3$ are also available	- Used for making exact solutions only	

Method of making solution from mass:- Used to make high concentration solutions [These are always produced before making lesser concentration solutions as they give low % errors)

Q. Create a $0.1\ mol\ dm^{-3}$ solution of NaOH, volume $250cm^3$, from a 5.0g sample of solid NaOH.

1) Calculate moles and then mass of solid to be used:-

$$n = c \times V$$
$$= 0.1 \times 250 \times 10^{-3}$$
$$= 0.025\ mol$$

$$n = \frac{mass}{Mr/Ar} \Rightarrow n \times Mr/Ar = mass$$
$$= 0.025 \times (23+16+1)$$
$$= 1.0\ g$$

2) Measure out 1.0g of NaOH in a beaker using an electronic balance.
- First measure mass of container and zero the scale
- then put in the NaOH

3) Add enough water → typically $50cm^3$ or $100cm^3$ and stir so that all of the NaOH dissolves

4) Transfer the solution to a volumetric flask.

5) Transfer washings of beaker to volumetric flask

6) Wash stirrer and transfer its washings to flask too.

7) Add water to $250\,cm^3$ mark.

Method of making solution from existing solution through dilution :- Used to make low concentration solutions from high concentration solution

↳ so that % error is reduced

Q. Create a $0.005\,mol\,dm^{-3}$ of $Ca(OH)_2$, volume $250\,cm^3$ from $0.1\,mol\,dm^{-3}$ $Ca(OH)_2$ solution.

$$\underbrace{n}_{const} = c \times V \Rightarrow c_1 V_1 = n = c_2 V_2 \Rightarrow c_1 V_1 = c_2 V_2$$

$$\Rightarrow 0.005 \times 250 \times 10^{-3} = 0.1 \times V_2$$

$$V_2 = 0.0125\,dm^3 \rightarrow 12.5\,cm^3$$

Could be made by suitable mass

$n = c \times V$

$n = 0.1 \times 250 \times 10^{-3} = 0.025\,mol$

$mass = n \times M_r / A_r$

$= 0.025 \times (40.1 + 32 + 2)$

$= 1.8525 \simeq 1.85g$

then follow steps as described in making solution by mass method

1) Measure out $12.5\,cm^3$ solution of $Ca(OH)_2$ in a measuring cylinder

2) Transfer solution to volumetric flask

3) Transfer washings of measuring cylinder to volumetric flask

4) Add water to $250\,cm^3$ mark

Q. Why is solution not made through mass when low concentrations are to be made?

Taking previous example :-

$n = c \times V$

$n = 0.005 \times 250\times10^{-3} = 1.25\times10^{-3}$

$\text{mass} = n \times M_r/A_r = (1.25 \times10^{-3}) \times (40.1 + 32 + 2) = \underline{0.092625\,g}$

- too small to measure

* small mass has high percentage error

absolute uncertainty $\leftarrow \dfrac{\Delta x}{\underset{\hookrightarrow \text{principal}}{x}} \times 100 = \%\ \text{error}$

Apparatus & Diagram

6) Burette
(Clamping needed)

Colourless:-

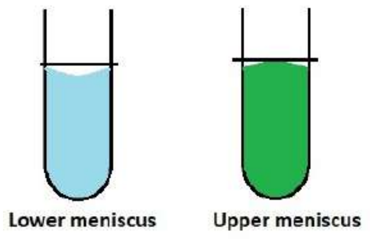
Read as $21.70\ cm^3$

Uses

- measures accurate amounts of volume upto 2 d.p. in cm^3
- used when very exact volumes are to be found out → titrations
- Since the burette is graduated to $0.1cm^3$, it's uncertainty is $\pm0.05cm^3$

Additional details / Limitations:-

- avoid parallax error
-

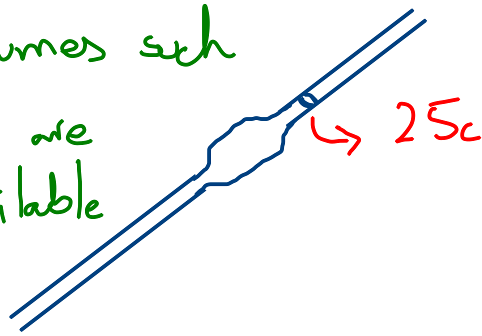
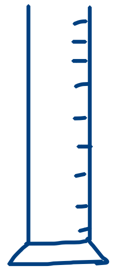
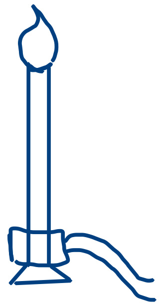
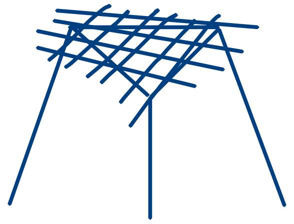


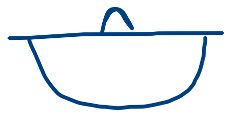
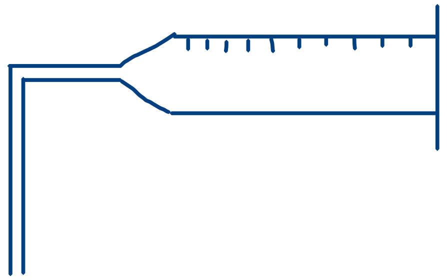
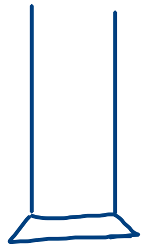
Read lower meniscus for colourless solutions

Read upper meniscus for coloured solutions

- slow method of transfer
- rinse burette with solution which will be filled in burette → makes sure concentration is as expected
- After filling, let some solution drain out to remove air.

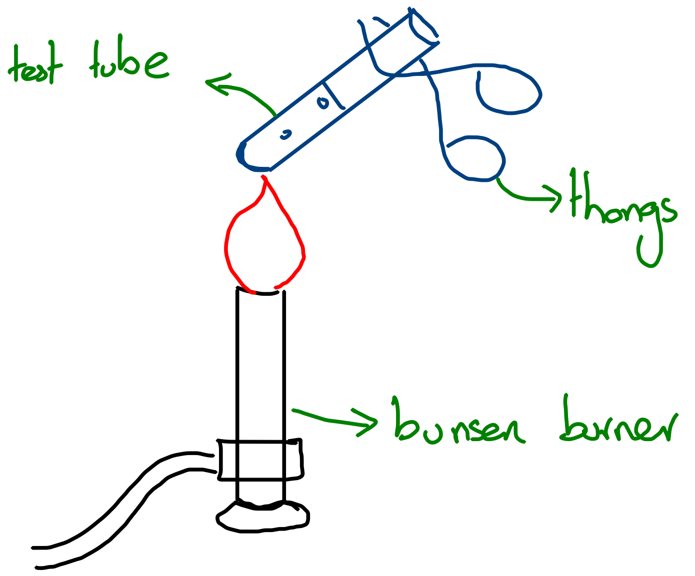
- Volumes are usually found out using subtraction of final and initial volumes → Error becomes $0.05 + 0.05 = \pm0.1\ cm^3$

Apparatus & Diagram	Uses	Additional details/Limitations:-
7) Pipette + pipette filler other volumes such as 10cm^3 are also available  ↳ 25cm^3 mark	- Very accurately transfer only 25 cm^3 of solution - an error of 25cm^3 - fast method of transfer	- You cannot vary the amount of volume
8) Measuring Cylinder 	- quickly measure out volumes → for substances in excess/catalysts - can be used to collect gas in downward displacement of water	- not very accurate
9) Bunsen burner 	- for strong heating - 800°C - 1000°C : flame temperature	- heat is not distributed evenly - Wear gloves - Use thongs to handle hot apparatus - donot use to heat flammable substances - incase of harmful gases being produced, conduct experiment in fume cupboard
10) Tripod stand + wire gauze 	- used to distribute heat evenly - can be used to heat crucible or beaker as it provides a platform	- to reduce heat losses, use insulation such as styrofoam cups, lids, insulating jacket

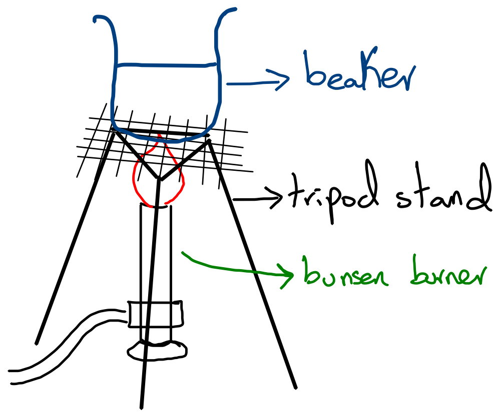
Apparatus & Diagram	Uses	Additional details/ Limitations:-
11) Ceramic Crucible 	- used for strong heating	
12) Gas syringe 	- used to collect and measure volumes of gas	- Keep surrounding temperature constant - prevents suck back effect
13) Gas jar 	-used to collect gases	
14) Electronic balance 	-used to measure mass	

Heating and cooling techniques:-

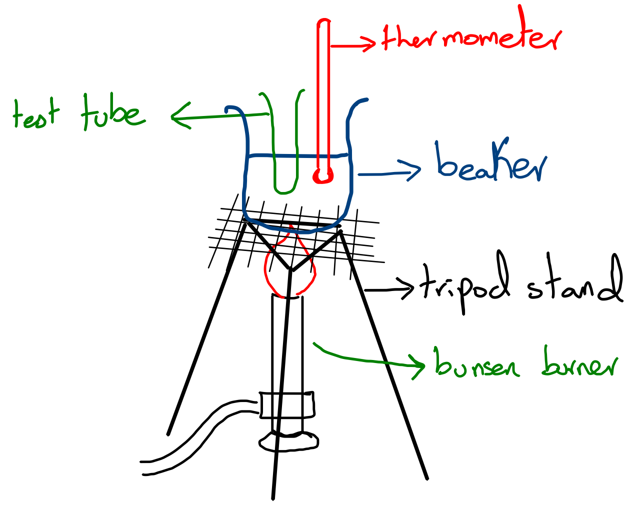
1) Heating for short time/small samples



2) Heating for long amount of time/large amounts of sample/uniform heating:-

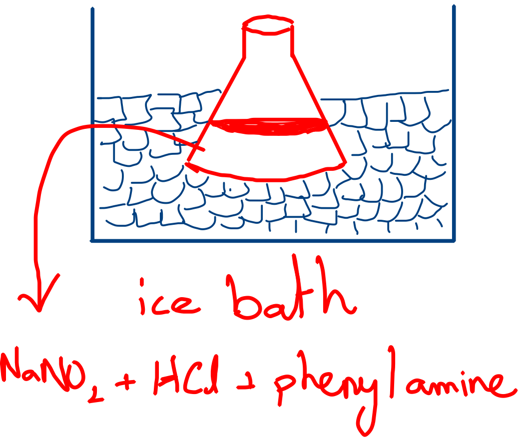


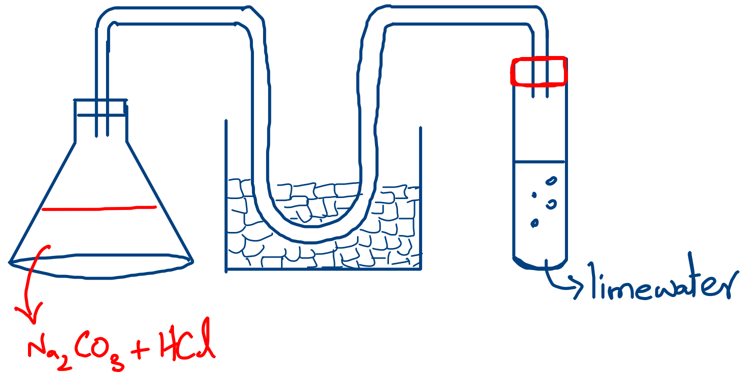
3) Heating flammable substances/temperature needs to be kept constant, very uniform heating



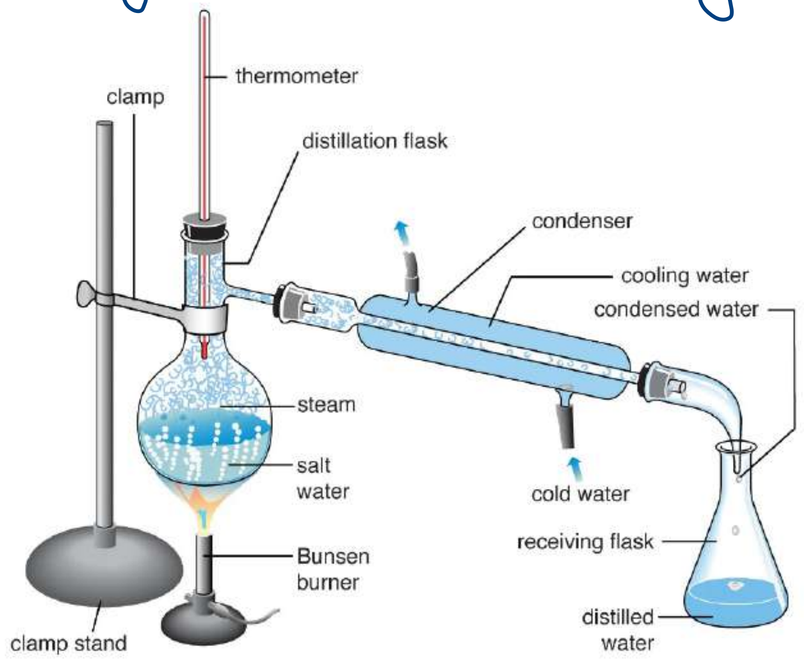
Water bath

4) Cooling the apparatus, just enough (Condensation of gas not required)





5) Liebeg condenser used to cool gases into liquids during distillation



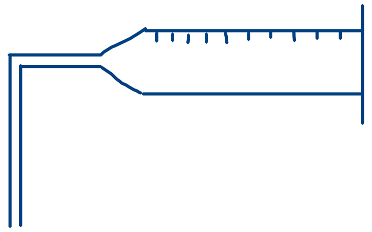
Safety Precautions:-

- Wear gloves
- Use thongs to handle hot apparatus
- donot use to heat flammable substances
- incease of harmful gases being produced, conduct experiment in fume cupboard
- to reduce heat losses, use insulation suchas styrofoam cups, lids, insulating jacket

Techniques involved when dealing with gases:-

1) Use a gas syringe to measure volume of gas collected

Gas syringe:-



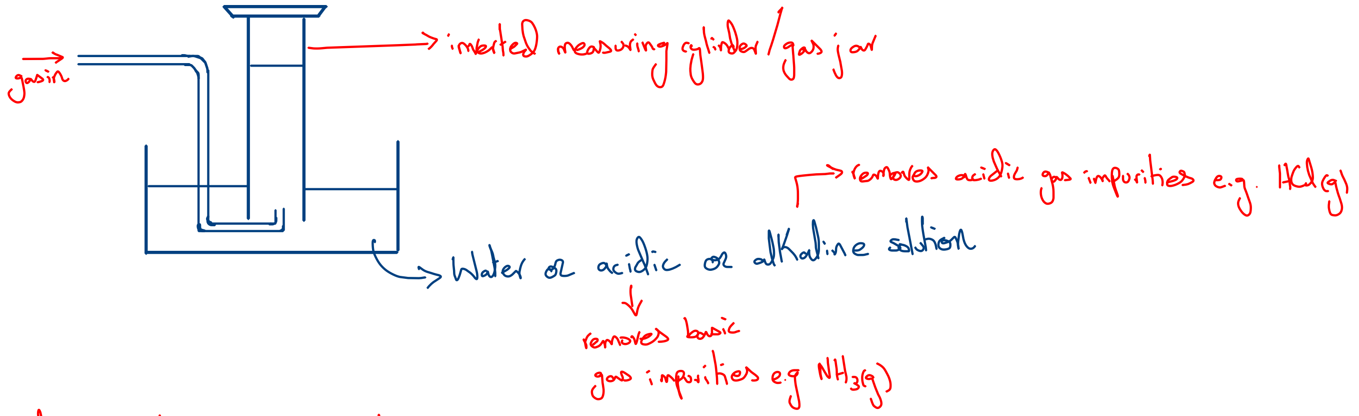
* Can be used for both soluble and insoluble gases

* Discard initial samples of gas as they are contaminated by air.

* suckback effect cannot happen here

Safety Precautions:- 1) Discard initial samples of gas as they are contaminated by air
2) Remove delivery tube before stopping heating to prevent suckback.

2) Downward displacement of water / Aqueous solution:-

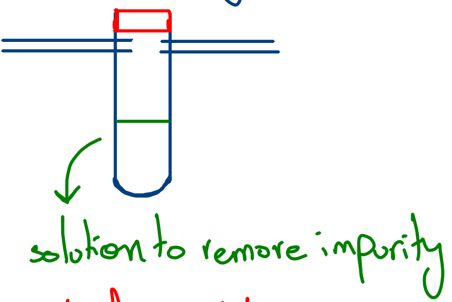


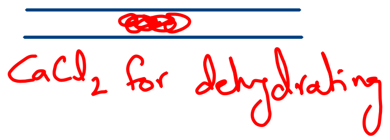
* used to collect insoluble gases e.g. H_2 or Cl_2

* used to remove impurities

* cannot be used to collect water soluble gases e.g. NH_3, HCl

3) Purifying a gas sample:-

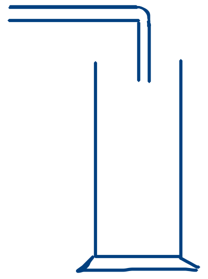




- Acidic solution removes basic gases
- basic solution removes acidic gases
- KOH removes Carbon dioxide (CO_2)
- $H_2SO_{4(aq)}$ or $CaCl_{2(s)}$ to remove water vapour (Dehydration)

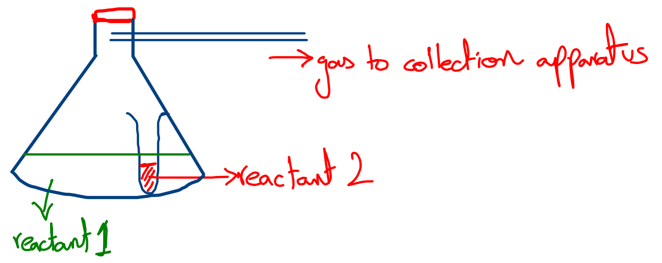
4) Upward / downward delivery:-

upward delivery for gases less dense than air (NH_3)



downward delivery for gases more dense than air (CO_2)

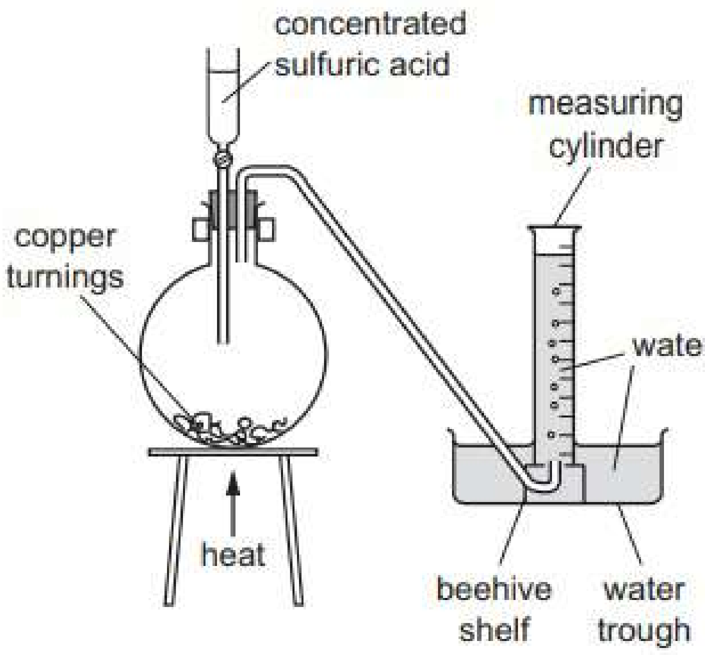
5) Mixing 2 samples together while making sure no gas escapes:-



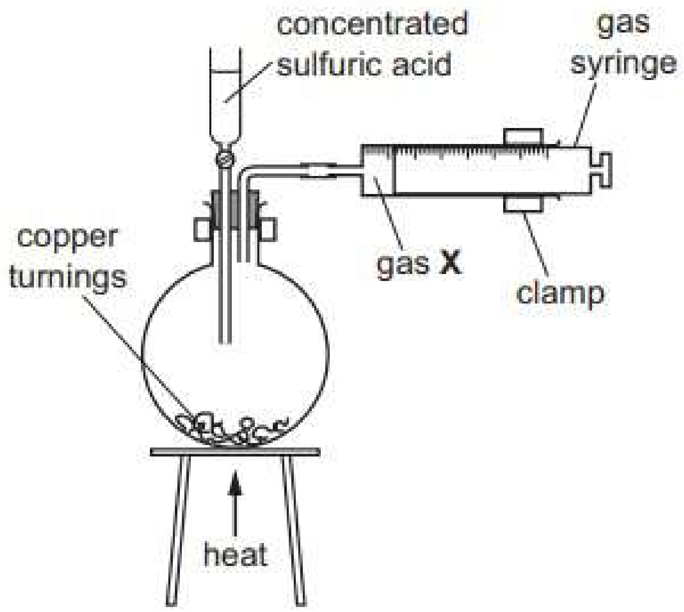
The small test tube tips over upon shaking

1 A student plans to carry out an experiment to find the relative molecular mass, M_r, of a soluble acidic gas, **X**, by finding the mass of a measured volume of gas **X**. Gas **X** can be prepared by the reaction between concentrated sulfuric acid and copper.

Two methods of gas collection are available to the student, as shown.



method **A**



method **B**

(a) **(i)** Explain why the first sample of gas collected from either apparatus should not be used for the M_r determination of gas **X**.

It is contaminated by air

[1]

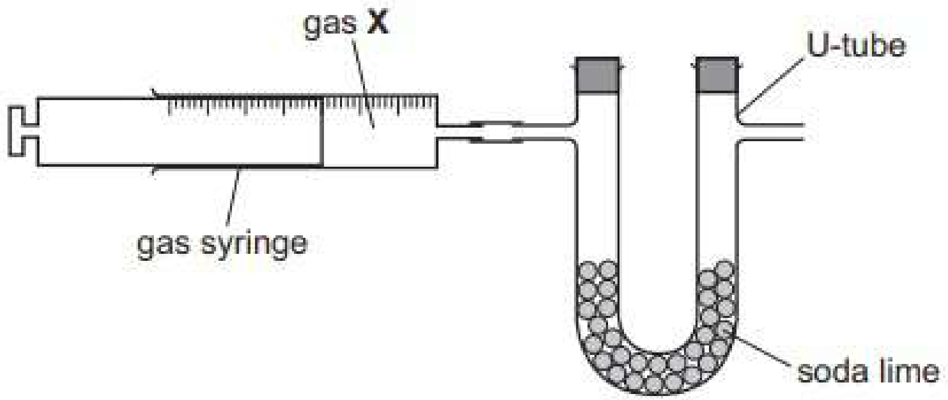
(ii) Give **two** reasons, other than your answer to **(a)(i)**, why, for this experiment, the apparatus in method **A** is less suitable than the apparatus in method **B** for collecting a sample of gas **X**.

reason 1 The gas produced may be soluble in water

reason 2 There is a danger of suck back effect

[2]

(b) The student is told to use the U-tube shown to find the mass of a sample of gas **X**.



A 100.0 cm^3 sample of pure gas **X** is placed in a gas syringe. The gas syringe is attached to a U-tube containing small lumps of solid soda lime, a mixture of sodium hydroxide and calcium hydroxide. All of gas **X** is slowly passed into the U-tube and the mass of gas **X** absorbed determined.
The temperature and the pressure of the room are recorded.

(i) State the measurements that are needed to determine the mass of gas **X** absorbed.

mass of U-tube + soda lime before experiment

mass of U-tube + soda lime after experiment [1]

(ii) Suggest why soda lime is used to absorb gas **X**.

Soda lime is alkaline so a neutralisation reaction occurs.

[1]

(c) Gas **X** can cause respiratory distress.

State an appropriate precaution that should be taken when doing this experiment.

Carry out experiment in fume cupboard

[1]

(d) Experiment 1 is carried out at a temperature of 21 °C, a pressure of 9.8×10^4 Pa, and uses 100 cm³ of pure gas **X**.

(i) Calculate the number of moles of gas **X** present in Experiment 1.

You should assume that gas **X** behaves like an ideal gas and so use $PV = nRT$.

$R = 8.31\,\text{J}\,\text{K}^{-1}\,\text{mol}^{-1}$

$$pV = nRT$$
$$(9.8\times10^{4})(100\times10^{-6}) = n\,(8.31)(21+273.15)$$
$$n = 4.01\times10^{-3}$$

moles of gas **X** = 4.01×10^{-3} mol [2]

(ii) The sample of gas **X** is found to have a mass of 0.251 g.

Explain how the student should use this information and their results to determine the M_r of **X**.

Mr/Ar = mass/mol So divide $0.251/4.01\times10^{-3} = 62.6$

[1]

(iii) Not all of gas **X** is absorbed by the soda lime.

State what effect, if any, this has on the student's calculated value of the M_r of gas **X**.

It will be less. $\left(\downarrow M_r = \dfrac{\text{mass}\downarrow}{\text{mol}}\right)$

[1]

(iv) In Experiment 2, the same mass of gas **X** is used, but the student did not record the temperature and pressure.

The calculated M_r of **X** for Experiment 2 is higher than the value calculated by the student for Experiment 1.

State and explain how the value of $\frac{P}{T}$ changes from Experiment 1 to Experiment 2.

The value of P/T decreases as calculated moles are lesser so P/T is less

[1]

$$\uparrow M_r = \frac{\text{mass}\ \text{same}}{\text{mol}\downarrow} \Rightarrow PV = nRT \Rightarrow \downarrow\frac{P}{T} = n\downarrow\frac{R}{V}$$

(e) State how the reliability of the results in Experiment 1 could be improved.

Repeat and average the results

[1]

(f) A different gas, methylamine, is alkaline.

State a change that would have to be made to the apparatus so that the M_r of methylamine could be determined.

Instead of soda lime use H_2SO_4 in U-tube

[1]

Titrations :-

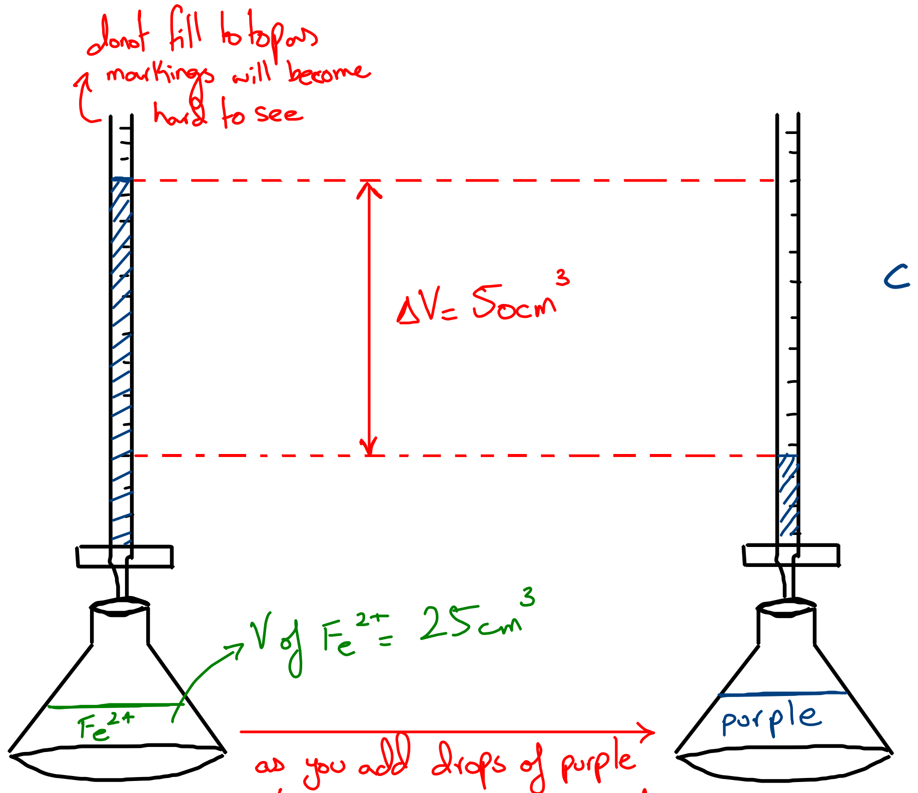
1) Acid - base :- Refer to AS Brønsted Lowry Acids and Bases.

2) Redox titration :- Used to find the concentration of Fe^{2+} ions

+2 to +3 : oxidised

$$\overset{+7\ -2}{MnO_4^-} + 5Fe^{2+} + 8H^+ \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

+7 to +2 : reduced



c of $KMnO_4 = 0.01\ mol\,dm^{-3}$

as you add drops of purple $KMnO_4$, the solution keeps returning to colourless state unless all of Fe^{2+} has reacted. (Shaking of conical flask is necessary to ensure reactants are mixed well together)

Q. The concentration of a solution of Fe^{2+} ions is to be found out. For this purpose, $25cm^3$ of acidified Fe^{2+} solution was titrated with $0.01\ mol\ dm^{-3}$ $KMnO_4$ solution. A total volume of $50cm^3$ of $KMnO_4$ was used. What is the concentration of Fe^{2+} ions in solution?

Apparatus used to measure out $25cm^3$ of Fe^{2+} solution: pipette

mol of $KMnO_4$ used $= c \times V = 0.01 \times 50\times10^{-3} = 5\times10^{-4}$ mol

mol of Fe^{2+} reacted acc/ balanced equation $= 5 \times 5\times10^{-4} = 2.5\times10^{-3}$ $\Rightarrow c = n/V = 2.5\times10^{-3}/25\times10^{-3} = \boxed{0.1\ mol\,dm^{-3}}$

2 'Lawn sand' is spread over the grass in gardens to reduce the growth of moss. Lawn sand is a mixture of sand and iron(II) sulfate crystals, $FeSO_4.7H_2O$.

Lawn sand usually contains 6–10% $FeSO_4.7H_2O$ by mass.

To determine the exact percentage by mass of $FeSO_4.7H_2O$ present in a sample of lawn sand, a student devises the following experiment.

step 1 Use a known mass of lawn sand to prepare 250.0 cm³ of solution **A** containing $Fe^{2+}(aq)$ ions. Solution **A** must have dilute sulfuric acid, $H_2SO_4(aq)$, added to it **before** it is made up to 250 cm³.

step 2 To determine the concentration of $Fe^{2+}(aq)$ in solution **A**, titrate a 25.00 cm³ sample of solution **A** against 0.0200 mol dm⁻³ aqueous potassium manganate(VII), $KMnO_4(aq)$.

The reaction which takes place during the titration is shown.

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$$

(a) (i) The end-point of the titration is 25.00 cm³ of 0.0200 mol dm⁻³ $KMnO_4(aq)$.

Determine the concentration of $Fe^{2+}(aq)$ that was present in the 25.00 cm³ sample of solution **A** at the start of the titration.

n of $KMnO_4$ used $= c \times V = 0.02 \times 25\times10^{-3}$

$= 5\times10^{-4}$

n of $Fe^{2+} = 5\times10^{-4} \times 5 = 2.5\times10^{-3}$

$C = \frac{2.5\times10^{-3}}{25\times10^{-3}} = 0.1$ mol dm^{-3}

concentration of $Fe^{2+}(aq)$ = 0.100 mol dm⁻³ [1]

If you were unable to calculate the concentration in **(i)**, assume for **(ii)** and **(iii)** that the concentration of $Fe^{2+}(aq)$ is 0.300 mol dm⁻³. This is **not** the correct answer.

(ii) Determine the mass of lawn sand needed to prepare the 250.0 cm³ of solution **A** at the concentration calculated in **(i)**.

Assume that lawn sand contains 8% $FeSO_4.7H_2O$ by mass.
[A_r: Fe, 55.8; S, 32.1; O, 16.0; H, 1.0]

n of $Fe^{2+} = 2.5\times10^{-3}$ in 25 cm^3

$2.5\times10^{-3} : 25$
$x : 250$

$x = 0.025$ mol of Fe^{2+} = mol of $FeSO_4.7H_2O$

mass = mol $\times M_r/A_r = 0.025 \times [55.8 + 32.1 + 4(16) + 14 + 7(16)]$

$= 6.9475$g

$\frac{6.9475}{y} \times 100 = 8 \Rightarrow y = 86.84375$

mass of lawn sand = 86.84 g [3]

(iii) Solution **A** must contain enough $H^+(aq)$ ions for the reaction to take place during the titration.

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$$

Use the concentration of $Fe^{2+}(aq)$ from **(i)** to determine the minimum volume of 2.00 mol dm⁻³ $H_2SO_4(aq)$ which must be added to prepare the 250.0 cm³ of solution **A**.

$Fe^{2+} : H^+$
$5 : 8$
$0.025 : x$

$x = 0.04$ mol of H^+ needed

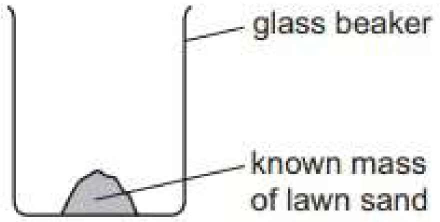
each mol of H_2SO_4 gives 2 mol of H^+

mol of $H_2SO_4 = 0.02$ mol

$C = n/V \Rightarrow 2 = \frac{0.02}{V} \Rightarrow V = 0.01$ dm^3 $\times 1000$
$= 10$cm^3

volume = 10 cm³ [2]

(b) Describe a method to prepare 250.0 cm³ of solution **A** starting with a glass beaker which contains the known mass of lawn sand determined in **(a)(ii)** as shown.



Assume that common laboratory apparatus is available.
You may find it helpful to write your answer as a series of smaller steps.

1) Add distilled water to dissolve the lawn sand
2) Filter the residue
3) Transfer the solution to a volumetric flask
4) Rinse the residue and add its washings to volumetric flask
5) Transfer the washings of the beaker to the volumetric flask
6) Add 10cm³ of H_2SO_4 using a measuring cylinder
7) Fill the volumetric flask with distilled water uptil 250 cm³ mark

[5]

(c) State the colour change in the conical flask at the end-point of the titration.

from colourless to purple

[1]

(d) Aqueous potassium manganate(VII) is a powerful oxidising agent.

Suggest the effect, if any, on the end-point volume if the student acidified the mixture with dilute hydrochloric acid, $HCl(aq)$, instead of dilute sulfuric acid, $H_2SO_4(aq)$. Explain your answer.

effect, if any, on the end-point volume higher

explanation Some of the MnO_4^- ions would get used up in oxidising Cl^- ions.

[2]

2 Ethanedioic acid is a white crystalline solid.

If excess aqueous potassium hydroxide, KOH(aq), is added to dilute ethanedioic acid, $H_2C_2O_4(aq)$, full neutralisation occurs and potassium ethanedioate, $K_2C_2O_4(aq)$, forms.

$$H_2C_2O_4(aq) + 2KOH(aq) \rightarrow K_2C_2O_4(aq) + 2H_2O(l)$$

If a small amount of potassium hydroxide is added, **partial** neutralisation takes place and not all H^+ ions in the acid are replaced by K^+ ions.

Instead an acid salt forms, which crystallises to form a solid with the formula $K_aH_b(C_2O_4)_c{\bullet}dH_2O$.

The letters *a*, *b* and *c* represent a ratio of the numbers of species present in the compound and may not necessarily be whole numbers. The relative number of water molecules associated with one formula of the compound is represented by *d*.

A student attempted to determine the values of *a*, *b*, *c* and *d* in a sample of an acid salt, $K_aH_b(C_2O_4)_c{\bullet}dH_2O$.

(a) The student wants to make a 250.0 cm³ aqueous solution of $K_aH_b(C_2O_4)_c{\bullet}dH_2O$, solution **A**.

The student adds 1.89 g of $K_aH_b(C_2O_4)_c{\bullet}dH_2O$ into a 100 cm³ beaker.

Describe the next steps the student should take to make solution **A**, containing exactly 1.89 g of $K_aH_b(C_2O_4)_c{\bullet}dH_2O$.

Give the name and capacity of the apparatus which should be used and describe how the student should ensure the volume is exactly 250.0 cm³.

Write your answer using a series of numbered steps.

1) Use distilled water to dissolve the acid salt in the beaker.
2) Transfer the solution to a 250cm³ Volumetric flask
3) Rinse the beaker and transfer its washings to the volumetric flask
4) Use distilled water to fill the volumetric flask upto 250cm³

[4]

(b) **Determining the number of moles of $C_2O_4^{2-}$ present**

Ethanedioate ions, $C_2O_4^{2-}(aq)$, react with manganate(VII) ions, $MnO_4^-(aq)$, in acidified conditions, as shown.

$$5C_2O_4^{2-}(aq) + 16H^+(aq) + 2MnO_4^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(g)$$

$MnO_4^-(aq)$ ions are a very deep purple in colour. All other species appear colourless.

The reaction takes place above a temperature of 70 °C.

The student carries out a redox titration using the following steps.

step 1 The student rinses and fills a burette with 0.0200 mol dm⁻³ $MnO_4^-(aq)$.

step 2 The student uses a pipette to transfer 25.0 cm³ of solution **A** into a conical flask.

step 3 The student adds 20 cm³, an excess, of 0.5 mol dm⁻³ $H_2SO_4(aq)$ to the conical flask.

step 4 The conical flask is heated until a temperature of about 80 °C is reached.

step 5 The student adds $MnO_4^-(aq)$ from the burette until an end-point is reached.

The student repeats the titration until concordant readings are achieved.

	rough	titration 1	titration 2	titration 3
final burette reading / cm³	25.05	24.50	26.60	24.50
initial burette reading / cm³	0.10	0.10	0.10	0.10
titre / cm³	25.05	24.40	26.50	24.40

The student determines the average titre to be 24.40 cm³.

(i) When emptying the pipette in **step 2**, the student touches the surface of the solution in the flask with the tip of the pipette.

Suggest why the student does this.

So that nothing is left in the pippete and exactly 25cm³ of solution A is transfered [1]

(ii) Suggest the most appropriate piece of apparatus to measure $H_2SO_4(aq)$ in **step 3**.

measuring cylinder [1]

(iii) Suggest why the student starts each titration with an initial burette reading of 0.10 cm³ rather than the usual 0.00 cm³.

$KMnO_4$ is coloured so if burette is filled to top, initial value will be very hard to read. [1]

(iv) What is meant by the term *concordant readings*?

titres with readings within 0.1cm³ of each other. [1]

(v) State the change of colour seen in the mixture in the conical flask at the end-point.

from colourless to purple [1]

(vi) Determine the number of moles of $C_2O_4^{2-}$ ions in the $250.0\,cm^3$ of solution **A**, $K_aH_b(C_2O_4)_c{\cdot}dH_2O$.

Give your answer to **three** significant figures.

n of $MnO_4^- = c \times V = 0.0200 \times 24.4 \times 10^{-3} = 4.88 \times 10^{-4}$

$MnO_4^- : C_2O_4^{2-}$

$2 : 5$

$4.88 \times 10^{-4} : x$

n of $C_2O_4^{2-}$ in $25\,cm^3 = 1.22 \times 10^{-3}$

$1.22 \times 10^{-3} : 25\,cm^3$

$y : 250\,cm^3$

$y = 0.0122$ mol

moles of $C_2O_4^{2-}$ ions in $250.0\,cm^3$ of solution **A** = 0.0122 mol [3]

If you were unable to calculate an answer to **(b)(vi)**, then you may use the value 1.18×10^{-2} mol for your calculations in **(c)**. This is **not** the correct value.

(c) The student then does an acid–base titration of solution **A** to determine the values of a and b in $K_aH_b(C_2O_4)_c{\cdot}dH_2O$.

(i) Suggest the name of a suitable reagent to use in this titration.

$NaOH_{(aq)}$ [1]

(ii) The student finds the concentration of H^+(aq) in solution **A** is $6.10 \times 10^{-2}\,mol\,dm^{-3}$.

Use this value and your answer to **(b)(vi)** to determine the ratio of c to b to **two** decimal places.

Then deduce the value of a in $\overset{+2}{(K_aH_b)}\overset{-2}{(C_2O_4)_c}{\cdot}dH_2O$ to **two** decimal places.

mol of H^+ in $250\,cm^3 = c \times V = 6.10 \times 10^{-2} \times 250 \times 10^{-3}$

$= 0.01525$

$c : b$

$0.0122 : 0.01525$

$1 : x$

$x = 1.25$

$a + b = 2$

$a + 1.25 = 2$

$a = 0.75$

ratio $c:b$ = 1 : 1.25

value of a = 0.75

[3]

(iii) Use your answer to **(b)(vi)**, **(c)(ii)** and other information given in the question to determine the mass of 1 mol of $K_aH_b(C_2O_4)_c{\cdot}dH_2O$ and hence determine the value of d to the nearest whole number.
[A_r: K, 39.1; H, 1.0; C, 12.0; O, 16.0]

If you were unable to calculate an answer to **(c)(ii)**, then you may use a = 0.30 and ratio $c:b$ = 1 : 1.60. These are **not** the correct values.

$Mr/Ar = \frac{Mass}{mol} = \frac{1.89}{0.0122} = 154.91$

theoretical mass of 1 mol = mol × Mr/Ar = 154.91

Mr/Ar of dehydrated $K_aH_b(C_2O_4) = (39.1)(0.75) + 1.25 + 2(12) + 4(16)$

$= 118.575$

mass of 1mol dehydrated = mol × Mr/Ar = 118.575

mass of $H_2O = 154.91 - 118.575 = 36.335$

mol of $H_2O = \frac{mass}{Mr/Ar} = \frac{36.335}{18} \simeq 2$

mass of 1 mol of $K_aH_b(C_2O_4)_c{\cdot}dH_2O$ = 154.9 g

value of d = 2

[2]

(d) A second student uses another method to determine d. Crystals of the sample, with known values of a, b and c, are heated in a crucible to remove the water molecules.

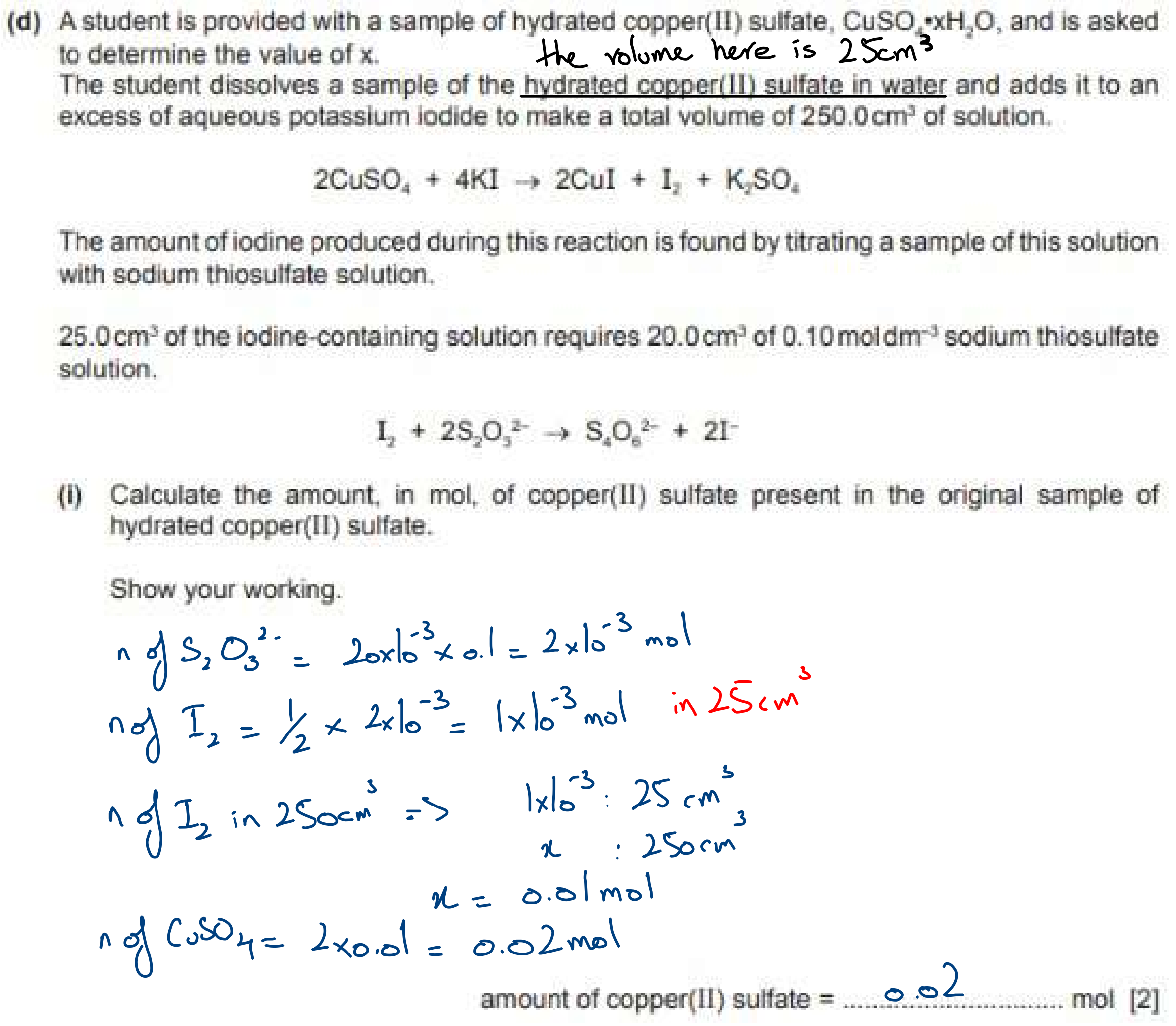
Construct a results table to show the readings that would need to be taken during this experiment.

mass of empty crucible /g	
mass of crucible + crystals before heating /g	
mass of crucible + crystals after heating /g	
mass of crucible + crystals after reheating /g	
mass of crucible + crystals after reheating again /g	

[2]

3) Iodometry :- Used to determine the concentration of an oxidising agent using iodine as the intermediary.

- First an oxidising agent is reacted with an iodide salt usually Potassium iodide (KI)
- As a result $I^-_{(aq)}$ ions are oxidised to $I_{2(aq)}$ (-1 to 0)
- The resulting $I_{2(aq)}$ is titrated with Sodium thiosulphate ($Na_2S_2O_3$) of Known concentration
- The moles of I_2 reacted and hence the moles of the oxidising agent and its concentration is then found out.



(d) A student is provided with a sample of hydrated copper(II) sulfate, $CuSO_4{\cdot}xH_2O$, and is asked to determine the value of x.

the volume here is 25cm³

The student dissolves a sample of the hydrated copper(II) sulfate in water and adds it to an excess of aqueous potassium iodide to make a total volume of 250.0 cm³ of solution.

$$2CuSO_4 + 4KI \rightarrow 2CuI + I_2 + K_2SO_4$$

The amount of iodine produced during this reaction is found by titrating a sample of this solution with sodium thiosulfate solution.

25.0 cm³ of the iodine-containing solution requires 20.0 cm³ of 0.10 mol dm⁻³ sodium thiosulfate solution.

$$I_2 + 2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2I^-$$

(i) Calculate the amount, in mol, of copper(II) sulfate present in the original sample of hydrated copper(II) sulfate.

Show your working.

n of $S_2O_3^{2-} = 20\times10^{-3} \times 0.1 = 2\times10^{-3}$ mol

n of $I_2 = \frac{1}{2} \times 2\times10^{-3} = 1\times10^{-3}$ mol in 25cm³

n of I_2 in 250cm³ ⇒ $1\times10^{-3} : 25\text{ cm}^3$

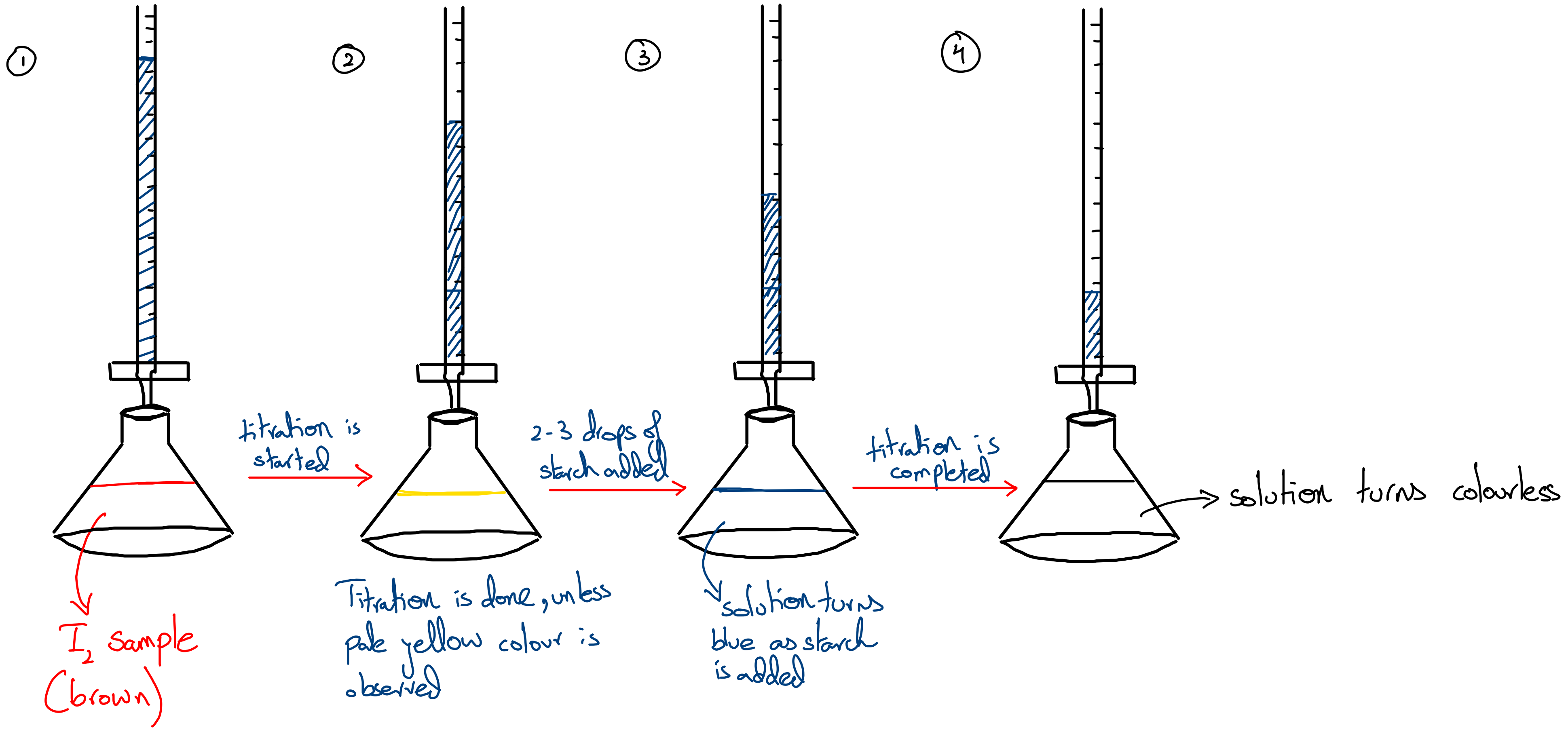
$x : 250\text{cm}^3$

$x = 0.01$ mol

n of $CuSO_4 = 2\times0.01 = 0.02$ mol

amount of copper(II) sulfate = ...0.02.......... mol [2]

Concentration? ⇒ $c = \frac{n}{V} = \frac{0.02}{25\times10^{-3}} = 0.8 \text{ mol dm}^{-3}$



1 Brass is an alloy of copper and zinc. Typical copper concentrations vary from 50% to 85%, depending upon the properties needed in the alloy. There may be small amounts of other metals present.

A student found a method to determine the percentage of copper in a sample of brass.

A known mass of brass powder is reacted with excess concentrated nitric acid. Both the copper and the zinc and any other metals present are oxidised into aqueous ions by the nitric acid. The amount of $Cu^{2+}(aq)$ ions present can be determined by a titration technique.

step 1 Use a weighing boat to accurately weigh by difference approximately 2 g of brass powder and place the brass into a small glass beaker.

step 2 In a fume cupboard add **approximately** 20 cm^3 of concentrated nitric acid to the brass in the beaker. Allow the brass to completely react to form solution **A**.

The equation for the reaction is shown.

$$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

step 3 Dilute **all** of solution **A** to form exactly 250.0 cm^3 of solution **B**.

step 4 Place 25.00 cm^3 of solution **B** into a conical flask.

step 5 Use a dropping pipette to add aqueous sodium carbonate, $Na_2CO_3(aq)$, to solution **B** in the conical flask until there is no more acid present.

step 6 Add approximately 20 cm^3 of aqueous potassium iodide, KI(aq), to the conical flask. A white precipitate forms as well as a brown solution of aqueous iodine, $I_2(aq)$.

step 7 Fill a burette with 0.100 mol dm^{-3} sodium thiosulfate solution, $Na_2S_2O_3(aq)$, so it is ready for the titration in **step 8**.

step 8 Carry out a titration of the aqueous iodine produced in the conical flask against the 0.100 mol dm^{-3} $Na_2S_2O_3(aq)$.

(a) Outline how the student should accurately weigh by difference in **step 1** in order that the exact mass of brass transferred into the small glass beaker is known. Include a results table, with appropriate headings, ready for the student to fill in.

The student should first measure the mass of the weighing boat + brass and after transferring should measure the mass of the weighing boat again. This should be done using an electronic balance and both masses be subtracted to get mass of brass

Mass of weighing boat + brass /g	Mass of empty weighing boat /g	mass of brass added /g

[2]

(b) Suggest why it is necessary to do **step 2** in a fume cupboard.

NO_2 which is a toxic gas is released [1]

(c) Outline how the student should carry out **step 3**. Include the name and capacity of the suitable piece of apparatus in which solution **B** should be prepared.

A 250 cm^3 volumetric flask must be used. Transfer solution A to volumetric flask Then add washings of beaker to flask and then add distilled water until 250 cm^3 mark.

[2]

(d) Name the apparatus needed to transfer solution **B** into the conical flask in **step 4**.

pipette [1]

(e) State how the student would know there was no more acid present in the mixture in **step 5**.

no more efferescence seen [1]

(f) The student is given 200 cm^3 of 0.100 mol dm^{-3} $Na_2S_2O_3(aq)$.

Outline how the student should use this solution to fill the burette in **step 7** so it is ready for titration. Include any relevant procedures the student should follow to ensure the burette is correctly filled before any readings are taken.

Rinse the burette with $Na_2S_2O_3(aq)$ and then fill the burette with it. After filling, allow some solution to drain out to remove air.

[2]

(g) The titration table the student used is shown.

(i) Complete the table and calculate the mean titre to be used in calculating the percentage of copper in brass.
Show your working.

titration number	rough	1	2	3	
final burette reading / cm^3	20.50	40.25	19.90	39.65	
initial burette reading / cm^3	0.00	20.60	0.00	19.90	
titre / cm^3	20.50	19.65	19.90	19.75	

$$\frac{19.65 + 19.75}{2} = 19.70$$

mean titre = 19.70 cm^3 [2]

(ii) The burette used by the student has graduations of 0.10 cm^3.

Determine the percentage error in the titre measured in titration number 2.

Show your working.

$$\frac{2(0.05)}{19.90} \times 100 = 0.5025$$

percentage error = 0.503% [1]

(iii) Other than a change in apparatus, suggest one change to the experiment which would lead to a reduction in the percentage error in a measured titre.

increase mass of brass

or use less concentration of $Na_2S_2O_3$ [1]

(h) **Steps 1–8** were repeated, this time using 1.88 g of brass. The end-point of the titration was found to be 16.50 cm^3.

The equations for the reactions occurring are shown.

equation 1 $2Cu^{2+}(aq) + 4I^-(aq) \rightarrow 2CuI(s) + I_2(aq)$
(step 6)

equation 2 $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$
(step 8)

(i) Determine the number of moles of I_2 formed when excess KI(aq) was added to 25.00 cm^3 of solution **B** in **step 6**.

Use the data from the repeated experiment in your calculations.

n of $Na_2S_2O_3 = c \times V = 0.100 \times 16.50 \times 10^{-3} = 1.65 \times 10^{-3}$

n of $I_2 = \frac{1.65 \times 10^{-3}}{2} = 8.25 \times 10^{-4}$

moles of I_2 = 8.25×10^{-4} [2]

(ii) Use your answer to **(h)(i)** to determine the mass of Cu^{2+} ions in solution **A** and therefore the percentage by mass of copper in this sample of brass.
If you were unable to obtain an answer to **(h)(i)**, assume the number of moles of I_2 to be 8.85×10^{-4} mol. This is **not** the correct value.
[A_r: Cu, 63.5]

n of $Cu^{2+} = (8.25 \times 10^{-4}) \times 2 = 1.65 \times 10^{-3}$

m of $Cu^{2+} = n \times Mr/Ar = 1.65 \times 10^{-3} \times 63.5 = 0.104775g$

$25cm^3 = 0.104775$
$250cm^3 = x$
$x = 1.04775$

$\%$ by mass $= \frac{1.04775}{1.88} \times 100$
$= 55.73$

percentage by mass of copper in the sample of brass = 55.7% [3]

(i) A small percentage of silver is sometimes found in some brass alloys.

In **step 2**, when concentrated nitric acid is added, silver metal is oxidised to silver ions, $Ag^+(aq)$.

At the end of **step 6** the $Ag^+(aq)$ ions no longer remain in solution.

Explain why.

Ag^+ displaces K^+ from its salt KI and forms AgI, a white ppt [1]

[Total: 19]

1 Iodide ions, I^-, and persulfate ions, $S_2O_8^{2-}$, react according to the following equation.

$$2I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(aq) + 2SO_4^{2-}(aq)$$

The rate of reaction between these ions can be determined from the time it takes for a certain amount of iodine, $I_2(aq)$, to be produced.

- A mixture of solutions is prepared, containing known volumes of
  - aqueous ammonium persulfate, $(NH_4)_2S_2O_8(aq)$,
  - aqueous sodium thiosulfate, $Na_2S_2O_3(aq)$,
  - starch indicator.
- A known volume of aqueous potassium iodide, KI(aq), is added to this mixture and a timer is started.
- After the reactants are mixed, they react slowly to produce iodine, $I_2(aq)$.
- Any iodine initially produced is removed by a reaction with thiosulfate ions.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

- Iodine, $I_2(aq)$, is continuously removed until all of the thiosulfate ions have been used up.
- After that time any $I_2(aq)$ that is produced turns the starch indicator blue.
- The time of the first appearance of the blue colour is recorded.
- This procedure is repeated with different volumes of reactants, keeping the total volume of the reaction mixture constant by adding the required volume of distilled water.

You are to plan a series of experiments to determine the effect of changing the concentration of iodide ions on the rate of reaction.

You are provided with the following materials.

solid ammonium persulfate, $(NH_4)_2S_2O_8(s)$
0.20 mol dm^{-3} aqueous KI, a source of $I^-(aq)$
0.0050 mol dm^{-3} aqueous $Na_2S_2O_3$, a source of $S_2O_3^{2-}(aq)$
starch indicator

(a) **(i)** Calculate the mass of $(NH_4)_2S_2O_8(s)$ that would be required to prepare 250 cm^3 of a standard solution of concentration 1.00 mol dm^{-3}.

[A_r values: N, 14.0; H, 1.0; S, 32.1; O, 16.0]

$n = c \times V = 1.00 \times (250 \times 10^{-3}) = 0.25$ mol

mass $= \text{mol} \times M_r / Ar = 0.25 \times (2(14) + 8 + 2(32.1) + 8(16))$

$= 57.05$ g

mass of $(NH_4)_2S_2O_8(s)$ = 57.1 g [1]

(ii) Describe how, after weighing the mass calculated in **(i)**, you would prepare this standard solution for use in your experiment.

Give the name and capacity, in cm^3, of any apparatus used.

Dissolve all of the solid in distilled water in a beaker.
Transfer contents of beaker to a volumetric flask of 250 cm^3.
Transfer the washings of the beaker to the flask and then
use distilled water to make solution upto 250 cm^3 mark.

[2]

(iii) Explain how the use of starch solution improves the accuracy of the experiment.

starch gives a sharp blue colour which is easily distinguishable as low concentrations of iodine produce pale yellow colour which is not visible easily. [1]

(b) A student planned five experiments to investigate the effect of iodide concentration, $[I^-]$, on the rate of reaction. The table shows the volumes used in experiment **1**.

Complete the table for experiments **2** to **5**.

experiment	volume of 1.00 mol dm^{-3} $(NH_4)_2S_2O_8(aq)$ / cm^3	volume of 0.20 mol dm^{-3} KI(aq) / cm^3	volume of water / cm^3	volume of 0.0050 mol dm^{-3} $Na_2S_2O_3(aq)$ / cm^3	volume of starch solution / cm^3
1	25.0	10.0	0.0	5.0	1.0
2	25.0	8.0	2.0	5.0	1.0
3	25.0	6.0	4.0	5.0	1.0
4	25.0	4.0	6.0	5.0	1.0
5	25.0	2.0	8.0	5.0	1.0

[3]

(c) In a different experiment, a student mixed the following solutions and measured the time taken for the reaction.

- 10.0 cm^3 of 1.00 mol dm^{-3} $(NH_4)_2S_2O_8(aq)$
- 5.0 cm^3 of 0.0050 mol dm^{-3} $Na_2S_2O_3(aq)$
- 5.0 cm^3 of 0.20 mol dm^{-3} KI(aq)
- 1.0 cm^3 of starch indicator

(i) The time taken for the blue colour to appear was 134 seconds (to the nearest second).

Calculate the rate of production of moles of I_2, in mol dm^{-3} s^{-1}.

n of $Na_2S_2O_3 = (5 \times 10^{-3}) \times 0.0050 = 2.5 \times 10^{-5}$

n of I_2 produced $= 1.25 \times 10^{-5}$

c of $I_2 = \frac{n}{V} = \frac{1.25 \times 10^{-5}}{(10+5+5+1) \times 10^{-3}} = 0.059524$

rate $= \frac{0.059524}{134} = 4.442 \times 10^{-6}$

rate of production of moles of I_2 = 4.44×10^{-6} mol dm^{-3} s^{-1} [3]

(ii) What should the student have done to make sure that the results were reliable?

Repeat the experiment and average the results

[1]

(iii) The 5.0 cm^3 of 0.0050 mol dm^{-3} $Na_2S_2O_3(aq)$ was measured using a 50 cm^3 burette which had graduations every 0.1 cm^3.

Calculate the maximum percentage error in the measured volume of this solution.

$\frac{2(0.05)}{5} \times 100 = 2$

percentage error = 2 % [1]

(d) A second student tried to perform the same experiment but found that the reaction mixture turned blue immediately after KI(aq) was added.

State what error the student had made.

He didnt add $Na_2S_2O_3$

[1]

(e) The following information gives some of the hazards associated with the chemicals used in the procedure.

Ammonium persulfate	Solid is **oxidising** and **hazardous to the environment**. Contact with combustible material may cause fire. It is classified as **health hazard**, is harmful if swallowed and is irritating to eyes, respiratory system and skin. Solutions equal to or more concentrated than 0.2 mol dm^{-3} should be labelled **health hazard** and **hazardous to the environment**. Solutions equal to or more concentrated than 0.05 mol dm^{-3} but less concentrated than 0.2 mol dm^{-3} should be labelled **health hazard**.
Potassium iodide	All solutions are low hazard.
Sodium thiosulfate	All solutions are low hazard.

Describe **one** relevant precaution, other than eye protection and a lab coat, that should be taken to keep the risk associated with the chemicals used to a minimum. Explain your answer.

Ammonium persulfate must be stated along with its hazard **and** linked to the precaution. ......

Ammonium persulfate is a skin irritant so wear gloves
OR ......
Ammonium persulfate is an irritant to the respiratory system; do the experiment in fume cupboard / face mask

OR [1]
Ammonium persulfate is harmful if swallowed so avoid mouth contact / wear face mask

OR 14]
Ammonium persulfate is oxidising so avoid contact with flammable / combustible materials.

OR
Ammonium persulfate is harmful / hazardous to the environment so do not dispose of down the drain / use (large quantities) of water to dilute before disposal

Enthalpy Changes:- Revisit Enthalpy Changes from AS

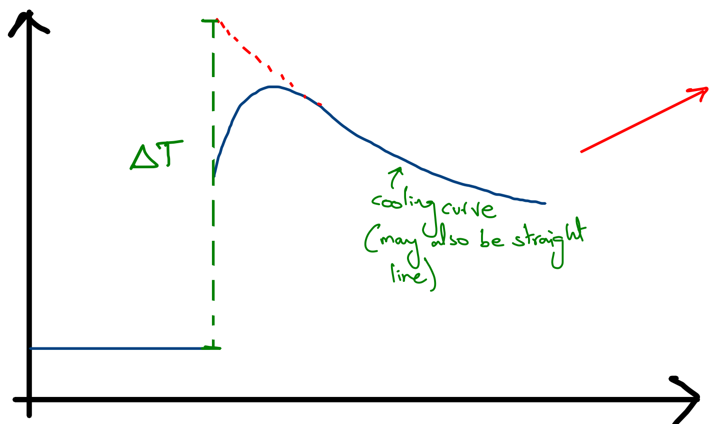
Important things to remember:-

- Heat energy released or absorbed => $Q = mc\Delta T$ (Put - sign if exothermic and + sign if endothermic)
- Convert Q in J to $KJ\ mol^{-1}$ by dividing by mol and multiplying by 10^{-3}

- $c = 4.18\ J g^{-1} K^{-1}$
- mass of solution in g = volume of solution in cm^3
- ΔT can be taken in °C or K.

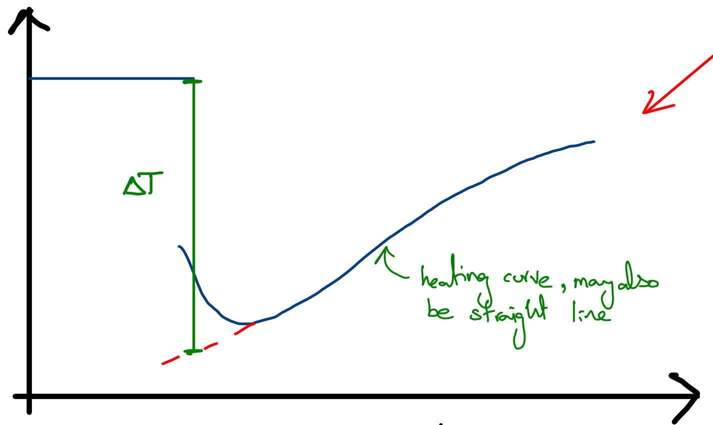
Problems that account for inaccuracy:-

1) Heat losses ⟶ minimised by use of styrofoam cups, beaker, insulating jacket and lids
2) In case of enthalpy changes of combustion, incomplete combustion may occur.
3) Small mass/volume used causes a greater % uncertainty
4) Precision of the thermometer is limitting
5) Experiment carried out under non standard conditions



exothermic reaction

extrapolation of the heating or cooling curve is done to find out actual ΔT as thermometer takes time to register changes in temperature



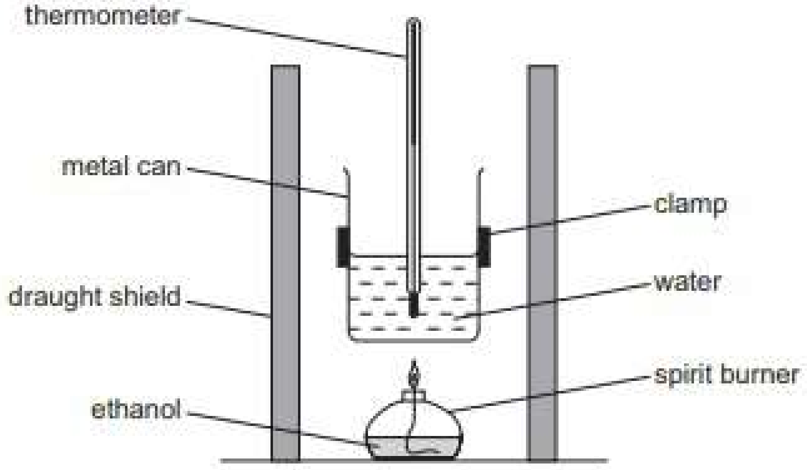
endothermic reaction

* Apparatus discussed in past papers ahead.

* Revise Hess' Law.

2 A student carried out a series of experiments to determine the enthalpy change of combustion of ethanol, C_2H_5OH.

A diagram of the apparatus is shown below.



The ethanol in the spirit burner was burned to heat a measured mass of water in the metal can. The student recorded the initial and final mass of ethanol and the initial and final temperature of the water for each experiment.

(a) Process the results in the table to calculate the amount of ethanol burned and the energy transferred to the water in each experiment.

Record all answers to **three** significant figures.
[Specific heat capacity of water, $c = 4.18\,J\,g^{-1}\,K^{-1}$]
[A_r: C, 12.0; H, 1.0; O, 16.0]

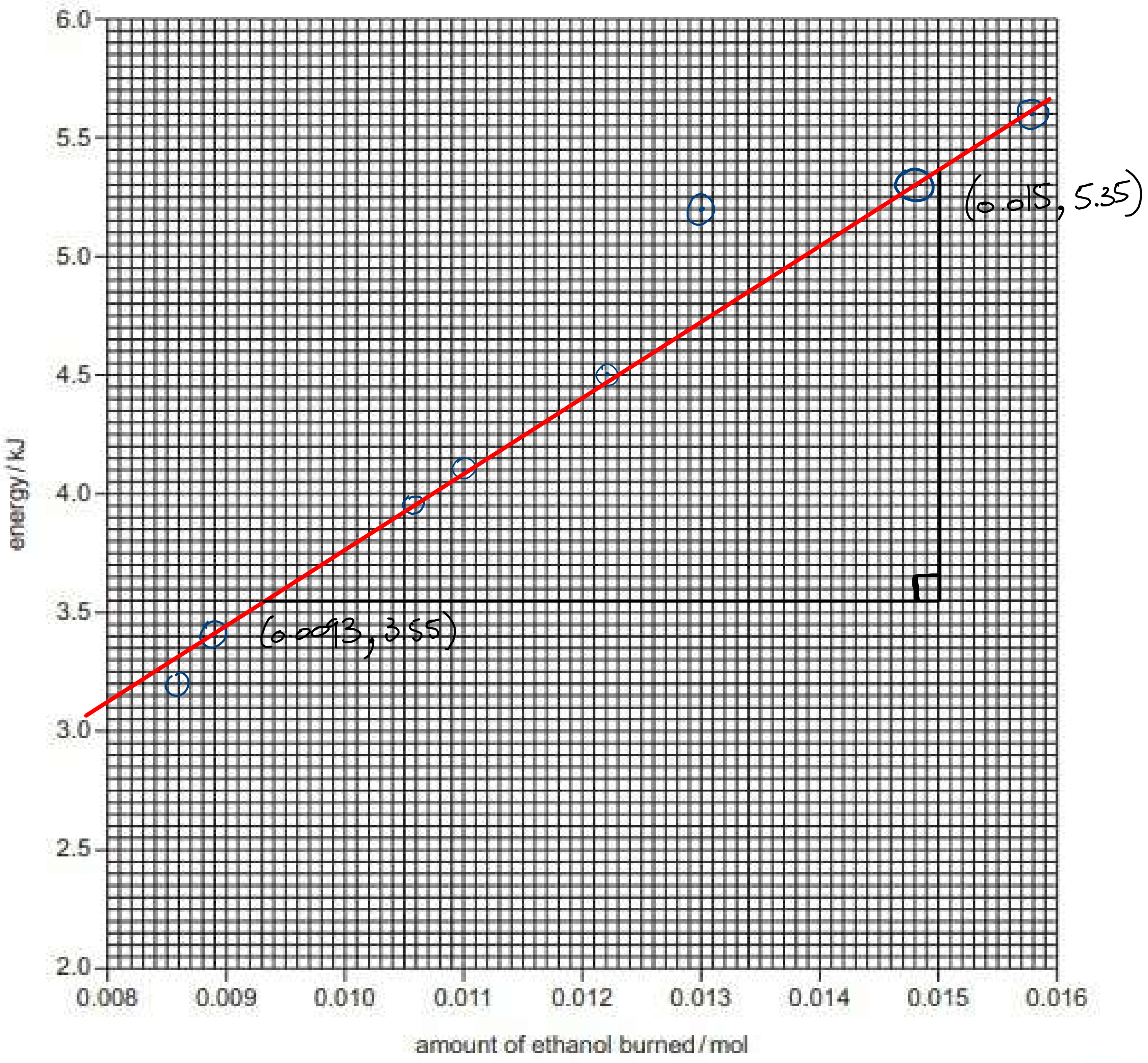
$$n = \frac{mass}{Mr/Ar} = \frac{0.391}{24+6+16} = 8.50 \times 10^{-3}$$

experiment number	mass of ethanol burned/g	temperature change/°C	mass of water heated/g	amount of ethanol burned/mol	energy transferred to the water/kJ
1	0.391	19.5	40.0	8.50×10^{-3}	3.26
2	0.488	23.6	40.0	0.0106	3.95
3	0.506	24.5	40.0	0.0110	4.10
4	0.559	26.9	40.0	0.0122	4.50
5	0.727	33.6	40.0	0.0158	5.62
6	0.597	31.1	40.0	0.0130	5.20
7	0.410	20.3	40.0	8.91×10^{-3}	3.39
8	0.681	31.7	40.0	0.0148	5.30

$Q = mc\Delta T = (40)(4.18)(19.5)/1000 = 3.26$

[2]

(b) Plot a graph on the grid below to show how the energy transferred to the water varies with the amount of ethanol burned. Draw the line of best fit.



[2]

(c) Write the number of the experiment which gave the result which was most anomalous.

Experiment 6 [1]

(d) The gradient of the graph gives the magnitude of the enthalpy change of combustion of ethanol.

Determine the gradient of your graph. State the co-ordinates of the **two** points you used for your calculation. Record the value of the gradient to **three** significant figures.

co-ordinates of two points used (0.0093, 3.55) (0.015, 5.35)

$$m = \frac{5.35 - 3.55}{0.015 - 0.0093} = 315.789$$

gradient = 316 kJ mol^{-1} [2]

(e) Under the same conditions of temperature and pressure as these experiments, the accepted value for the enthalpy change of combustion of ethanol is –1370 kJ mol^{-1}.

Explain why this value is a negative number.

The reaction is exothermic [1]

(f) **(i)** Calculate the maximum percentage error in the measurement of each mass used in experiment 1.

mass measured	maximum error in a single reading	maximum percentage error / %
0.391 g of ethanol burned	0.0005 g	$\frac{2(0.0005)}{0.391} \times 100 = 0.256\%$
40.0 g of water	0.05 g	$\frac{0.05}{40} \times 100 = 0.125\%$

[1]

(ii) Another student repeated the experiments using the method described under the same conditions. The value obtained for the enthalpy change of combustion of ethanol was –612 kJ mol^{-1}.

Suggest a reason why the errors calculated in **(i)** do not fully account for the difference between the student's value and the accepted value for enthalpy change of combustion, –1370 kJ mol^{-1}.

The percentage errors are too small to account for the big difference.
The heat losses are more significant. [1]

[Total: 10]

1 The enthalpy change of reaction, ΔH_r, for the decomposition of sodium hydrogencarbonate, $NaHCO_3(s)$, cannot be measured directly.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

A student must carry out **two** separate experiments and use the results of these experiments to determine the enthalpy change of reaction for the decomposition of sodium hydrogencarbonate.

(a) Suggest why the enthalpy change of reaction, ΔH_r, for the decomposition of sodium hydrogencarbonate cannot be measured directly.

It is impossible to measure the temperature change during direct heating

[1]

In both experiments the student used a weighing boat. A weighing boat is a small vessel used to contain solid samples when they are weighed.

Experiment 1 Reaction between sodium carbonate, $Na_2CO_3(s)$, and dilute hydrochloric acid, $HCl(aq)$

step 1 The student added approximately 3 g of $Na_2CO_3(s)$ to a weighing boat and accurately measured the combined mass of the weighing boat and $Na_2CO_3(s)$. This mass was recorded in Table 1.1.

step 2 The student used a measuring cylinder to measure 50 cm³ (50g) of 2 mol dm⁻³ $HCl(aq)$.

step 3 The experiment was carried out and the results were recorded in Table 1.2.

step 4 The student reweighed the empty weighing boat and recorded the mass in Table 1.1.

Table 1.1 mass results from Experiment 1

mass of weighing boat and $Na_2CO_3(s)$ / g	4.15
mass of empty weighing boat after addition of $Na_2CO_3(s)$ to $HCl(aq)$ / g	0.97
mass of $Na_2CO_3(s)$ added / g	

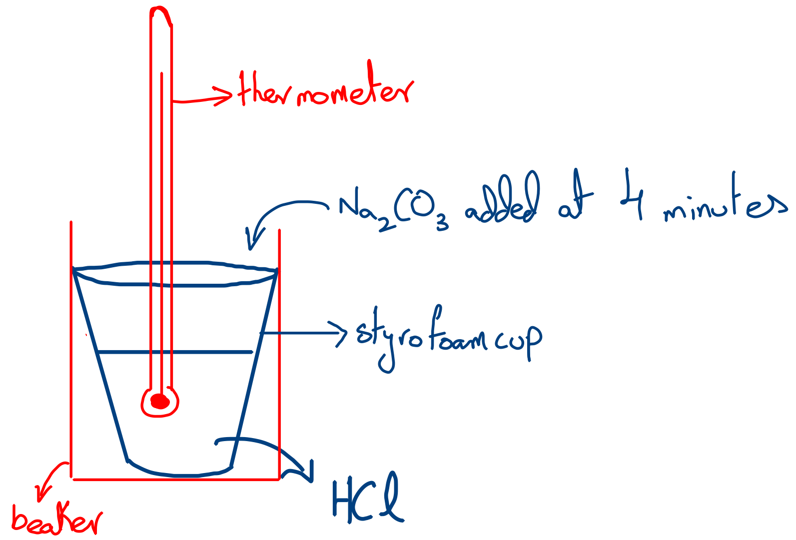
Table 1.2 temperature results from Experiment 1

time / minutes	0	1	2	3		5	6	7	8	9	10
temperature of mixture / °C	20.0	19.8	19.8	19.8		24.6	24.7	24.5	24.3	24.1	23.9

(b) Outline how the student carried out **step 3** of the experiment.
You may find it helpful to write your answer as a series of smaller steps.

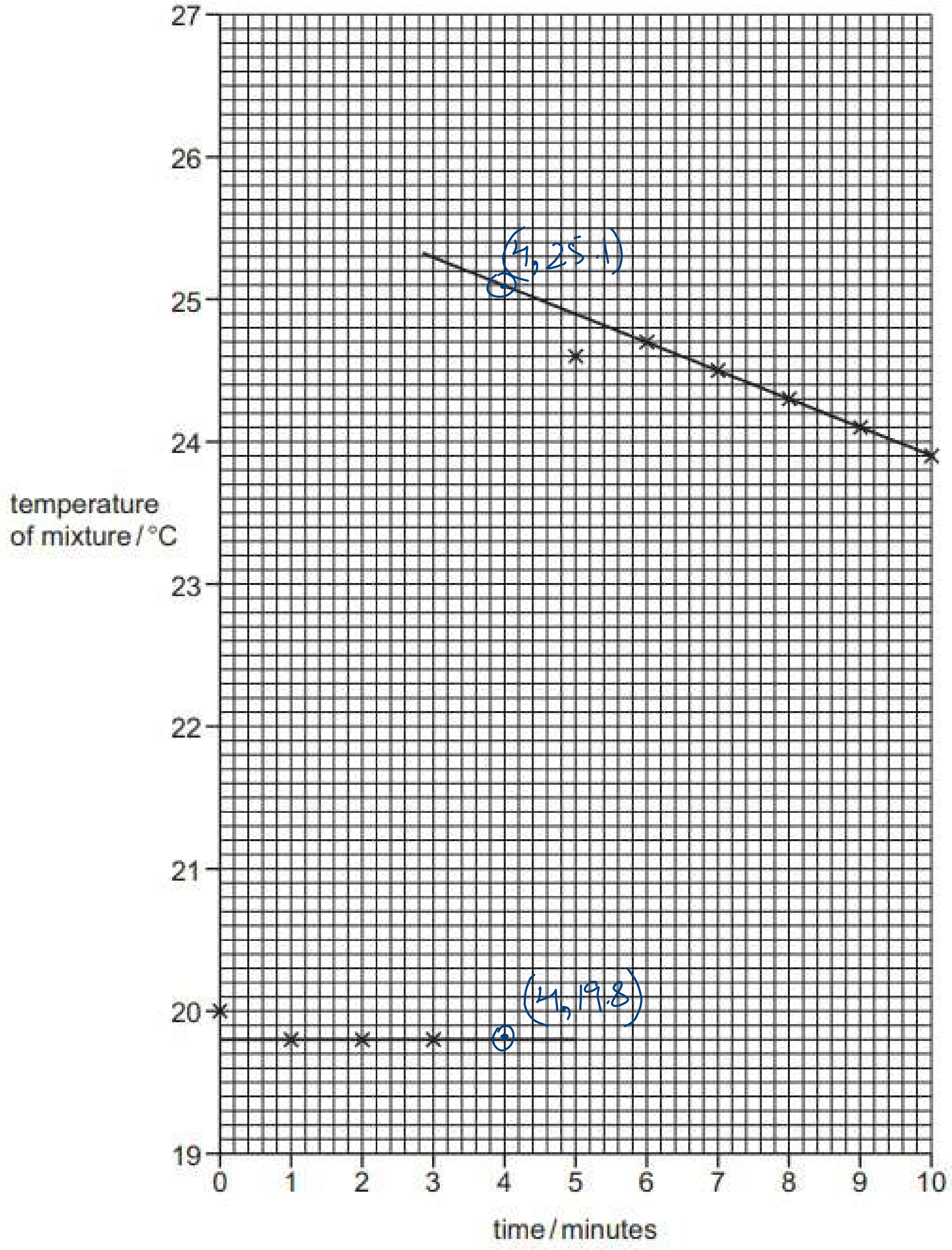
The apparatus is setup as shown. The student measures the temperature every minute. He waits for the first 3 minutes to get a steady initial temperature and then adds the Na_2CO_3 to the solution at 4 minutes

Draw a labelled diagram of the apparatus.



[3]

The student plotted a graph of the results and drew **two** lines of best fit which were both extrapolated as shown.



(c) Use the graph to determine the theoretical temperature increase at 4 minutes.

25.1 − 19.8 = 5.3

theoretical temperature increase = 5.3 °C [1]

(d) Use Table 1.1 on page 2 to determine the mass of $Na_2CO_3(s)$ which was added to the $HCl(aq)$. Use this value and your answer to **(c)** to determine the enthalpy change, ΔH_1, for the reaction shown.

$$Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Give your answer to **three** significant figures.
[Assume that the specific heat capacity of the solution is 4.18 J g^{-1} K^{-1}.]
[A_r: Na, 23.0; C, 12.0; O, 16.0]

m of Na_2CO_3 = 4.15 − 0.97 = 3.18g

n of $Na_2CO_3 = \frac{3.18}{2(23) + 12 + 3(16)} = 0.03$ mol

$Q = mc\Delta T = 50 \times 4.18 \times 5.3 = 1107.7$ J

$Q = \frac{1107.7 \times 10^{-3}}{0.03} = 36.9$ KJ mol^{-1}

ΔH_1 = − 36.9 kJ mol^{-1} [3]

(e) (i) Explain why the student did **not** add the $Na_2CO_3(s)$ to the $HCl(aq)$ at 0 minutes.

so that the acid reaches room temperature and the initial temperature is steady [1]

(ii) Suggest why the temperature measured at 5 minutes was lower than the temperature measured at 6 minutes.

The reaction was not complete

[1]

Experiment 2 Reaction between sodium hydrogencarbonate, $NaHCO_3(s)$, and dilute hydrochloric acid, $HCl(aq)$

step 1 The student weighed an empty weighing boat and recorded the mass in Table 1.3.

step 2 The student added exactly 4.20 g of $NaHCO_3(s)$ to the weighing boat and recorded the mass in Table 1.3.

step 3 The student carried out the same experimental procedure as in **steps 2** and **3** of Experiment 1.

Table 1.3 mass results from Experiment 2

mass of empty weighing boat / g	0.95
mass of weighing boat and $NaHCO_3(s)$ / g	5.15
mass of $NaHCO_3(s)$ added / g	

(f) Explain why the method of determining the mass of solid added in Experiment 2 is less accurate than the method of determining the mass of solid added in Experiment 1.

The difference was not taken so the exact mass transferred will be different than this value [1]

(g) (i) In Experiment 2 a 50 cm³ measuring cylinder was used to measure the 50 cm³ of $HCl(aq)$. The 50 cm³ measuring cylinder had 1 cm³ graduations.

Calculate the maximum percentage error in measuring 50 cm³ of $HCl(aq)$ with this 50 cm³ measuring cylinder.

$$\frac{0.5}{50} \times 100 = 1$$

maximum percentage error = 1 % [1]

(ii) Explain why measuring the concentration of the 2 mol dm⁻³ HCl more precisely would **not** affect the result of the experiment.

HCl is in excess

[1]

(iii) Suggest what the student should change to reduce the percentage error associated with the temperature readings **without** changing the apparatus.

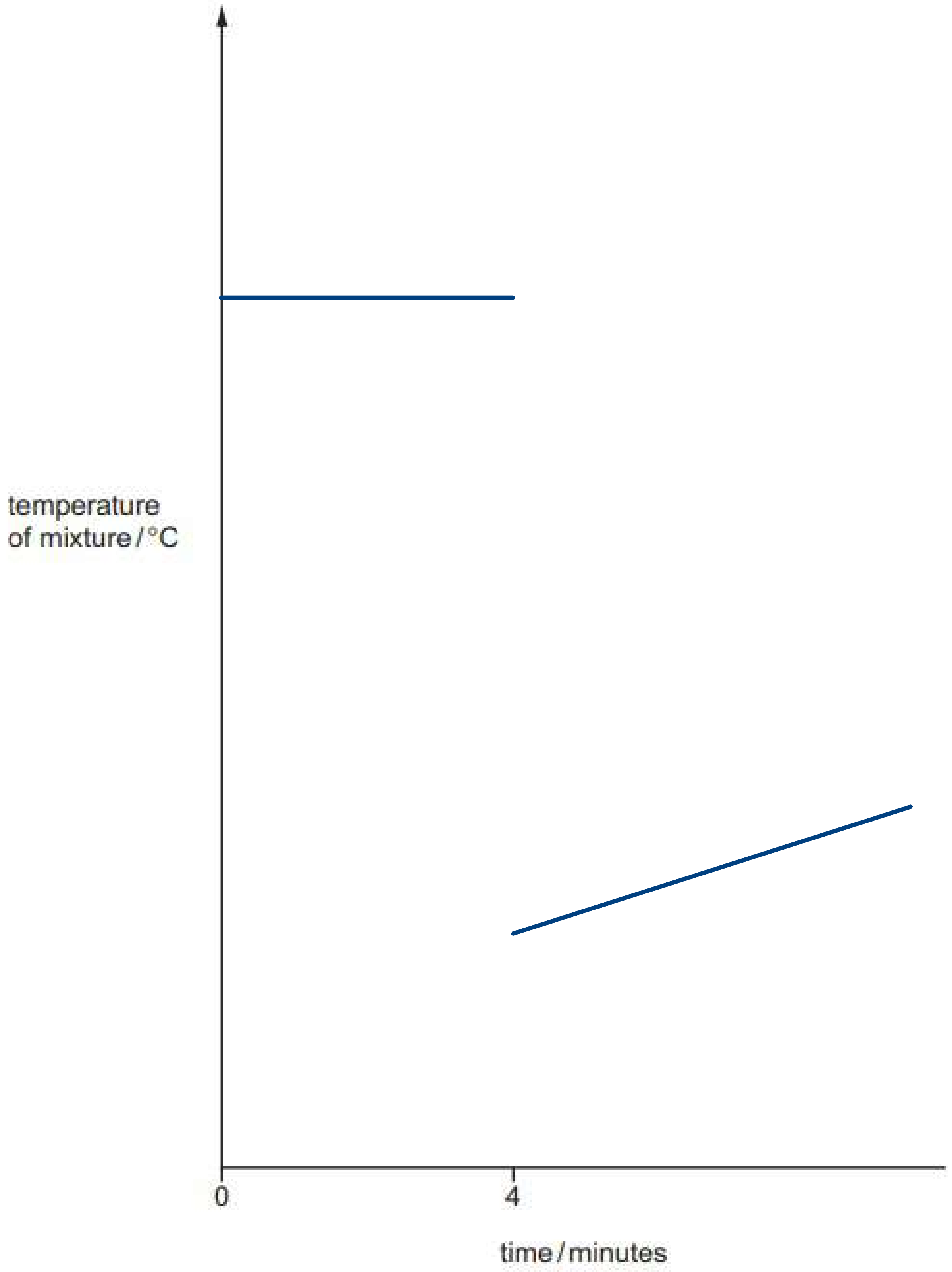
use more Na_2CO_3 so that a larger ΔT is recorded and the percentage uncertainty decreases [1]

or use smaller volume $\longrightarrow Q = mc\Delta T \Rightarrow \uparrow\Delta T = \frac{Q}{\downarrow mc} \rightarrow$ less % error

(h) The student used the results from Experiment 2 and correctly determined the enthalpy change for the reaction between $NaHCO_3(s)$ and $HCl(aq)$, ΔH_2, to be +24.2 kJ mol⁻¹.

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g) \qquad \Delta H_2 = +24.2\,kJ\,mol^{-1}$$

Use the axes to draw a sketch graph of the expected results of Experiment 2.

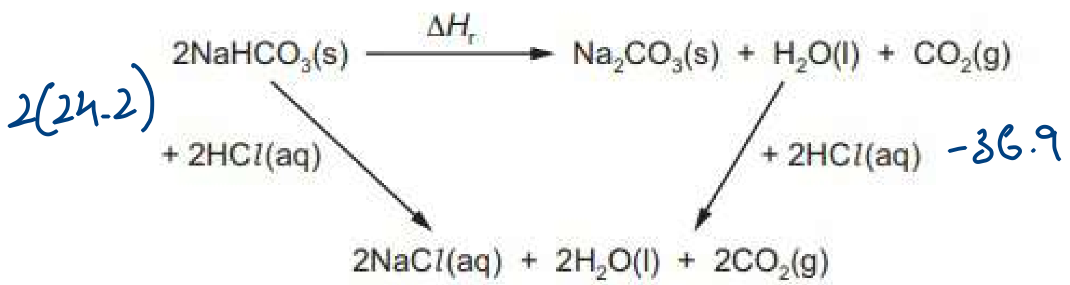


[2]

(i) Use ΔH_1 from **(d)** and ΔH_2 from **(h)** to determine the enthalpy change of reaction, ΔH_r, for the decomposition of $NaHCO_3(s)$.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

An energy cycle has been drawn for you.



If you were unable to calculate ΔH_1 in **(d)**, assume $\Delta H_1 = -26.7\ kJ\,mol^{-1}$. This is **not** the correct value of ΔH_1.

$\Delta H_r = 2(24.2) - (-36.9) = 85.3$

ΔH_r = 85.3 kJ mol^{-1} [2]

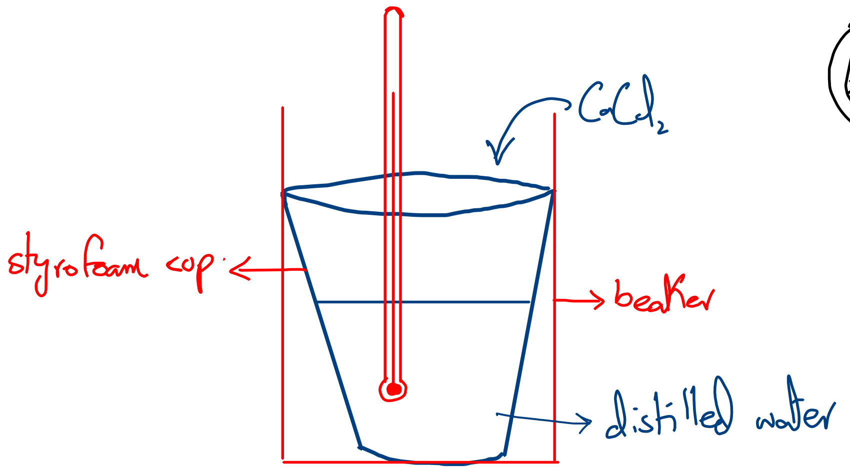
1 The pain of muscle strains and swellings can be eased by using heat packs. As a source of heat, some heat packs use the energy released when anhydrous calcium chloride dissolves in water.

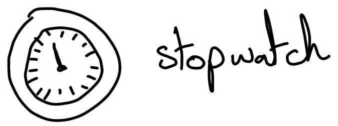
$$CaCl_2(s) \xrightarrow{\text{water}} Ca^{2+}(aq) + 2Cl^-(aq)$$

A heat pack consists of a bag of water, inside which a smaller bag contains anhydrous calcium chloride. When pressure is applied to the heat pack, the smaller bag bursts releasing the anhydrous calcium chloride into the water. The heat pack is shaken to speed up dissolving. Energy is released which warms the heat pack.

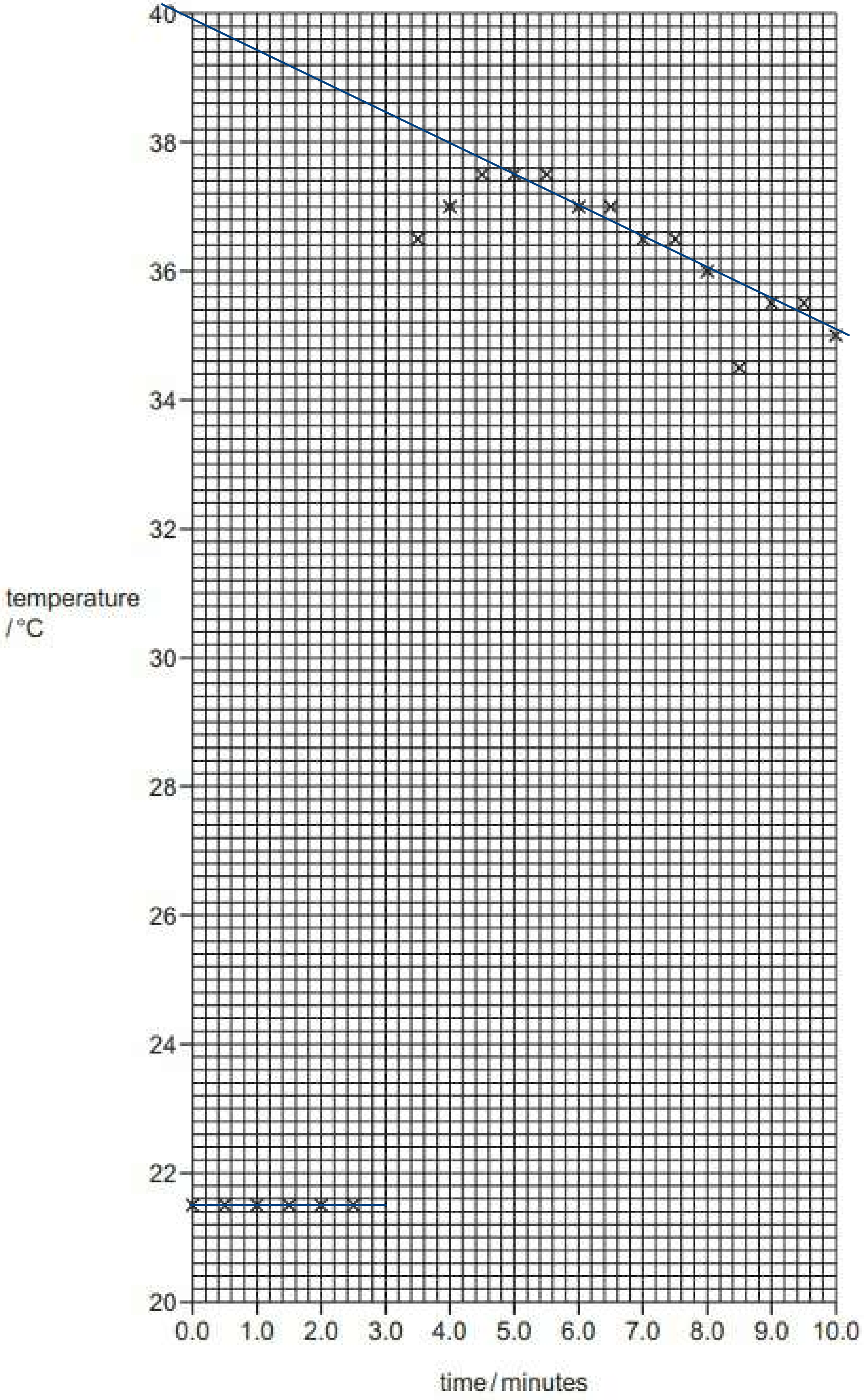
A student carried out an experiment to determine the enthalpy change when anhydrous calcium chloride dissolves in distilled water. The results the student obtained are plotted on the graph on page 4.

(a) By considering the graph of results, draw a labelled diagram of the experimental set-up that the student could have used to produce the graph shown.
Label the apparatus and chemicals required to measure the **two** variables.





[2]



(b) Explain why the student took readings between 0.0 minutes and 2.5 minutes.

to ensure the water reached a constant room temperature

[1]

(c) Explain why the student did not take a reading at 3.0 minutes.

Calcium chloride was added at 3 minutes

[1]

(d) Explain why the temperature continued to increase between 3.5 minutes and 4.5 minutes.

reaction was not complete

[1]

(e) Draw two straight lines of best fit on the grid. Extrapolate these lines to estimate the theoretical temperature rise at 3.0 minutes. Give your answer to **one** decimal place.

$38.5 - 21.5 =$

theoretical temperature rise at 3.0 minutes = 17 °C [2]

(f) One of the results is anomalous. This occurred because the student took the thermometer out of the solution and then replaced it just before the reading was taken.

The time at which the anomalous reading was taken was 8.5 minutes.

Explain why these actions led to the anomalous point.

The thermometer takes time to reach the temperature of the solution

[1]

(g) Explain why stirring the mixture would make this experiment more reliable.

The $CaCl_2$ would dissolve faster and heat distribution would be uniform.

[1]

(h) Anhydrous calcium chloride is classified as a moderate health hazard. It is an irritant.

Apart from wearing eye protection, state **one** other relevant safety precaution the student should have taken.

Wear gloves

[1]

(i) The student found the value for the enthalpy change of solution of anhydrous calcium chloride to be $-82.5\,\text{kJ mol}^{-1}$.
A manufacturer produces a heat pack that contains 75.0 g of water.

Calculate the mass of anhydrous calcium chloride the manufacturer must use in the inner bag to produce a rise in temperature of 30.0 °C.
The specific heat capacity of water, $c = 4.18\,\text{J g}^{-1}\text{K}^{-1}$.
[A_r: Ca, 40.1; C*l*, 35.5]

$Q = mc\Delta T$

$Q = 75 \times 4.18 \times 30 = 9405\ \text{J}$

$\text{moles used} = \dfrac{9405 \times 10^{-3}}{82.5} = 0.114\ \text{mol}$

$\text{mass} = \text{mol} \times Mr/Ar = 0.114 \times (40.1 + 71) = 12.7\text{g}$

mass of anhydrous $CaCl_2$ = 12.7 g [2]

[Total: 12]

2 Naphthalene, $C_{10}H_8$, is an aromatic organic compound that is a solid at room temperature and pressure.

The enthalpy change of fusion, ΔH_{fusion}, of naphthalene is the enthalpy change that occurs when naphthalene melts. It cannot be directly measured.

$$C_{10}H_8(s) \rightleftharpoons C_{10}H_8(l)$$

The melting point of naphthalene depends on its purity. When other compounds dissolve in liquid naphthalene its melting point is lowered.

Diphenylamine, $(C_6H_5)_2NH$, dissolves in liquid naphthalene. By varying the amount of diphenylamine and measuring the melting points of the different mixtures, the enthalpy change of fusion of naphthalene can be calculated.

(a) The equation that links the melting point of a naphthalene-diphenylamine mixture with its enthalpy change of fusion, ΔH_{fusion}, is shown.

$$\log Y = A - \frac{\Delta H_{\text{fusion}}}{2.30 \times RT_m}$$

$$\underset{\overset{\downarrow}{y}}{\log Y} = \underset{\overset{\downarrow}{m}}{-\left(\frac{\Delta H}{2.30R}\right)} \underset{\overset{\downarrow}{x}}{\frac{1}{T_m}} \underset{\overset{\downarrow}{+c}}{+A}$$

Y = mole fraction of naphthalene

R = molar gas constant, 8.31 J K^{-1} mol^{-1}

T_m = melting point of naphthalene in K

A is a constant

(i) A series of experiments is performed using the same amount, 0.100 mol, of naphthalene each time.

Calculate the mass of naphthalene, $C_{10}H_8$, that should be used for each of these experiments.

[A_r: C, 12.0; H, 1.0]

mass = mol × Mr/Ar = 0.100 × (12(10) + 8)

= 12.8

mass of naphthalene = 12.8 g [1]

(ii) The melting point and freezing point of a substance are the same. The melting point, T_m, of a substance can be found by recording the temperature at which the substance freezes, measured when crystals first start to appear on cooling.

The results of a series of experiments using 0.100 mol of naphthalene and different masses of diphenylamine are shown.

Process the results to complete the table.

Record all your data to **three** significant figures.

The mole fraction of naphthalene, Y, is calculated as shown.

$$Y = \frac{n_N}{n_N + n_D}$$

n_N = amount in moles of naphthalene, $C_{10}H_8$, = 0.100

n_D = amount in moles of diphenylamine, $(C_6H_5)_2NH$

amount of $(C_6H_5)_2NH$, n_D / mol	mole fraction of $C_{10}H_8$, Y	temperature at which crystals appear, T_m / K	$\frac{1}{T_m}$ / 10^{-3} K^{-1}	log Y
0.00	1.00	353	2.83	0.00
0.00888	0.918	349	2.87	-0.0372
0.0178	0.849	345	2.90	-0.0711
0.0266	0.790	341	2.93	-0.102
0.0355	0.738	338	2.96	-0.132
0.0444	0.693	334	2.99	-0.159
0.0533	0.652	331	3.02	-0.186
0.0621	0.617	329	3.04	-0.210
0.0769	0.565	325	3.08	-0.248

[3]

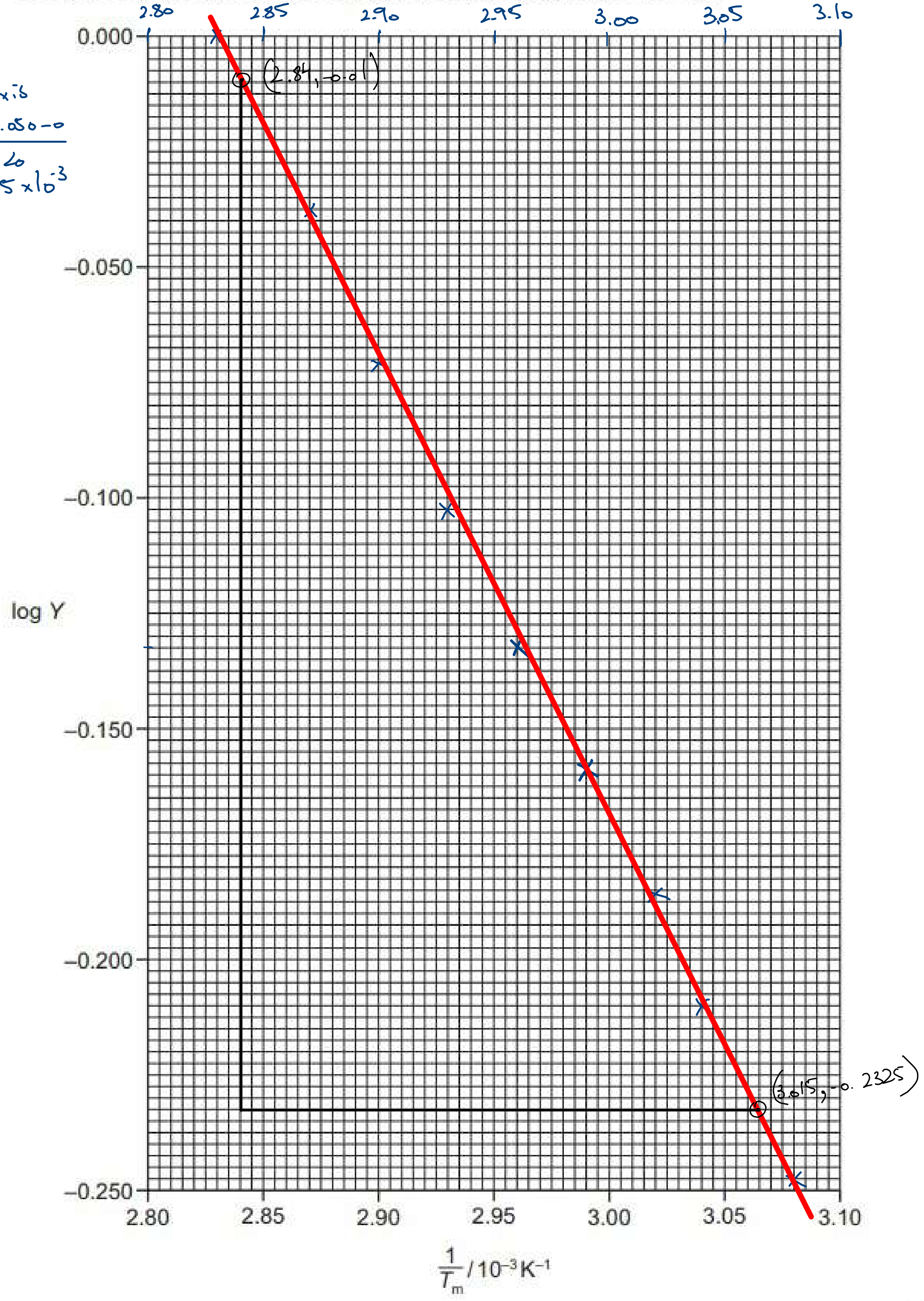
(b) Plot a graph on the grid to show the relationship between $\frac{1}{T_m}$ and log Y.

Use a cross (×) to plot each data point. Draw the straight line of best fit.

y axis

$= \frac{-0.050 - 0}{20}$

$= 2.5 \times 10^{-3}$



[2]

(c) (i) Use the graph to determine the gradient of the line of best fit. State the co-ordinates of both points you used in your calculation.

co-ordinates 1 (2.84, −0.01) co-ordinates 2 (3.015, −0.2325)

$$\frac{-0.2325 + 0.01}{(3.015 - 2.84) \times 10^{-3}} = -1271$$

gradient = −1270 K [2]

(ii) Use your answer to **(c)(i)** to determine the value of the enthalpy change of fusion of naphthalene, ΔH_{fusion}, in kJ mol⁻¹.

$$-1270 = \frac{-\Delta H}{2.30 \times \frac{8.31}{1000}} \Rightarrow \Delta H = 24.3$$

ΔH_{fusion} = 24.3 kJ mol⁻¹ [1]

(d) (i) Do you consider the results obtained to be reliable? Explain your answer.

Yes as all points are very close or on the line

[1]

(ii) Different literature values for the enthalpy change of fusion of naphthalene suggest that 10.00 g of naphthalene require between 1.45 kJ and 1.47 kJ to melt.

Use this information to calculate the range of ΔH_{fusion} values of naphthalene, $C_{10}H_8$, given in literature.

Use your values to comment on the accuracy of the experimental procedure.
[A_r: C, 12.0; H, 1.0]

> If you were not able to calculate ΔH_{fusion} in **(c)(ii)**, you may use 18.4 kJ mol^{-1}, but this may not be the correct answer.

$n = \frac{mass}{Mr/Ar} = \frac{10}{12(10)+8} = \frac{5}{64}$

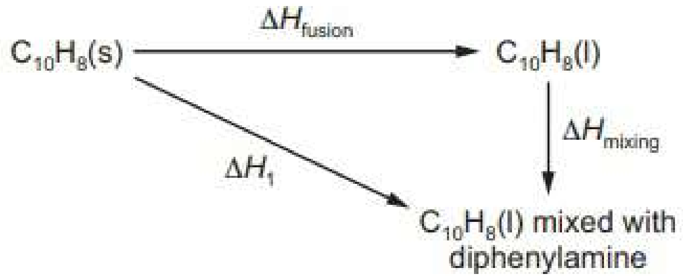
$\frac{1.45}{5/64} = 18.56 \approx 18.6$ $\frac{1.47}{5/64} = 18816 \approx 18.8$

range 18.6 – 18.8 kJ mol^{-1}

comment experiment not very accurate as the range does not contain actual value / is very different to actual value

[2]

(e) The enthalpy change calculated in this reaction is actually ΔH_1, shown in the Hess' cycle.



It is assumed that the enthalpy change when $C_{10}H_8(l)$ and diphenylamine are mixed, ΔH_{mixing}, is zero, and therefore $\Delta H_1 = \Delta H_{fusion}$.

State how the value of ΔH_1 compares to the value of ΔH_{fusion} if the mixing of naphthalene and diphenylamine is endothermic.

Explain your answer.

ΔH_1 would be more positive as ΔH_{mixing} is positive so $\Delta H_1 = \Delta H_{fusion} + \Delta H_{mixing}$ is positive

[2]

(f) A student incorrectly uses a value for the M_r of diphenylamine that is too low. This produces incorrect values for the mole fraction, Y.

(i) Predict the effect this will have on the calculated values of Y.

Explain your answer.

$$Y = \frac{n_N}{n_N + n_D}$$

$\uparrow n_D = \frac{mass}{M_r / Ar \downarrow}$

n_N = amount in moles of naphthalene
n_D = amount in moles of diphenylamine

n_D would be larger and so Y would be smaller

[1]

(ii) The student uses the incorrectly calculated value of Y from **(f)(i)** in the determination of ΔH_{fusion}.

$$\log Y = A - \frac{\Delta H_{fusion}}{2.30 \times RT_m}$$

Predict how the student's calculated value of ΔH_{fusion} is different from the actual value.

Explain your answer.

ΔH_{fusion} will be more positive as log Y would be more negative resulting in a gradient that is more negative

[2]

1 Zinc metal reacts with aqueous copper(II) sulfate.

$$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$$

The enthalpy change of this reaction, ΔH, can be determined by adding excess zinc powder to a measured volume of 0.500 mol dm^{-3} aqueous copper(II) sulfate.

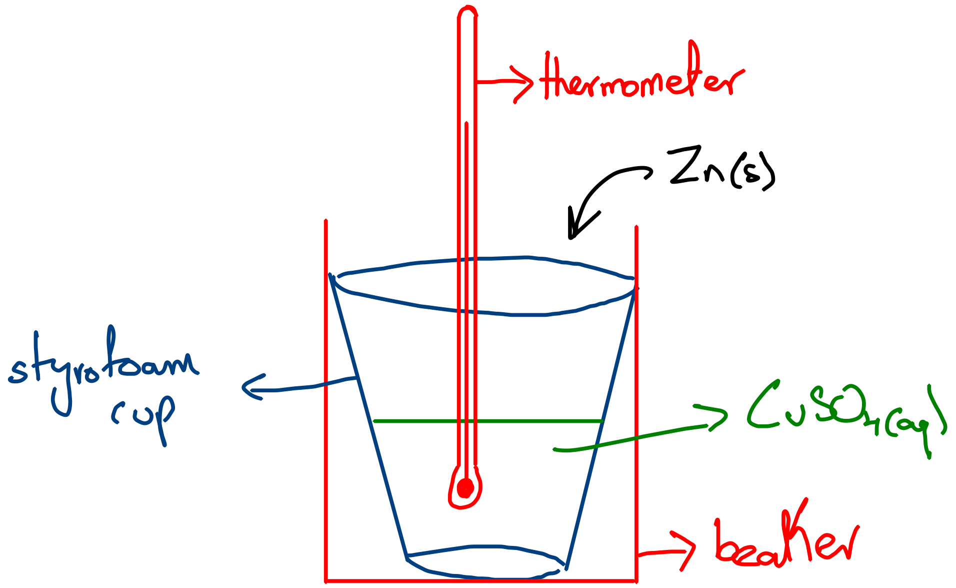
The temperature of 25.0 cm^3 of 0.500 mol dm^{-3} aqueous copper(II) sulfate is recorded for three minutes. At four minutes 3 g, an excess, of zinc powder is added and the mixture is continuously stirred. The temperature is recorded at times shown in the table.

time / min	0	1	2	3		$4\frac{1}{2}$	5	$5\frac{1}{2}$	6	$6\frac{1}{2}$	7	8	9	10
temperature / °C	18	19.5	19.5	19.5		32.5	38	36	34	33	32.5	31.5	31	31

(a) Use the results table to deduce the graduations on the thermometer that is used to record these temperature readings.

1°C [1]

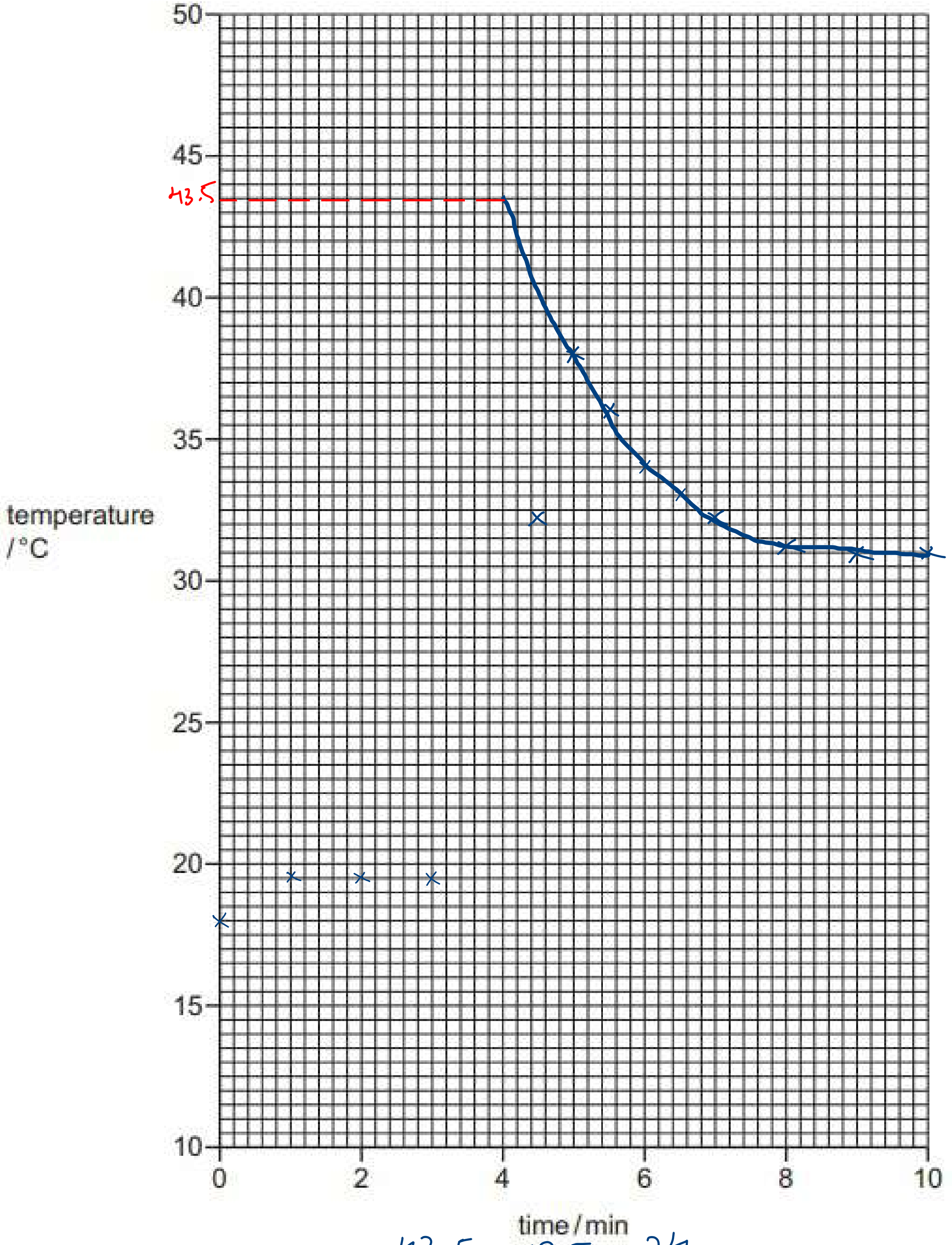
(b) Draw a labelled diagram of the apparatus set up at one minute.



[2]

(c) Plot a graph of temperature (*y*-axis) against time (*x*-axis). Use a cross (**×**) to plot each data point. Draw a line of best fit during cooling.

Extrapolate the cooling curve back to four minutes and determine the temperature change during the reaction.



43.5 − 19.5 = 24

temperature change = 24 °C [2]

(d) Use the formula $\Delta H = -mc\Delta T$ to determine the enthalpy change of reaction, ΔH, in $kJ\,mol^{-1}$.

Assume:

- mass of $1.00\,cm^3$ of solution = 1.00 g
- $c = 4.18\,J\,g^{-1}\,K^{-1}$.

$\Delta H = -mc\Delta T$
$= -25 \times 4.18 \times 24$
$= -2508$

$n = c \times V$
$= 0.5 \times 25 \times 10^{-3}$
$= 0.0125$

$\Rightarrow \dfrac{-2568 \times 10^{-3}}{0.0125} = -200$

ΔH = ...-200.... $kJ\,mol^{-1}$ [2]

(e) Heat loss is a major source of error in the results of this experiment.

Suggest how the following changes would affect the amount of heat loss, if at all.

Explain your answer in each case.

(i) The mass of zinc is doubled.

effect on heat loss ....no change....

explanation ....Zinc is already in excess....

.... [1]

(ii) The concentration of 25.0 cm^3 of aqueous copper(II) sulfate is doubled. The amount of zinc used is still an excess.

effect on heat loss ....increases....

explanation ....double the moles of $CuSO_4$ are available so more

....energy is released.... [1]

(iii) The volume of 0.500 $mol\,dm^{-3}$ aqueous copper(II) sulfate is doubled. The amount of zinc used is still an excess.

effect on heat loss ....no considerable change....

explanation ....The number of moles have doubled but twice the

....volume has to be heated so they cancel each other out.... [1]

[Total: 10]

1 A student is asked to find the enthalpy change for the reaction between anhydrous magnesium sulfate and water.

$$MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4{\bullet}7H_2O(s)$$

7 bonds made → exothermic

This enthalpy change cannot be measured directly.

(a) Predict whether the enthalpy change for this reaction is positive or negative. Explain the reason for your prediction.

prediction exothermic

explanation 7 bonds were made

[1]

(b) The student decided to do two separate experiments.

Experiment 1
To find the enthalpy change of solution of anhydrous magnesium sulfate, $MgSO_4(s)$, 0.0250 moles of $MgSO_4(s)$ are dissolved in 50.0 cm³ distilled water.

Experiment 2
To find the enthalpy change of solution of hydrated magnesium sulfate, $MgSO_4{\bullet}7H_2O(s)$, 0.0250 moles of $MgSO_4{\bullet}7H_2O(s)$ are dissolved in 50.0 cm³ distilled water.

The results for **Experiment 1** are shown in the graph of temperature against time on page 3.

(i) Draw and extrapolate the cooling curve back to 180 seconds. Determine the temperature change during the reaction.

27.55 − 19.5 =

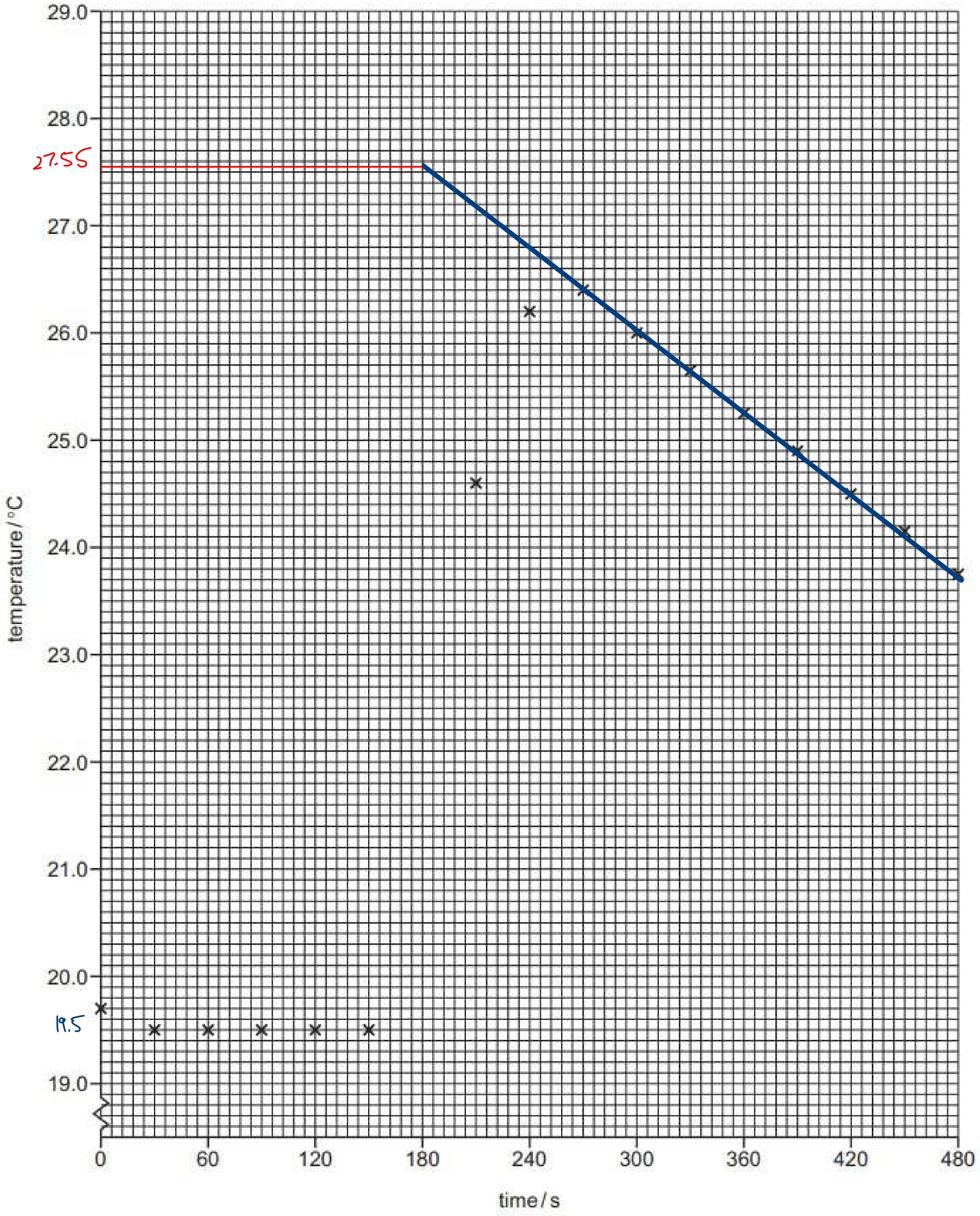
temperature change = 8.05 °C [1]

(ii) The anhydrous magnesium sulfate was not added when the timing started.

Explain why.

so that water reaches a constant room temperature

[1]



(iii) 3.01 g (0.0250 mol) of anhydrous magnesium sulfate is weighed.

Outline the next steps that should be taken in order to obtain the results in **Experiment 1**.

Write your answer using a series of numbered steps.

1) Measure 50 cm³ of distilled water using a measuring cylinder and transfer it into a styrofoam cup.

2) Place the styrofoam cup in a beaker and place a thermometer in the water

3) Measure the temperature of the water at 30s intervals until temperature becomes constant

4) Add $MgSO_4$ at 3 minutes and stir mixture at intervals [4]

(c) (i) The student realised that when dissolving 0.0250 moles $MgSO_4{\cdot}7H_2O(s)$, the amount of water present in the compound alters the total volume of water used in **Experiment 2**.

Calculate the volume of distilled water needed to make the total volume of water 50.00 cm³ in **Experiment 2**.

Give your answer to the nearest 0.05 cm³.

Assume that 1.00 cm³ of distilled water has a mass of 1.00 g.

[A_r: O, 16.0; H, 1.0]

$0.0250 \times 7 = 0.175$ mol of water added

mass $= \text{mol} \times M_r / A_r \Rightarrow 0.175 \times 18 = 3.15\text{g} \rightarrow 3.15\text{ cm}^3$

$50 - 3.15 = 46.85$

volume of distilled water = 46.85 cm³ [1]

(ii) State which piece of apparatus should be used to measure the volume of distilled water in **(c)(i)**.

Explain your answer.

apparatus burette

explanation it measures to 0.05cm³

[2]

(d) The temperature change when 0.0250 moles of $MgSO_4{\cdot}7H_2O(s)$ is added to the water is very small.

Suggest how the student should modify the experimental procedure to make the temperature change larger.

use a larger amount of moles of hydrated $MgSO_4$ [1]

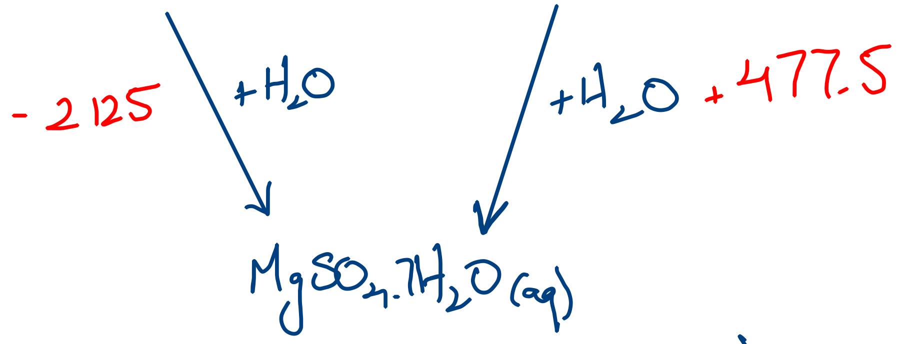
use a smaller volume of distilled water

(e) (i) The energy released by 0.0250 moles of $MgSO_4(s)$, in **Experiment 1**, is 2125 J. The energy absorbed by 0.0250 moles of $MgSO_4{\cdot}7H_2O(s)$, in **Experiment 2**, is 477.5 J.

Calculate the enthalpy change, Δ*H*, for the reaction.

Include a sign in your answer. Give your answer to **one** decimal place.

$$MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4{\cdot}7H_2O(s)$$



$$\Delta H = \frac{-2125}{0.0250} \times 10^{-3} - \left(\frac{477.5}{0.0250} \times 10^{-3}\right) = -104.1$$

enthalpy change, Δ*H* = -104.1 kJ mol⁻¹ [2]

(ii) The student noticed that some $MgSO_4(s)$ in **Experiment 1** was left undissolved.

State and explain the effect this would have on the value of the enthalpy change for the reaction in **Experiment 1**.

effect it would be less negative/exothermic

explanation fewer moles reacted.

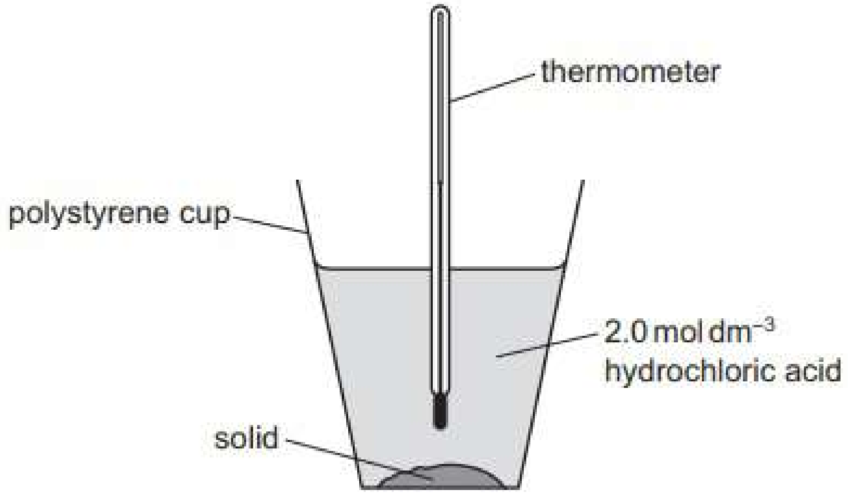
[1]

[Total: 14]

1 Potassium hydrogencarbonate, $KHCO_3$, decomposes when strongly heated to form potassium carbonate, K_2CO_3.

$$2KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(l)$$

A student plans to determine the value for the enthalpy change for this reaction, ΔH_r, which cannot be determined directly. The student carries out two separate experiments using the following apparatus.



Experiment 1 uses solid $KHCO_3$.

Experiment 2 uses solid K_2CO_3.

The following method is used for both experiments:

- Transfer 50.00 cm^3, an excess, of 2 mol dm^{-3} hydrochloric acid into a cup.
- After 2 minutes, record the temperature of the acid.
- Weigh approximately 0.0250 moles of solid.
- Add the solid to the acid, stir the mixture using a thermometer and record the temperature throughout the reaction.

Hazard information: 2 mol dm^{-3} hydrochloric acid is irritant, solid potassium hydrogencarbonate and solid potassium carbonate may cause irritation to the skin and eyes.

The equations for the two reactions are:

reaction 1 $KHCO_3(s) + HCl(aq) \rightarrow KCl(aq) + CO_2(g) + H_2O(l)$

reaction 2 $K_2CO_3(s) + 2HCl(aq) \rightarrow 2KCl(aq) + CO_2(g) + H_2O(l)$

(a) Suggest why it is not possible to measure ΔH_r for the decomposition reaction directly.

It is impossible to measure temperature change during direct heating [1]

(b) (i) Calculate the mass of 0.0250 moles of each solid. Give your answers to **three** decimal places.

[A_r: K, 39.1; H, 1.0; C, 12.0; O, 16.0]

mass = mol × M_r / A_r

$m_{KHCO_3} = 0.0250 \times (39.1 + 1 + 12 + 3(16)) = 2.5025$

$m_{K_2CO_3} = (0.0250) \times (2(39.1) + 12 + 3(16)) = 3.455$

mass of $KHCO_3$ = 2.503 g

mass of K_2CO_3 = 3.455 g

[1]

(ii) The masses of solid are measured using a electronic three decimal place balance. 0.001

Calculate the percentage error in the measurement of the mass of $KHCO_3$.

Show your working.

$\frac{0.001}{2.503} \times 100 = 0.0399$

percentage error = 0.040 [1]

(c) The student obtained the following results.

solid	initial temperature / °C	maximum / minimum temperature / °C	temperature change, ΔT / °C
$KHCO_3$	17.5	14.0	3.5
K_2CO_3	19.0	20.5	1.5

(i) Complete the table by calculating temperature change.

Use the formula $q = mc\Delta T$ to determine the energy change, q, that took place during **experiment 1**. Use q to calculate the enthalpy change of reaction 1, ΔH_1, in kJ mol^{-1}.

Include a sign in your answer.

Assume 1.00 cm^3 of solution has a mass of 1.00 g.
c = 4.18 J g^{-1} K^{-1}

$\Delta H_1 = \frac{50 \times 4.18 \times 3.5 \times 10^{-3}}{0.0250} = 29.26$

ΔH_1 = + 29.26 kJ mol^{-1} [2]

(ii) Use the formula $q = mc\Delta T$ to determine the energy change, q, that took place during **experiment 2**. Use q to calculate the enthalpy change of reaction 2, ΔH_2, in kJ mol^{-1}.

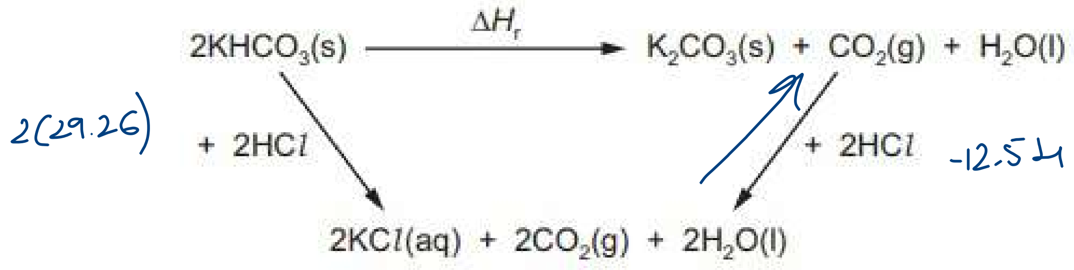
Include a sign in your answer.

Assume 1.00 cm^3 of solution has a mass of 1.00 g.
$c = 4.18\,J\,g^{-1}K^{-1}$

$$\Delta H_2 = \frac{50 \times 4.18 \times 1.5 \times 10^{-3}}{0.0250} = 12.54$$

ΔH_2 = – 12.54 kJ mol^{-1} [2]

(d) Use the following cycle to calculate ΔH_r.



$$2(29.26) - (-12.54) = +71.06$$

ΔH_r = +71.06 kJ mol^{-1} [2]

(e) A textbook states the value of the enthalpy change for the decomposition of potassium hydrogencarbonate as +76.0 kJ mol^{-1}.

Suggest **two** reasons why the experimental value is different to the actual value.

1 Heat losses

2 Amount of solid transferred may not be exactly 0.0250 mol

3. Non standard conditions [2]

(f) Suggest **one** improvement to the apparatus which would reduce the difference between the experimental value and the actual value.

Add a lid to the polystyrene cup [1]

(g) Name a suitable piece of apparatus which should be used to measure the volume of acid used in **experiment 1**.

burette [1]

(h) Apart from wearing safety glasses and a lab coat, state **one** safety precaution which must be taken during **experiment 1**. Explain your answer.

Wear gloves as chemicals are irritants.

[1]

[Total: 14]

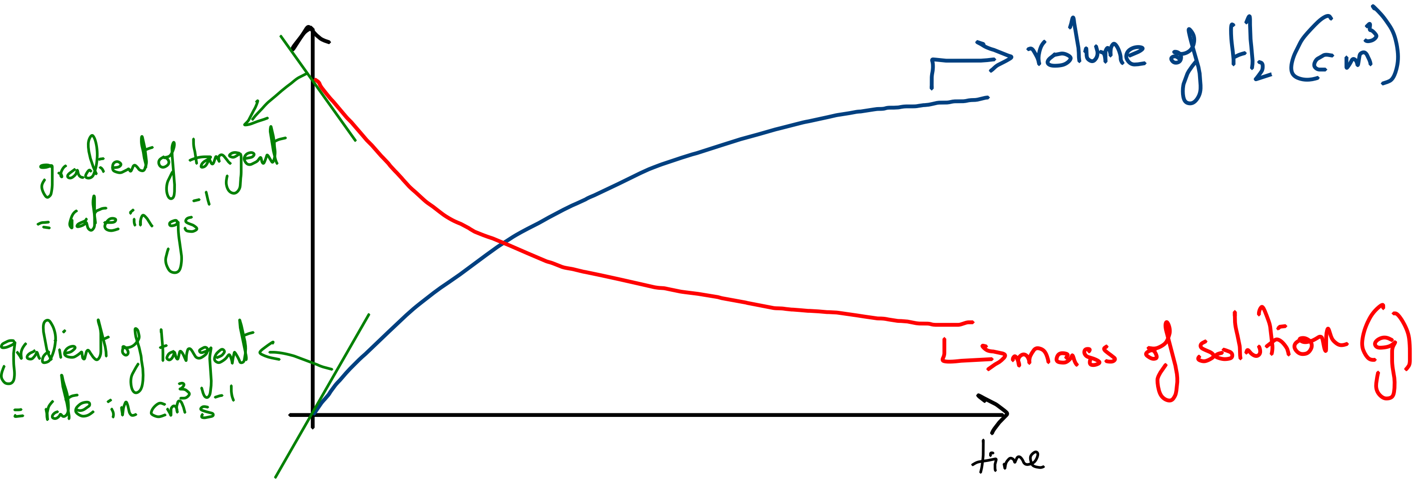
Rate of Reaction :- Revisit reaction Kinetics from A2

$$\text{Rate} = \frac{\text{change in concentration}}{\text{time}} \Rightarrow \text{Rate} \propto \frac{\text{change in mass/volume/conductivity/colour}}{\text{time}}$$

Q. $Ca_{(s)} + HCl_{(g)} \longrightarrow CaCl_{2(s)} + H_{2(g)}$

Describe 2 methods to calculate rate of reaction :-

1) Collect $H_{2}(g)$ in a syringe and record volumes at intervals
2) Measure the mass of the solution at intervals



* Always initial rates are compared

Rate equation :- $\text{rate} = K [A]^x [B]^y$

- if you want to find order w.r.t to A keep B in large excess
- Keep Temperature, Pressure, surface area constant as these affect K.

1 Propanone, CH_3COCH_3, is an organic liquid which is soluble in water.

Aqueous propanone reacts with aqueous iodine. The reaction is catalysed by $H^+(aq)$ ions.

$$CH_3COCH_3(aq) + I_2(aq) \rightarrow CH_3COCH_2I(aq) + HI(aq)$$

The order of reaction with respect to iodine can be determined experimentally.

An experiment is carried out using the following solutions.

- solution **A**, $25.0\,cm^3$ of $1.00\,mol\,dm^{-3}$ $CH_3COCH_3(aq)$
- solution **B**, $25.0\,cm^3$ of $1.00\,mol\,dm^{-3}$ $H_2SO_4(aq)$
- solution **C**, $50.0\,cm^3$ of $0.200\,mol\,dm^{-3}$ $I_2(aq)$

The solutions are mixed to start the reaction. At certain time intervals, a $10.0\,cm^3$ portion of the mixture is withdrawn and transferred to a conical flask containing excess sodium hydrogencarbonate, $NaHCO_3(aq)$. This prevents any further significant reaction taking place by removing the $H^+(aq)$ ions. The concentration of unreacted $I_2(aq)$ in each $10.0\,cm^3$ portion of the mixture can then be determined by titration with aqueous thiosulfate ions, $S_2O_3^{2-}(aq)$.

(a) State the size and type of apparatus needed to prepare a suitable volume of a standard solution of $1.00\,mol\,dm^{-3}$ $CH_3COCH_3(aq)$ from liquid propanone.

Calculate the mass of propanone needed to prepare this standard solution.
[A_r: C, 12.0; H, 1.0; O, 16.0]

apparatus volumetric flask, $250\,cm^3$

$n = c \times V = 1 \times 250 \times 10^{-3}$

$n = 0.25$ mol

mass = mol $\times M_r/A_r$ = $0.25 \times (3(12) + 6 + 16)$

$= 14.5g$

mass of propanone 14.5 g
[3]

(b) Solutions **A**, **B** and **C** need to be added in a specific order and the clock started as the third solution is added.

(i) Suggest the best order of adding the solutions.

1 A

2 B

3 C

[1]

(ii) Explain your choice.

A & C are the reactants so C should be mixed last so that reaction starts when timer starts [1]

(c) Each $10.0\,cm^3$ portion of mixture removed from the main reaction is added to a separate solution of sodium hydrogencarbonate, $NaHCO_3(aq)$, in a conical flask to remove $H^+(aq)$ ions.

(i) Which piece of apparatus should be used to transfer each $10.0\,cm^3$ portion of mixture to the conical flask?

$10.0\,cm^3$ pipette [1]

(ii) Suggest **two** reasons why $NaHCO_3(aq)$ is preferred to $NaOH(aq)$ as the reagent used to remove $H^+(aq)$ ions.

reason 1 $NaHCO_3$ will produce effervescence, when effervescence finishes we will know all H^+ ions are removed

reason 2 NaOH may react with I_2 and propanone to give iodoform reaction

[2]

(d) The unreacted iodine in each 10.0 cm³ portion of the mixture is titrated against 0.100 mol dm⁻³ aqueous thiosulfate ions, $S_2O_3^{2-}(aq)$, to determine the concentration of $I_2(aq)$ in the mixture at the time that the 10.0 cm³ portion was withdrawn.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

(i) A 10.0 cm³ portion of mixture is removed at time = 0. This is before any of the 0.200 mol dm⁻³ $I_2(aq)$ had reacted.

Calculate the volume of 0.100 mol dm⁻³ $S_2O_3^{2-}(aq)$ needed to react with the iodine present in this 10.0 cm³ portion of mixture.

n of $I_2 = c \times V = 0.200 \times 50 \times 10^{-3} = 0.01$ mol in 50cm³

n of I_2 in 10cm³ = $\begin{matrix} 100 & : & 0.01 \\ 10 & : & x \end{matrix}$ ⇒ $x = 1 \times 10^{-3}$

n of $S_2O_3^{2-} = 2 \times 1 \times 10^{-3} = 2 \times 10^{-3}$

$V = \frac{n}{c} = \frac{2 \times 10^{-3}}{0.100} = 0.02\ dm^3 \times 1000 = 20 cm^3$

volume 0.100 mol dm⁻³ $S_2O_3^{2-}(aq)$ = 20.0 cm³ [3]

(ii) Suggest the name of a suitable indicator to use in the titration and state its colour change.

indicator starch

colour change blue to colourless

[2]

(e) State **two** variables which must be recorded in this experiment.

For each variable, state the units.

variable 1 time units seconds

variable 2 volume of thiosulfate units cm³

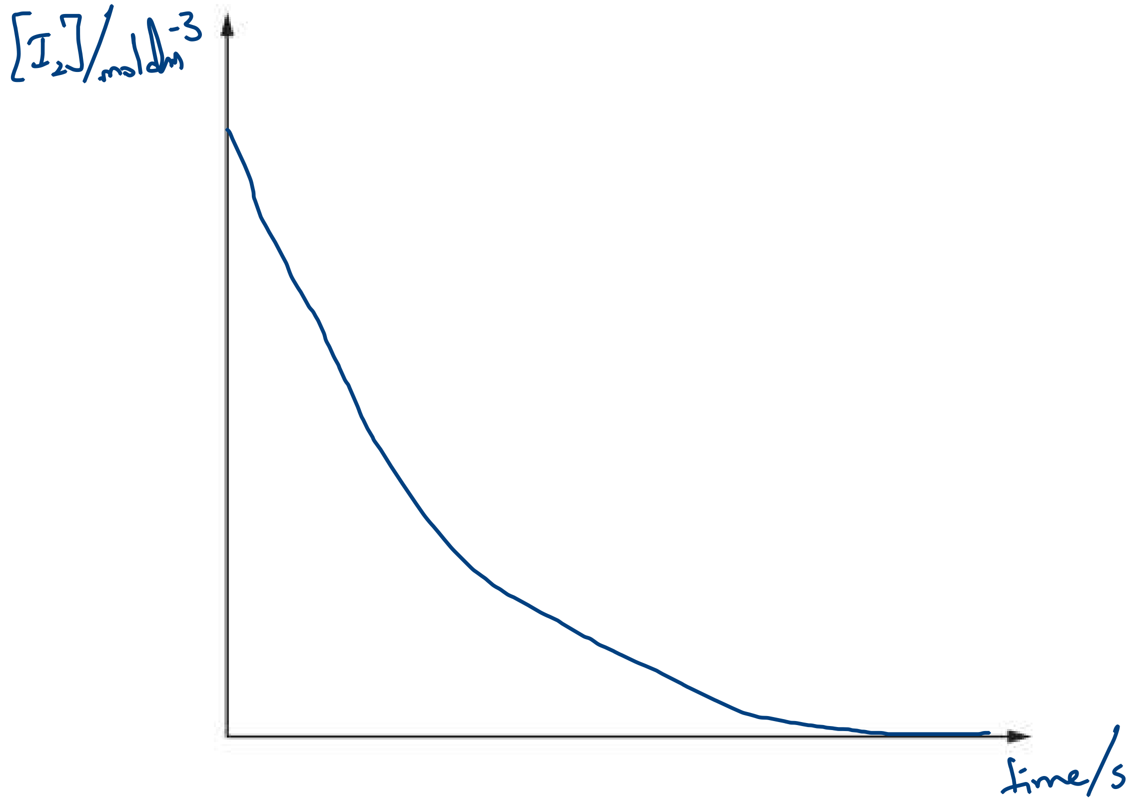
[2]

(f) State **one** other variable which must be controlled in this experiment.

temperature [1]

(g) The order of reaction with respect to iodine is expected to be first order.

(i) Use the axes below to draw a sketch graph of how the concentration of iodine changes during the experiment. Label both axes.



[2]

(ii) How could the graph be used to prove that the order of reaction with respect to iodine is first order?

Half life will be constant

calculate half lives for atleast 3 different points on the graph

[1]

(h) A student suggested that the temperature at which the experiment was carried out would affect the order of reaction with respect to iodine.

State if the student was correct and explain your answer.

incorrect. Temperature affects rate of reaction, it does not affect order of reaction

[1]

[Total: 20]

1 The reaction between hydrogen peroxide, H_2O_2(aq), and iodide ions, I^-(aq), takes place in acidic conditions.

$$H_2O_2(aq) + 2I^-(aq) + 2H^+(aq) \rightarrow I_2(aq) + 2H_2O(l) \qquad \textbf{reaction 1}$$

The rate of this reaction can be found by measuring the time taken for a given amount of iodine, I_2(aq), to form.

This is done by adding a known amount of thiosulfate ions, $S_2O_3^{2-}$(aq), and allowing the I_2(aq) formed in **reaction 1** to react with the $S_2O_3^{2-}$(aq).

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq) \qquad \textbf{reaction 2}$$

After the $S_2O_3^{2-}$(aq) ions have all reacted in **reaction 2**, any further I_2(aq) formed in **reaction 1** can be detected using an indicator.

A student carried out a series of experiments to determine the order of reaction with respect to the concentration of I^-(aq) ions in **reaction 1**.

The student prepared the following solutions.

solution A 0.100 mol dm^{-3} KI(aq)

solution B 0.0500 mol dm^{-3} $Na_2S_2O_3$(aq)

The student also had access to the following chemicals.

solution C 0.100 mol dm^{-3} H_2O_2(aq)
0.2 mol dm^{-3} H_2SO_4(aq)
distilled water
a suitable indicator

(a) The student prepared solution **A** in a 250 cm^3 volumetric flask.

(i) The student used a balance accurate to two decimal places and a weighing boat. A weighing boat is a small container used to hold solid samples when they are weighed.

Determine the mass, in g, of KI needed to prepare 250.0 cm^3 of solution **A**.
[A_r: K, 39.1; I, 126.9]

$n = c \times V = 0.1 \times 250 \times 10^{-3} = 0.025$ mol

mass = mol $\times Mr/Ar = 0.025 \times (39.1 + 126.9) = 4.15$g

mass = 4.15 g [2]

(ii) The student weighed the empty weighing boat. The student then added solid KI to the weighing boat until the mass of KI calculated in **(i)** was reached. The student transferred all of the KI from the weighing boat into a 100 cm^3 beaker.

Describe how the student could check that the mass of KI transferred into the 100 cm^3 beaker was exactly the same as the mass calculated in **(i)**.

reweigh the empty weighing boat and subtract the masses.
The difference will work out to be 4.15g [1]

(iii) The student dissolved the KI in the 100 cm^3 beaker in distilled water and transferred the solution formed into a 250 cm^3 volumetric flask. Distilled water was added to the volumetric flask until the volume of the solution was exactly 250 cm^3. Care was taken to avoid parallax errors.

Describe:

- how the student should transfer all the KI solution from the beaker into the 250 cm^3 volumetric flask
- how the student should fill the volumetric flask exactly up to the 250 cm^3 mark.

The student should stir the mixture to make sure all KI is dissolved and then transfer to volumetric flask. He should rinse the beaker and stirrer and add their washings to volumetric flask too. He should add distilled water and when nearing 250cm^3, should use a dropper to add distilled water dropwise. [2]

(b) The student rinsed a burette with solution **A** before filling it with solution **A**.

Explain why this improves the accuracy of the results.

The solution in the burette is at the expected concentration

[1]

(c) The student was given a solution of 0.400 mol dm^{-3} $Na_2S_2O_3$(aq).

Determine the volume, in cm^3, of this solution that should be added to a 100 cm^3 volumetric flask to prepare 100.0 cm^3 of solution **B**. Give your answer to **two** decimal places.

$n_1 = n_2$

$c_1 \times V_1 = c_2 \times V_2$

$0.400 \times V_1 = 0.05 \times 100 \times 10^{-3}$

$V_1 = 0.0125 \times 1000 = 12.5$ cm^3

volume = 12.50 cm^3 [1]

(d) Experiment 1 was carried out using a series of steps.

step 1 The student used a measuring cylinder to measure 25 cm³ of 0.2 mol dm⁻³ H_2SO_4(aq). This was transferred to a conical flask.

step 2 The student added 20.00 cm³ of distilled water from a burette to the conical flask.

0.1 mol dm^{-3} KI

step 3 The student added 5.00 cm³ of solution **A** from a burette to the conical flask.

0.05 mol dm^{-3} $Na_2S_2O_3$

step 4 The student added 5.00 cm³ of solution **B** from a burette to the conical flask.

step 5 The student added 1.0 cm³ of indicator from a teat pipette to the conical flask.

step 6 The student used a burette to add 10.00 cm³ of solution **C** to a small beaker. The contents of the beaker were added to the conical flask and a stopclock was started immediately. The stopclock was stopped when the I_2 formed caused the indicator to change colour.

In Experiments 2–6 the student repeated **steps 1–6** but using the volumes of distilled water and solution **A** given in the table.

The student carried out two trials of each experiment.

experiment	volume of H_2SO_4(aq) / cm³	volume of distilled water / cm³	volume of solution **A**, *v* / cm³	volume of solution **B** / cm³	volume of indicator / cm³	time for the indicator to change colour, *t* / s	
						trial 1	trial 2
1	25.0	20.00	5.00	5.00	1.0	218	220
2	25.0	15.00	10.00	5.00	1.0	112	113
3	25.0	12.50	12.50	5.00	1.0	100	▨
4	25.0	10.00	15.00	5.00	1.0	77	76
5	25.0	5.00	20.00	5.00	1.0	59	59
6	25.0	0.00	25.00	5.00	1.0	47	49

(i) In Experiment 3, trial 2, the indicator changed colour as soon as the student added solution **C** to the conical flask. No results were recorded for Experiment 3, trial 2.

Suggest which step the student did **not** carry out in Experiment 3, trial 2.

He forgot to add $Na_2S_2O_3$ - step 4 [1]

(ii) Suggest why the results shown in the table could be considered reliable.

The repeated times are close to one another

[1]

(iii) What was the percentage error in the burette reading for measuring the volume of solution **A** in Experiment 5?

$$\frac{(2 \times 0.05)}{20.00} \times 100 = 0.5$$

percentage error = 0.500 % [1]

(iv) Suggest why a measuring cylinder was used to measure the volume of H_2SO_4(aq) rather than a more accurate piece of apparatus, such as a burette.

Acid is in excess [1]

(v) For Experiments 1–6, state:

- the independent variable volume of solution A
- the dependent variable. time

[2]

(e) The rate equation can be written as rate = $k[I^-]^n$ where $[I^-]$ is proportional to the volume of solution **A** and *n* is the order of reaction with respect to I^-.

(i) Use only the results of Experiments 1–6 given in **(d)** to complete the table where:

- *v* is the volume of solution **A** used in cm^3
- t_{av} is the average time taken in trial 1 and trial 2 in s.

Give all values to **three** significant figures.

trial 1	trial 2
218	220
112	113
100	
77	76
59	59
47	49

experiment	v / cm^3	log v	t_{av} / s	$(1/t_{av})$ / s^{-1}	$\log(1/t_{av})$
1	5.00	0.699	219	0.00457	-2.34
2	10.00	1.00	113	0.00885	-2.05
3	12.50	1.10	100	0.0100	-2.00
4	15.00	1.18	76.5	0.0131	-1.88
5	20.00	1.30	59.0	0.0169	-1.77
6	25.00	1.40	48.0	0.0208	-1.68

[2]

(ii) Rate can be expressed as $(1/t_{av})$.
The rate equation can be expressed as shown.

$$\log(1/t_{av}) = n\log v + c$$

$$\overset{\downarrow}{y} = \overset{\downarrow}{m}\overset{\downarrow}{x} + \overset{\downarrow}{c}$$

where:

- c is a constant
- *v* is proportional to $[I^-]$.

On the grid:

- Plot a graph of $\log(1/t_{av})$ against log *v*. Use a cross (×) to plot each data point.
- Draw a line of best fit.

[2]

(iii) Use your graph to determine the gradient of the line of best fit. State the coordinates of both points you used in your calculation. Give the gradient to **three** significant figures.
Determine the order of reaction with respect to I^-(aq).

co-ordinates 1 (0.74, -2.30) co-ordinates 2 (1.27, -1.80)

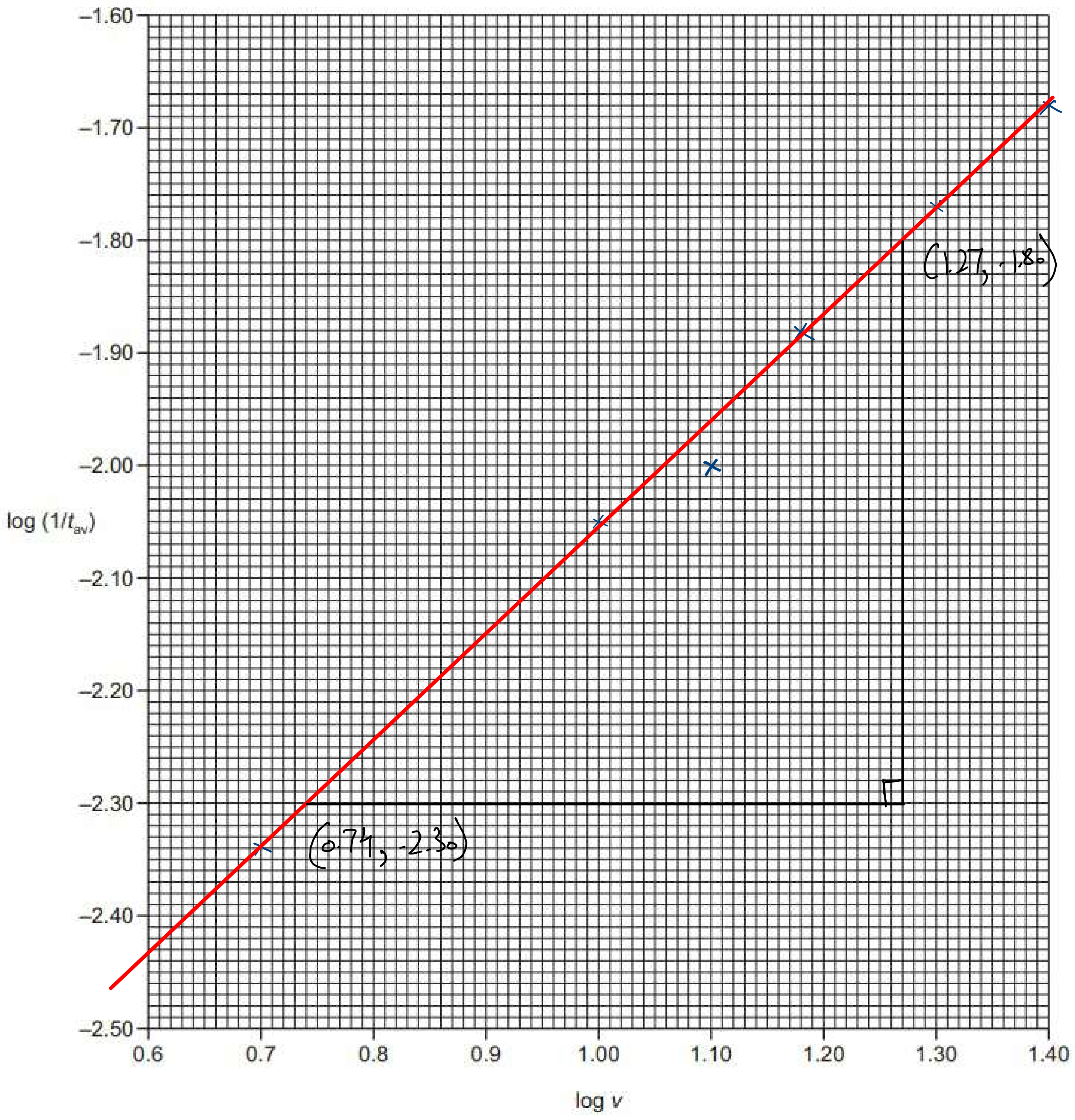
$$\frac{-1.80 + 2.30}{1.27 - 0.74} = 0.943$$

n = gradient ≃ 1

gradient = 0.943

order of reaction with respect to I^-(aq) = 1st

[3]



[Total: 20]

1 Hydrogen peroxide decomposes slowly at room temperature to give water and oxygen.

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

The **initial** rate of this reaction can be increased by the addition of a metal oxide catalyst.

A student is asked to investigate which metal oxide catalyst is best at increasing the **initial** rate of this reaction by using a method which involves the collection of oxygen.

The student is provided with the following metal oxides: copper(II) oxide, iron(III) oxide, manganese(IV) oxide, nickel(II) oxide and titanium(IV) oxide.

The student is also provided with an excess volume, of a known concentration, of aqueous hydrogen peroxide and any laboratory equipment needed.

(a) (i) State the independent variable.

type of metal oxide [1]

(ii) State the dependent variable.

volume of O_2 gas produced in a certain time [1]

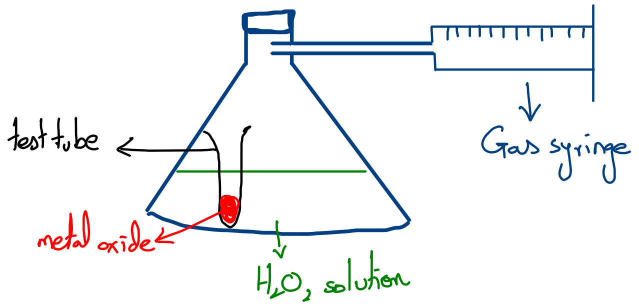
(b) State two variables that would need to be controlled.

1 the volume of H_2O_2 solution

2 Temperature

3. the amount and S.A of catalyst [2]

(c) Draw a labelled diagram of the assembled apparatus that could be used to carry out these experiments. The apparatus should allow the accurate recording of the oxygen produced.



[3]

(d) (i) What measurements need to be recorded during the course of each experiment to allow the **initial** rate to be determined?

volume of oxygen and time [1]

(ii) How is the **initial** rate determined using these measurements?

plot a volume time graph. Draw a tangent at t=0 and determine it's gradient [1]

(e) How can the student ensure that the results are reliable?

Repeat and average [1]

(f) Suggest an alternative method to investigate these reactions which does not include the collection of gas.

measuring the decrease in mass of the reaction vessel and time [1]

(g) Once the reaction has finished, how can the student demonstrate that the metal oxide has not been affected by the reaction?

filter the solid residue and measure its mass on a balance It will be same as mass before being added demonstrating that the catalyst has not been affected by the reaction [2]

(h) When aqueous hydrogen peroxide is stored there is a small hole in the lid of the bottle.

Suggest why this is necessary.

To allow oxygen gas being produced to move out and not build too much pressure inside. [1]

[Total: 14]

2 Benzenediazonium chloride, $C_6H_5N_2Cl$, is readily hydrolysed at temperatures above 5 °C, forming phenol, nitrogen gas and hydrochloric acid.

$$C_6H_5N_2Cl(aq) + H_2O(l) \rightarrow C_6H_5OH(aq) + N_2(g) + HCl(aq)$$

The progress of the reaction can be monitored by measuring the volume of gas produced over time. The volume of gas produced, V, after time, t, is proportional to the amount of benzenediazonium chloride that has been hydrolysed. The final volume of gas produced, V_{final}, is proportional to the original concentration of benzenediazonium chloride.

The order of reaction can be determined from these results.

(a) **(i)** The experimentally determined volumes of gas produced during the hydrolysis of benzenediazonium chloride at 50 °C are recorded below.

Process the results to allow you to plot a graph of $(V_{final} - V)$ against time, t.

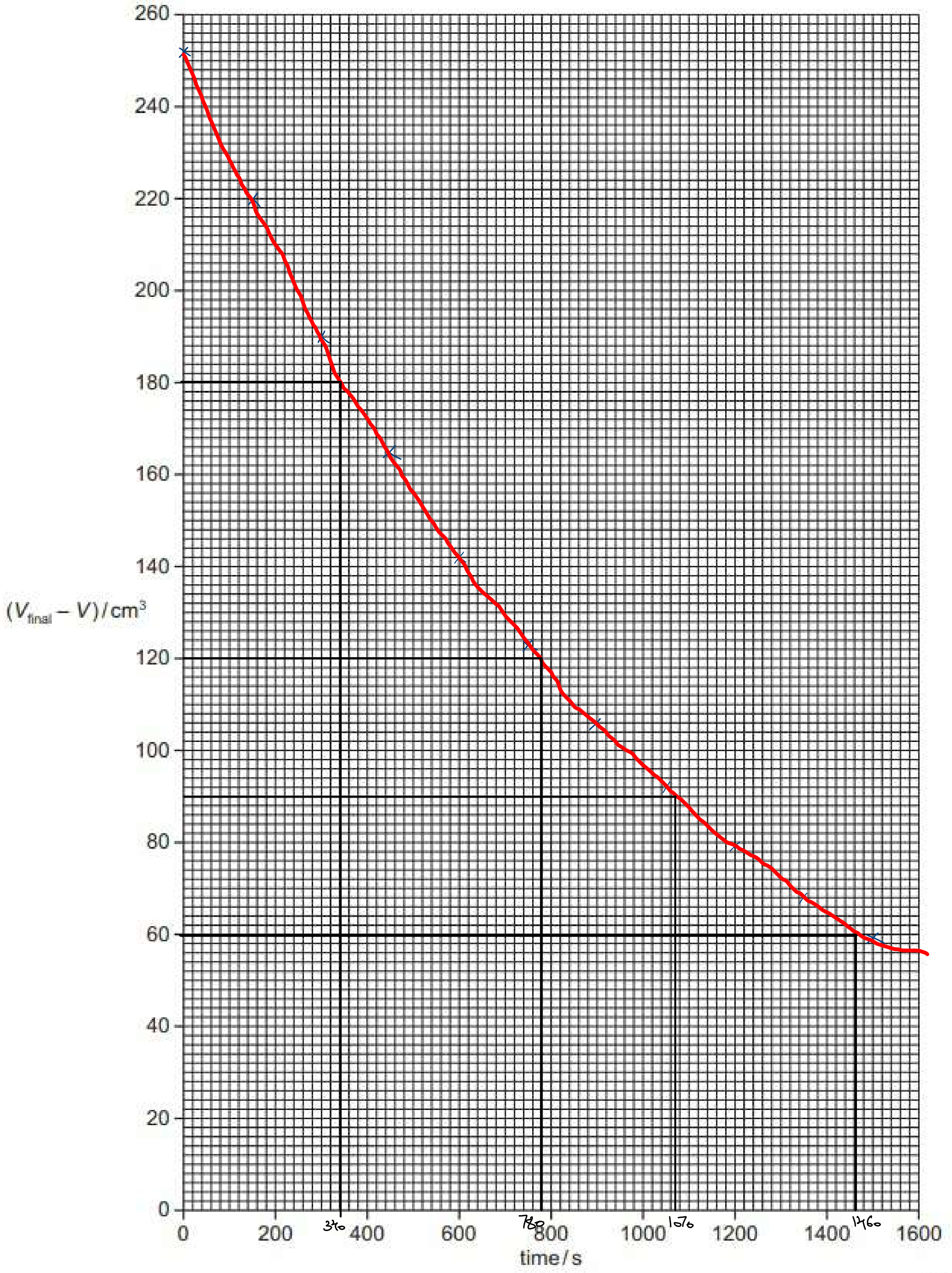
$V_{final} = 252\,cm^3$

time / s	volume, V / cm^3	$(V_{final} - V)$ / cm^3
0	0	252
150	32	220
300	62	190
450	87	165
600	110	142
750	129	123
900	146	106
1050	160	92
1200	173	79
1350	184	68
1500	193	59

[1]

(ii) Plot a graph to show how $(V_{final} - V)$ / cm^3 varies with time / s.

Use a cross (**×**) to plot each data point. Draw the curve of best fit.



[2]

(iii) Do you think the results obtained in **(i)** are reliable? Explain your answer.

Yes as most points are on the line

[1]

(iv) Use the graph to determine the half-life, $t_{½}$, of this reaction.

State the co-ordinates of both points you used in your calculation.

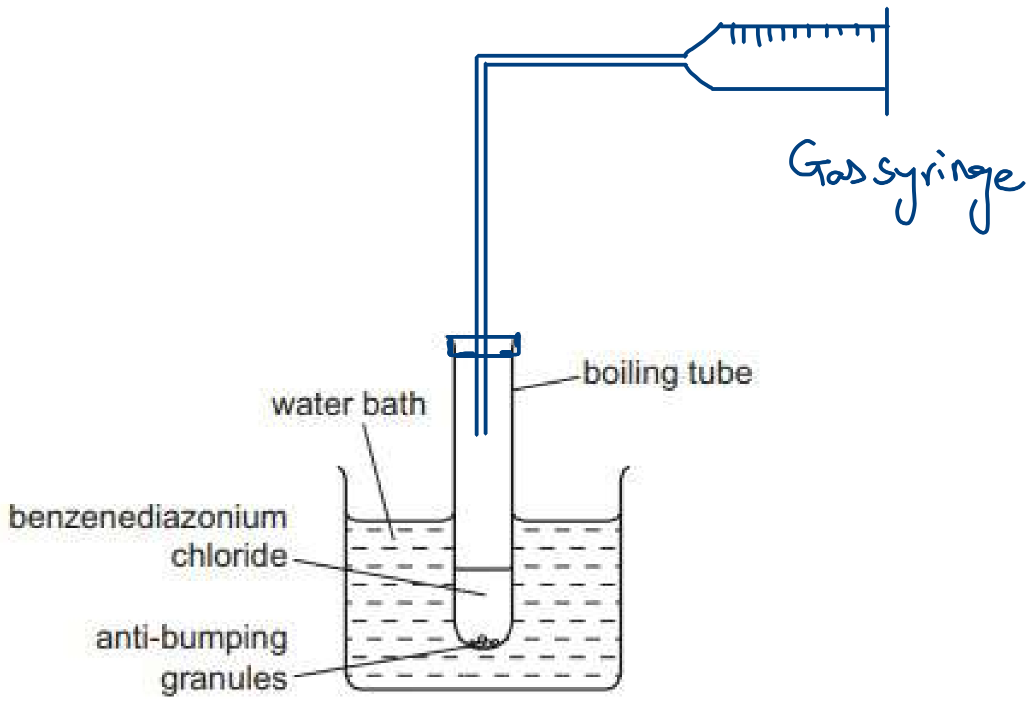
co-ordinates 1 (340, 180) co-ordinates 2 (1070, 90)

1070 - 340 =

half-life = 730 s

[2]

(b) A student set up an experiment to determine the order of the reaction in **(a)**. Part of the experimental set-up is shown below.



(i) Complete the diagram above to show the experimental set-up the student could have used to collect and measure the volume of gas evolved by the reaction. [2]

(ii) The water bath was set at 60 °C.

At a reaction temperature of 60 °C, the measurements made would be less accurate than measurements made at room temperature.

State why the measurements made at a higher temperature are less accurate.

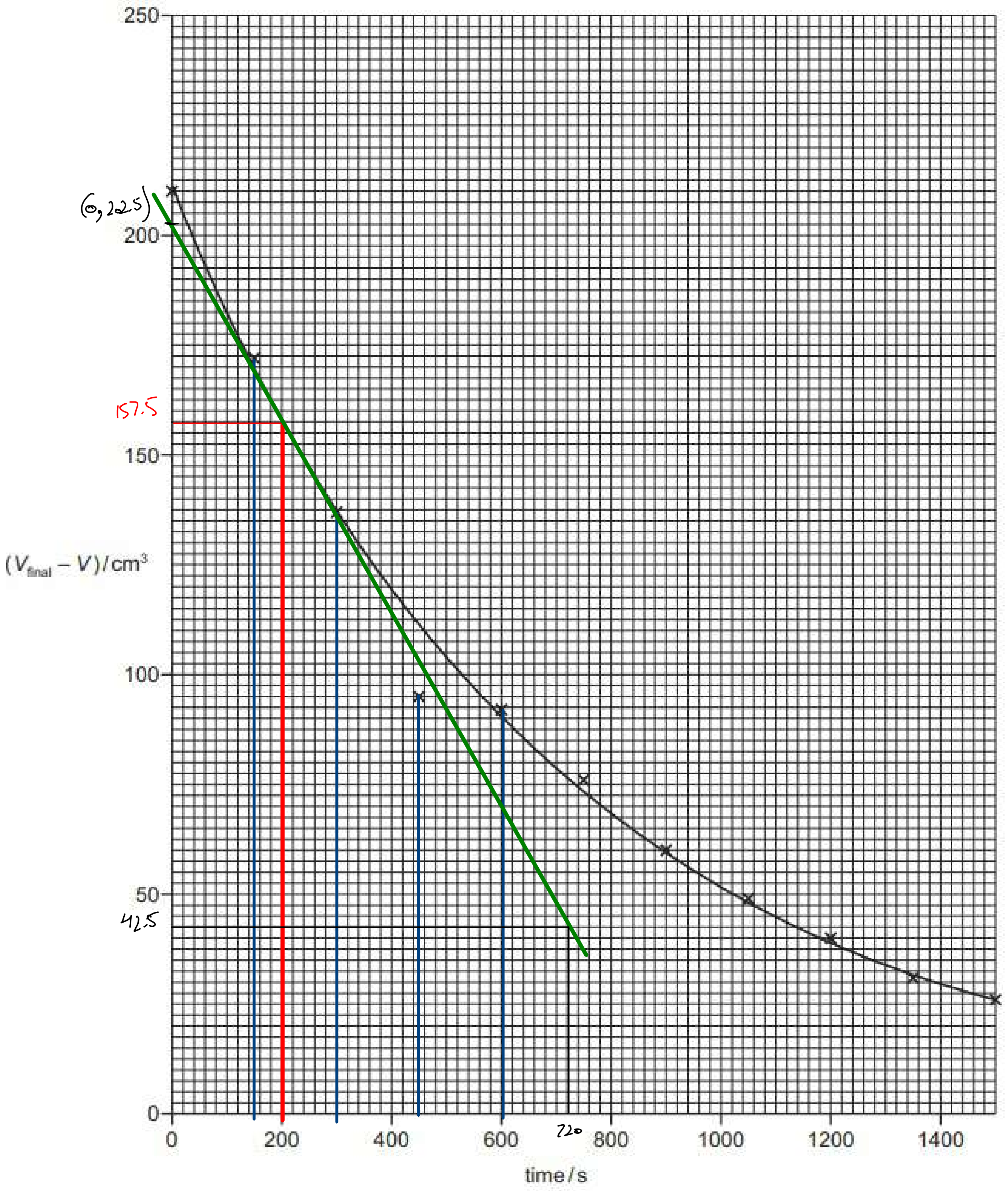
the rate of reaction would be too high so it would get really hard to read gas syringe at set intervals

State the effect this will have on the values of $V_{\text{final}} - V$.

lower for same time value

[2]

(c) The graph below shows the results obtained from another benzenediazonium chloride hydrolysis reaction performed at a different temperature.



(i) The point at time = 450 s is considered to be anomalous.

Suggest what caused the anomaly.

The reading was taken later than at 450s

[1]

(ii) The rate of reaction at different times can be calculated by drawing tangents to the best-fit line. The gradient of the tangent is equal to the rate of reaction, in $cm^3 s^{-1}$.

Use the graph in **(c)** to read the value of $(V_{final} - V)$ at time t = 200 s and write this value in the table below.

Draw a tangent to the curve at time t = 200 s. Use the tangent to determine the gradient at time t = 200 s.

State the co-ordinates of both points you used in your calculation.

co-ordinates 1 (0, 202.5) co-ordinates 2 (720, 42.5)

$$\frac{42.5 - 202.5}{720 - 0} = -0.222$$

gradient at 200 s = -0.222 $cm^3 s^{-1}$

Use your gradient to complete the table.

	time / s	$(V_{final} - V)/cm^3$	rate of reaction / $cm^3 s^{-1}$
1	200	158	-0.222
2	500	104	–0.143
3	600	91	–0.127
4	900	59	–0.0867
5	1000	52	–0.0720
6	1400	30	–0.0417

[4]

(iii) The concentration of benzenediazonium chloride is directly proportional to $(V_{final} - V)$.

Use the data in the table in **(ii)** to calculate the order of reaction with respect to benzenediazonium chloride.

You must show your working.

Comparing 4 & 6

Vfinal - V decreased by factor of $\frac{30}{59} = 0.51 \simeq 0.5$

rate decreased by factor of $\frac{-0.0417}{-0.0867} = 0.48 \simeq 0.5$

So 1st order

[2]

Electrochemistry:- Revisit Electrochemistry from A2

1 A more reactive metal will displace a less reactive metal from a solution of its salt. This reaction is exothermic. If the same reaction is set up in an electrochemical cell then, instead of an enthalpy change, electrical energy is produced and a cell voltage can be measured.

You are to plan an investigation of the reaction of three different metals (magnesium, iron and zinc) with aqueous copper(II) sulfate. You will plan to investigate whether there is a relationship between their cell potential values, $E^\ominus_{cell}$, and their enthalpy changes of reaction, ΔH_r.

$$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$$
$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$
$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Copper(II) sulfate solution is classified as a **moderate hazard**.

Zinc sulfate solution is classified as **corrosive**.

Iron(II) sulfate solution is classified as a **health hazard**.

(a) Predict how ΔH_r may change as $E^\ominus_{cell}$ increases. Give a reason for your prediction.

As $E^\ominus$cell increases, the ΔH_r will become more negative. The reaction will become more exothermic as the difference in reactivity increases [1]

(b) The first part of the investigation is to determine the enthalpy change, ΔH_r, for the reaction of the same number of moles of three powdered metals with 0.500 mol dm^{-3} copper(II) sulfate.

When determining the ΔH_r for the reaction of the metals listed above with aqueous copper(II) sulfate,

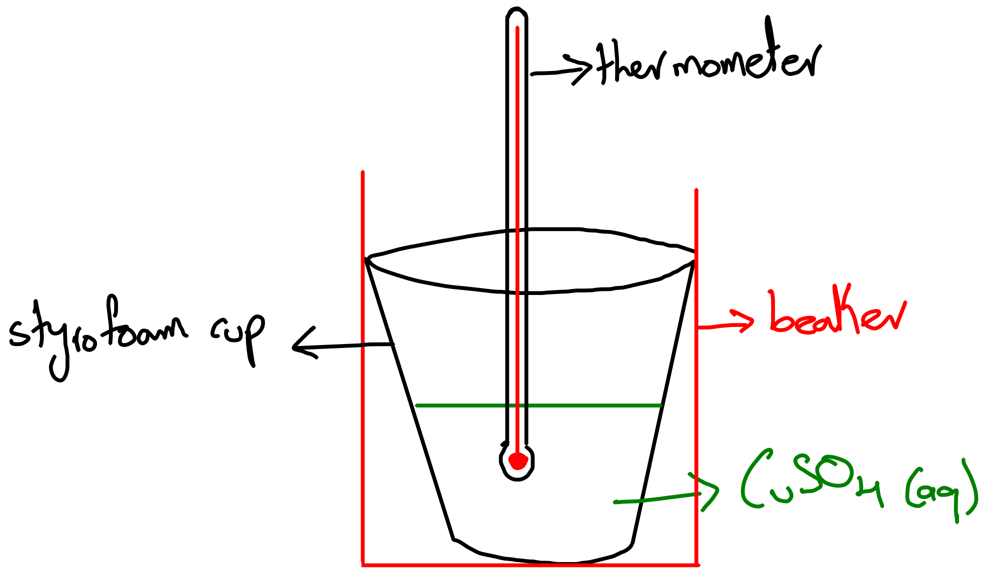
the independent variable is, type of powdered metal

the dependent variable is. temperature change or enthalpy change

[2]

You are provided with a sample of powdered metal and 50.0 cm^3 of 0.500 mol dm^{-3} aqueous copper(II) sulfate.

(c) (i) Draw a fully labelled diagram to show how the apparatus should be set up to allow you to determine the increase in temperature of aqueous copper(II) sulfate.
You should use apparatus normally found in a school or college laboratory.



[1]

(ii) State the measurements you would make in your experiment.

Mass of weighing boat + metal and mass of empty weighing boat after metal transferred and initial temperature before adding metal and highest temperature after adding metal [2]

(iii) Other than eye protection, state **one** precaution you would take to make sure that the experiment proceeds safely.

wear gloves

[1]

(iv) For the reaction with magnesium, calculate the mass of magnesium, in g, you would use so that it is in a small excess. You must show your working.
[A_r: Mg, 24.3]

n of $CuSO_4$ = c×V = 0.500 × 50×10^{-3} = 0.025 mol

n of Mg(s) = 0.025 mol

mass = mol × Mr / Ar = 0.025 × 24.3 = 0.675 g

mass of Mg = ...1...... g [2]

(v) Explain why the metal used should be in powdered form rather than in strips.

larger surface area increases rate of reaction.

[1]

(vi) The aqueous copper(II) sulfate and metal mixture should be stirred continuously. Explain why.

To ensure uniformity in heating of solution

[1]

(d) In one experiment, the increase in temperature when excess magnesium powder is added to $50.0\,cm^3$ of $0.500\,mol\,dm^{-3}$ aqueous copper(II) sulfate is 58.5 °C.

Calculate the enthalpy change for this reaction, ΔH_r, in $kJ\,mol^{-1}$.
Assume the specific heat capacity, *c*, of the reaction mixture is $4.18\,J\,g^{-1}K^{-1}$.
Assume $1.0\,cm^3$ of $0.500\,mol\,dm^{-3}$ aqueous copper(II) sulfate has a mass of 1.0 g.
Include a sign in your answer.

$$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$$

$Q = mc\Delta\theta = (50)(4.18)(58.5)$

$= 12226.5\ J$

$\Delta H_r = \frac{12226.5\times10^{-3}}{0.025} = 489.06$

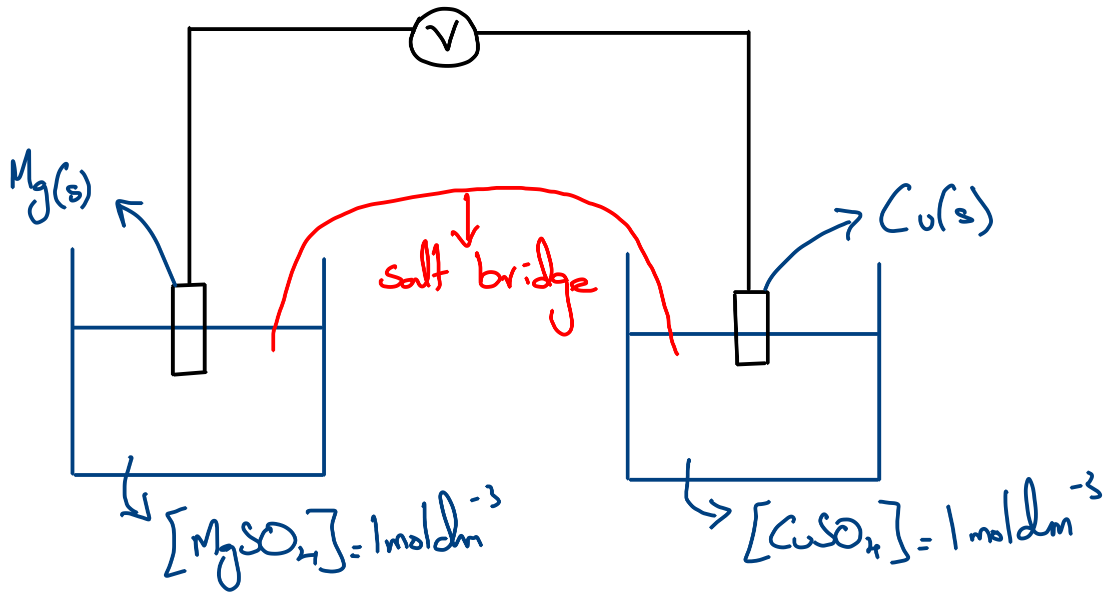
ΔH_r = ...-489...... $kJ\,mol^{-1}$ [2]

(e) The second part of the investigation involves determining the cell potential, E°_{cell}, for the three electrochemical cells.

cell reaction
$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Draw a diagram of the apparatus you would use to measure the E°_{cell} for the magnesium/copper cell. Your labels should include the **names** of the metals and the **names** and **concentrations** of the solutions you would use.

298 K



[3]

(f) Explain why the enthalpy change determination and cell potential determination should be carried out at the same temperature as each other.

so that values can be compared

[1]

(g) Accepted E°_{cell} values are shown for the cell reactions.

	cell reaction	E°_{cell}/V	ΔH_r
1	$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$	+2.72	-489
2	$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$	+1.10	-420
3	$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$	+0.78	-400

Use your prediction in **(a)**, your answer to **(d)** and data from the table to predict ΔH_r values for reactions 2 and 3.
Complete the table with these values. [1]

[Total: 18]

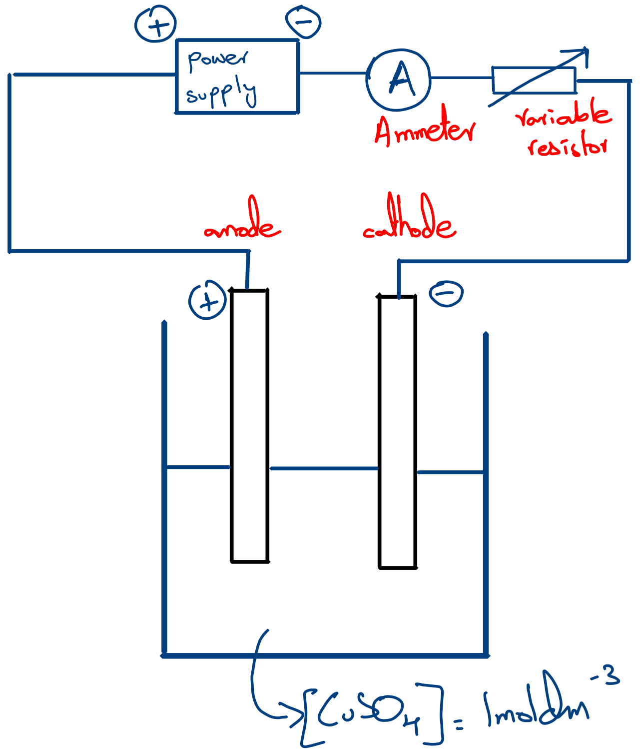
1 The Faraday constant is the charge in coulombs, C, carried by 1 mole of electrons.

(a) A student plans an electrolysis experiment to determine the Faraday constant.

The student was supplied with the following.

- 1.0 mol dm^{-3} copper(II) sulfate
- clean, dry copper foil electrodes, labelled ‘anode’ and ‘cathode’
- balance
- stop-clock
- ammeter
- other equipment suitable for carrying out electrolysis

Draw a labelled diagram of the apparatus and chemicals the student should use in their electrolysis experiment. Include in your diagram the circuit connecting the anode and cathode.



[2]

(b) Two of the hazards of using copper(II) sulfate solution are given below.

For each hazard, state a precaution, other than eye protection and a lab coat, that the student should take when carrying out the experiment.

hazard: copper(II) sulfate solution causes skin irritation

precaution ... wear gloves

hazard: copper(II) sulfate solution is toxic to aquatic life

precaution ... dont dispose down the drain

[2]

The student carried out the electrolysis for exactly 30 minutes with a current of 0.5A.

- After the electrolysis was finished, the student removed the electrodes.
- The electrodes were then carefully washed in water and then dipped in propanone.
- The electrodes were dried by allowing the propanone to evaporate.

(c) State the measurements the student would need to record to calculate the mass change of an electrode. Include the appropriate unit.

mass of electrode before and after experiment / g

[1]

(d) Calculate the charge passed through the copper(II) sulfate solution during the electrolysis experiment using the formula shown.

charge (C) = current (A) × time (s)

$= 0.5 \times 30 \times 60$

$= 900$

charge passed = 900 C [1]

(e) The mass change of the anode was −0.282 g.

Calculate the amount, in mol, of copper lost from the anode. Give your answer to 3 significant figures.
[A_r: Cu, 63.5]

$\text{mol} = \frac{\text{mass}}{M_r/A_r} = \frac{0.282}{63.5} = 4.44 \times 10^{-3}$

moles of copper lost from the anode = 4.44×10^{-3} mol [1]

(f) Use your answers to **(d)** and **(e)** to calculate the charge required to remove 1 mole of copper from the anode.

4.44×10^{-3} mol : 900

1 mol : x

$x = 202702.7027$

charge required to remove 1 mole of copper = 202703 C [1]

(g) The theoretical charge required to remove 1 mole of copper from the anode into solution as copper(II) ions is 193 000 C.
The Faraday constant is 96 500 C mol^{-1}.

Explain why the theoretical charge is twice the Faraday constant.

It is twice as 2 moles of electrons are required to remove 1 mole of electron.

[1]

(h) A possible source of error is not drying the anode at the start of the experiment.

Explain the effect, if any, on the calculated value of the Faraday constant if the anode is wet at the beginning of the experiment but dry at the end.

effect The value will be less

explanation the mass and therefore moles will be recorded as higher than actual.

mass ↑ → mol ↑ → charge required to remove 1 mole ↓ [1]

(i) The student wanted to ensure that the anode was completely dry at the end of the experiment and decided to evaporate off the propanone using a blue Bunsen flame. The student noticed some blackening of the surface of the copper.

Suggest what caused this blackening.

incomplete combustion occurs so carbon soot is deposited

[1]

(j) The student calculated the mass change of the anode and the cathode after the experiment was complete.

mass change of anode = −0.282 g   mass change of cathode = +0.217 g

Suggest **one** reason why the mass gained at the cathode is **not** the same as the mass lost at the anode. Assume the student has recorded the mass changes correctly.

some copper falls off the electrode and gets deposited at the bottom

[1]

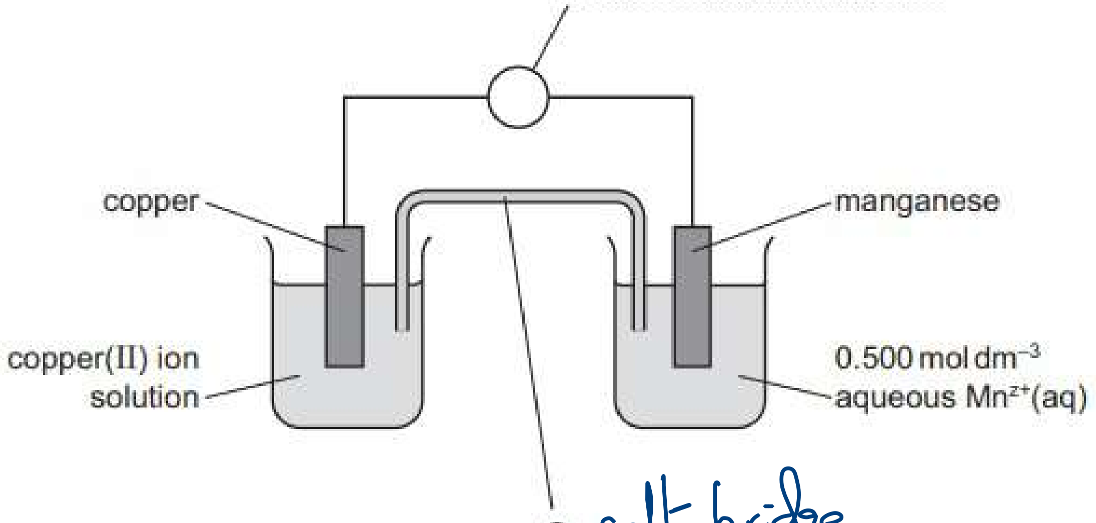
[Total: 12]

1 A student investigates the charge (z+) carried by aqueous manganese ions, $Mn^{z+}(aq)$. The electrochemical cell shown is set up for this investigation with the following two half-cells:

- a standard copper(II) ion / copper half-cell ($E^\circ = +0.340\,V$)
- a half-cell made from manganese and $0.500\,mol\,dm^{-3}$ $Mn^{z+}(aq)$.

(a) Label the items **P** and **Q** and state the concentration of the copper(II) ion solution in the copper half-cell.

P voltmeter



Q salt bridge

concentration of the copper(II) ion solution in the copper half-cell = $1\,mol\,dm^{-3}$ [1]

(b) During the investigation the student plans to use solutions of $Mn^{z+}(aq)$ of lower concentration than $0.500\,mol\,dm^{-3}$.

(i) Calculate the volume of $0.500\,mol\,dm^{-3}$ $Mn^{z+}(aq)$ needed to prepare $100.0\,cm^3$ of $0.200\,mol\,dm^{-3}$ $Mn^{z+}(aq)$.

$$c_1V_1 = c_2V_2$$

$$(0.500)(V_1) = (0.200)(100\times10^{-3})$$

$$V_1 = 0.04\,dm^3 \times 1000 = 40\,cm^3$$

volume = 40 cm^3 [1]

(ii) Describe how, using a $100\,cm^3$ volumetric flask, the student should prepare exactly $100.0\,cm^3$ of $0.200\,mol\,dm^{-3}$ $Mn^{z+}(aq)$ using the volume of $0.500\,mol\,dm^{-3}$ $Mn^{z+}(aq)$ calculated in **(b)(i)** and standard school or college apparatus.

Transfer $40\,cm^3$ of $0.500\,mol\,dm^{-3}$ into a $100\,cm^3$ volumetric flask. Use distilled water to fill the volumetric flask upto $100\,cm^3$ mark.

[2]

The cell potential of the electrochemical cell in **(a)** is measured. The $0.500\,mol\,dm^{-3}$ $Mn^{z+}(aq)$ is then replaced by the $0.200\,mol\,dm^{-3}$ solution and the cell potential is measured again. This is repeated for other lower concentrations of $Mn^{z+}(aq)$. All measurements are made at 25 °C.

(iii) The results of the experiment are shown in the table.

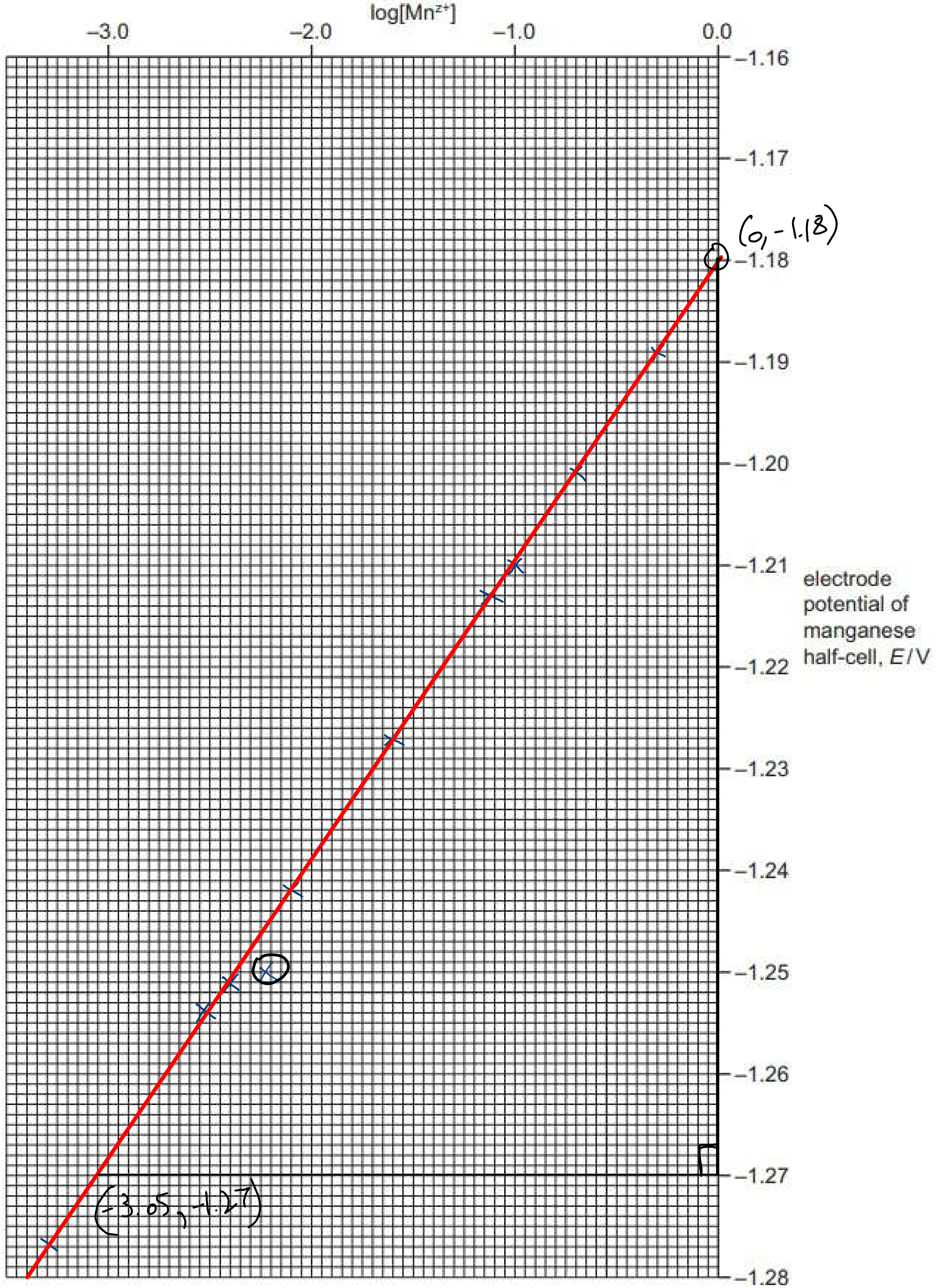
Complete column three of the table, calculating $\log[Mn^{z+}]$ to **two decimal places**.
Complete column four of the table, calculating E, the electrode potential of each manganese half-cell, to **three decimal places**, using the equation shown.

$$E(\text{manganese half-cell}) = E_{cell} + 0.340\,V$$

$[Mn^{z+}]$ / mol dm⁻³	cell potential, E_{cell}/V	$\log[Mn^{z+}]$	electrode potential of each manganese half-cell, E/V
5.0×10^{-1}	–1.529	–0.30	–1.189
2.0×10^{-1}	–1.541	-0.70	-1.201
1.0×10^{-1}	–1.550	-1.00	-1.210
7.5×10^{-2}	–1.553	-1.12	-1.213
2.5×10^{-2}	–1.567	-1.60	-1.227
8.0×10^{-3}	–1.582	-2.10	-1.242
6.0×10^{-3}	–1.590	-2.22	-1.250
4.0×10^{-3}	–1.591	-2.40	-1.251
3.0×10^{-3}	–1.594	-2.52	-1.254
5.0×10^{-4}	–1.617	-3.30	-1.277

[2]

(c) Plot a graph of electrode potential of manganese half-cell (*y*-axis) against $\log[\mathrm{Mn}^{2+}]$ (*x*-axis).
Use a cross (×) to plot each data point. Draw a line of best fit.



[2]

(d) (i) Circle the most anomalous point on your graph. [1]

(ii) The student is careful to ensure that all solutions used are at the same temperature in all experiments.

Suggest a possible explanation for the position of the anomalous point circled in **(d)(i)** relative to the line of best fit.

The solution is more dilute than it should be

[1]

(e) Your graph is a plot of *E* against $\log[\mathrm{Mn}^{2+}]$ and can be analysed using the Nernst equation at 25 °C.

$$\overset{y\,=}{E} = \underset{c\,+}{E^{\circ}} + \underset{m}{\frac{0.059}{z}}\underset{x}{\left(\log[\mathrm{Mn}^{2+}]\right)}$$

z is the value of the charge carried by the manganese ion
E is the electrode potential / V
E° is the standard electrode potential / V

Use the Nernst equation and your graph to find the standard electrode potential, E°, of the manganese half-cell.

E° = -1.18 V [1]

(f) (i) Determine the gradient of the graph.
State the co-ordinates of both points you used for your calculation.
Record the value of the gradient to **three significant figures**.

co-ordinates 1 (0, -1.18) co-ordinates 2 (-3.05, -1.27)

$$\frac{-1.27+1.18}{-3.05-0} = 0.0295$$

gradient = 0.0295
[2]

(ii) Use your answer to **(f)(i)** and the Nernst equation to calculate the value of z to three significant figures and give the formula of the manganese ion.
Your calculation **must** show the use of the Nernst equation.

(If you were unable to calculate an answer to **(f)(i)** you may use the value 0.0197. This is **not** the correct value.)

$$0.0295 = \frac{0.059}{z}$$

$z = 2$

z = 2

formula of manganese ion = Mn^{2+}

[2]

(g) Lowering [Mn^{z+}] causes the value of the electrode potential of the manganese half-cell to become more negative.

Suggest why this happens.

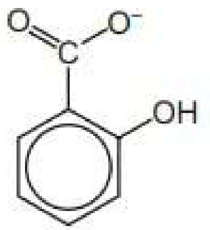
The equilibrium $Mn_{(s)} \rightleftharpoons Mn^{2+} + 2e^-$ shifts to the right hand side, more electrons are produced so more negative. [1]

[Total: 16]

2 Transition metal complex ions are coloured. The formula of a complex ion can be determined using colorimetry.

In colorimetry, light of a certain wavelength is passed through a complex ion solution. The absorbance of the light is proportional to the intensity of the colour of the solution. The more concentrated the complex ion solution, the more intense its colour and so the higher the absorbance.

A student carried out an experiment to determine the formula of the complex ion formed between aqueous iron(III) ions, Fe^{3+}(aq), and aqueous 2-hydroxybenzoate ions, $C_6H_4(OH)CO_2^-$, which have the structure shown.



(a) In the first step of the experiment the student prepared 100.0 cm^3 of 0.0500 mol dm^{-3} aqueous iron(III) nitrate.

(i) Determine the mass, in g, of solid hydrated iron(III) nitrate, $Fe(NO_3)_3.9H_2O$, needed to prepare 100.0 cm^3 of a 0.0500 mol dm^{-3} solution.
[A_r: Fe, 55.8; N, 14.0; O, 16.0; H, 1.0]

$n = c \times V = 0.05 \times 100 \times 10^{-3} = 5 \times 10^{-3}$ mol

mass = mol × M_r/A_r = $5 \times 10^{-3} \times (55.8 + 3(19) + 9(16) + 18 + 9(16))$

= 2.019g

mass of $Fe(NO_3)_3.9H_2O$ = 2.02 g [2]

(ii) Describe how, after weighing the mass determined in **(i)**, the student should prepare 100.0 cm^3 of 0.0500 mol dm^{-3} aqueous iron(III) nitrate.

In your answer you must give the name and capacity, in cm^3, of any apparatus used.

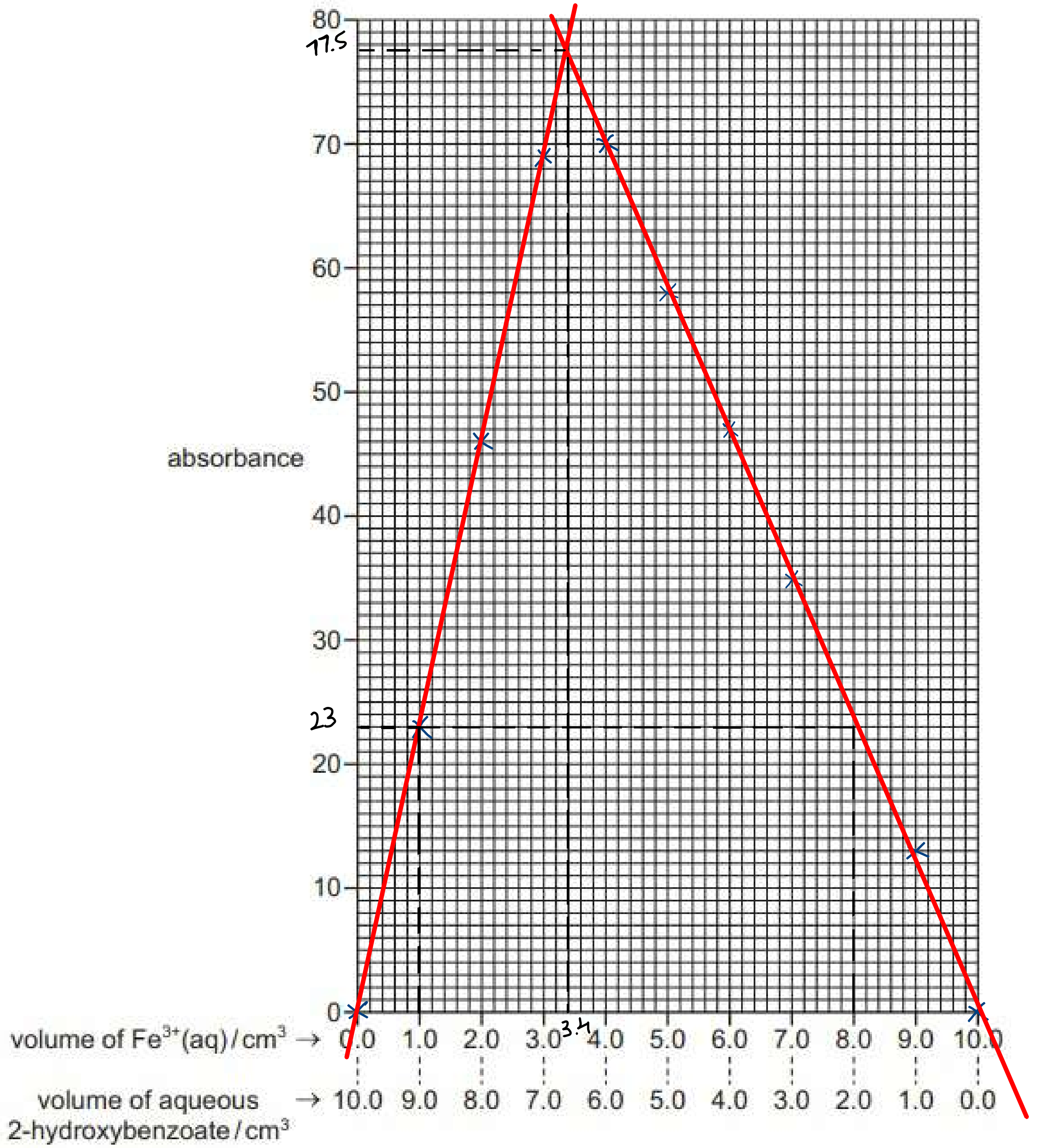
Transfer 2.02g into beaker and dissolve it using distilled water. Transfer the contents of the beaker to a 100cm³ volumetric flask. Transfer the washings of the beaker to the flask too and then using distilled water fill volumetric flask upto 100cm³ mark. [2]

(b) The student prepared solutions containing various combinations of 0.0500 mol dm^{-3} Fe^{3+}(aq) and 0.0500 mol dm^{-3} aqueous 2-hydroxybenzoate, as shown in the table.

The student placed a small sample of each solution into a colorimeter and measured the absorbance. The student made a mistake in test number **9** and did **not** measure the result.

test number	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**
volume of Fe^{3+}(aq) / cm^3	0.0	1.0	2.0	3.0	4.0	5.0	6.0	7.0		9.0	10.0
volume of aqueous 2-hydroxybenzoate / cm^3	10.0	9.0	8.0	7.0	6.0	5.0	4.0	3.0		1.0	0.0
absorbance	0	23	46	69	70	58	47	35		13	0

(i) Plot a graph on the grid to show the relationship between absorbance and the volumes of Fe^{3+}(aq) and aqueous 2-hydroxybenzoate used.
Use a cross (×) to represent each data point. Draw **two** lines of best fit. [2]



(ii) Use the graph in **(i)** to determine the volumes of $Fe^{3+}(aq)$ and aqueous 2-hydroxybenzoate which would give the maximum absorbance.

volume of $Fe^{3+}(aq)$ = 3.4 cm^3

volume of aqueous 2-hydroxybenzoate = 6.6 cm^3 [1]

(iii) The point of maximum absorbance shows where all of the ions are combined in the complex.

Use the volumes in **(ii)** to determine the number of moles of 2-hydroxybenzoate ions that form a complex with 1 mole of Fe^{3+} ions.

$$\frac{6.6}{3.4} \simeq 2$$

moles of 2-hydroxybenzoate ions = 2 [1]

(iv) $Fe^{3+}(aq)$ ions exist in aqueous solution as complex ions with the formula $[Fe(H_2O)_6]^{3+}(aq)$. 2-hydroxybenzoate ions, $C_6H_4(OH)CO_2^-$, are bidentate ligands.

Use this information and your answer to **(iii)** to suggest the formula of the complex ion formed between $Fe^{3+}(aq)$ ions and 2-hydroxybenzoate ions.

$[Fe(H_2O)_2(C_6H_4(OH)CO_2)_2]^+$ [1]

(v) Name the apparatus that should be used to measure the volumes of the solutions given in the table accurately.

burette [1]

(c) In test **9**, instead of mixing 8.0 cm^3 of $Fe^{3+}(aq)$ and 2.0 cm^3 of aqueous 2-hydroxybenzoate, the student mixed 16.0 cm^3 of $Fe^{3+}(aq)$ and 4.0 cm^3 of aqueous 2-hydroxybenzoate.

Use your graph in **(b)(i)** to suggest the absorbance that would have been measured if a sample of this solution had been analysed in the colorimeter.

absorbance = 23 % [1]

(d) In a colorimetry experiment, the absorbance of the solution follows the relationship shown.

$$A = \varepsilon c l$$

A is the absorbance (no units).
c is the concentration in $mol\,dm^{-3}$.
l is the path length of the light travelling through the solution in cm.
ε is the molar absorption coefficient (a constant).

Determine the unit of *ε*.

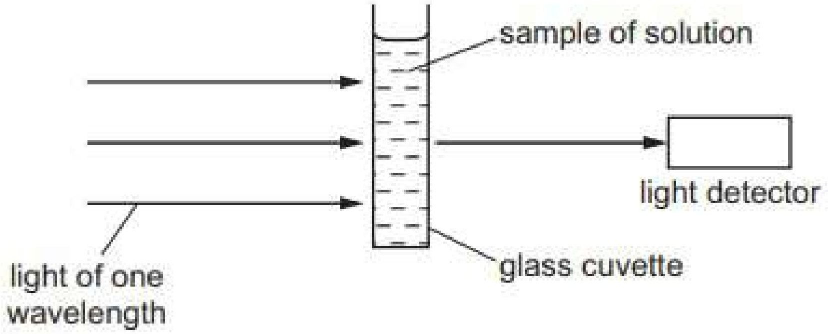
$$A = \varepsilon c l$$

$$\varepsilon = \frac{A}{c l} = \frac{\text{no units}}{mol\,dm^{-3} \times cm} = dm^3\,mol^{-1}\,cm^{-1}$$

unit = $dm^3\,mol^{-1}\,cm^{-1}$ [1]

[Total: 12]

2 When light passes through solutions of chemical compounds some of the light may be absorbed. The quantity of light absorbed is called the absorbance and it is measured by a spectrophotometer. A simplified diagram of a spectrophotometer is shown. A glass cuvette is a rectangular vessel.



(a) (i) A chemist placed distilled water in the glass cuvette. This was then put into the spectrophotometer and a reading taken.

Explain why this reading was taken.

To callibrate the instrument

[1]

(ii) Light passes through opposite sides of the cuvette. These two sides must be wiped with a cloth to ensure they are clean and dry.

Explain why this procedure makes the readings more accurate.

The light will not get absorbed by dirt / fingerprints

[1]

Manganese is added to steel to increase its strength. A spectrophotometer can be used to analyse the manganese content of steel. This is done by comparing the absorbance of a solution of $MnO_4^-(aq)$ prepared from a sample of steel, with the absorbance of solutions of known concentrations of $MnO_4^-(aq)$.

(b) 1.0 dm^3 of a standard solution of 0.0300 mol dm^{-3} MnO_4^- was prepared by a chemist using solid potassium manganate(VII), $KMnO_4$, measured using a two decimal place balance.

(i) Calculate the mass of $KMnO_4$ required to prepare this standard solution.
[A_r: K, 39.1; Mn, 54.9; O, 16.0]

$n = c \times V = 0.0300 \times 1 = 0.03$ mol

mass = mol $\times M_r/A_r = 0.03 \times (39.1 + 54.9 + 4(16))$

$= 4.74$

mass of $KMnO_4$ = 4.74 g [1]

(ii) Describe how the chemist should accurately prepare this standard solution using a sample of $KMnO_4$ of mass calculated in **(i)**. There is a 1.0 dm^3 volumetric flask available.

Dissolve 4.74g in a beaker using distilled water. Transfer the solution to a 1 dm^3 volumetric flask. Transfer the washings of the beaker to the volumetric flask. Fill the volumetric flask upto the 1 dm^3 mark.

[2]

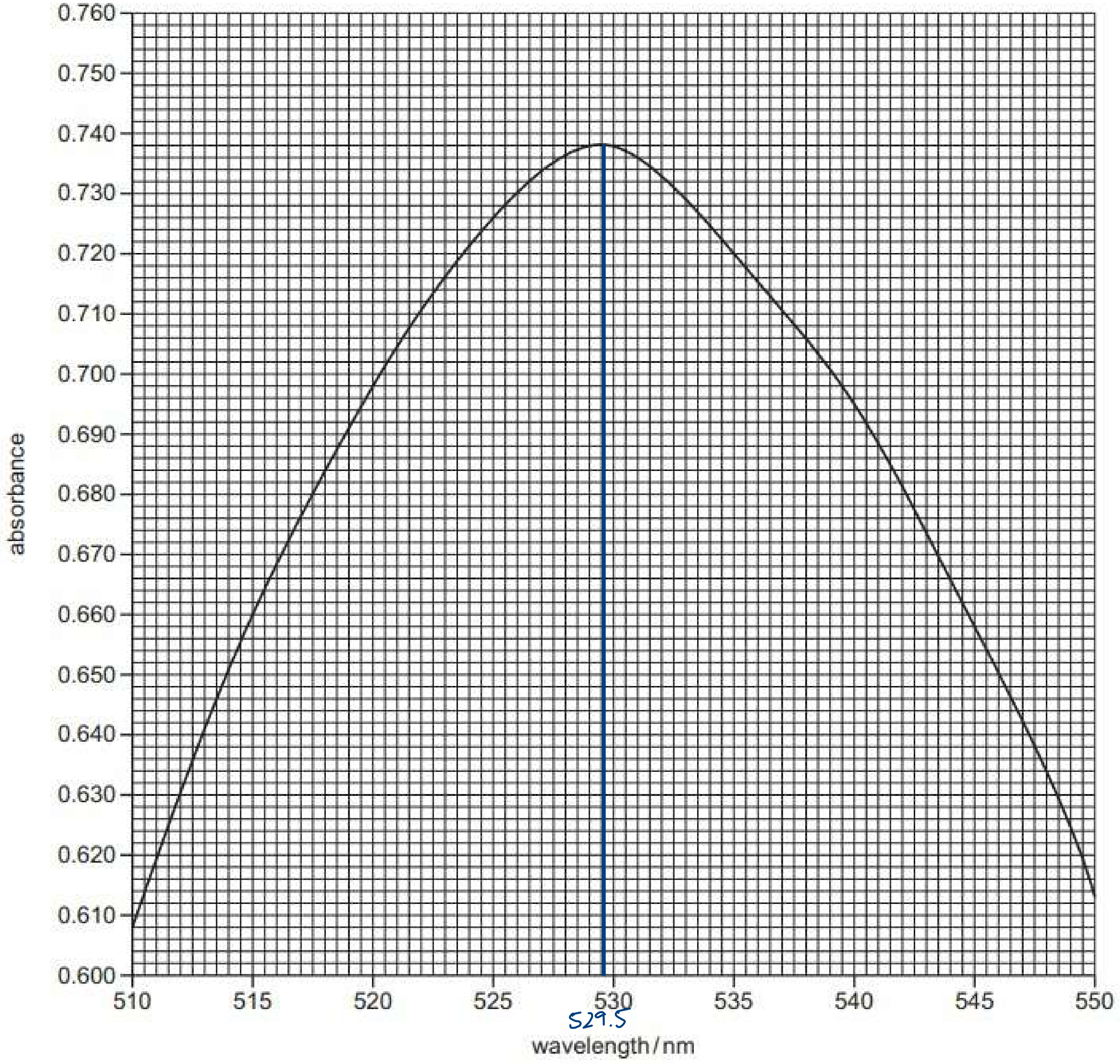
(iii) The chemist diluted this standard solution to 3.0×10^{-4} mol dm^{-3} for use in the experiments.

Explain why the chemist did not prepare a solution of this concentration directly, by dissolving the required mass of $KMnO_4$ in 1.0 dm^3 of water.

A small mass would have to be used which would have a high percentage error and be really difficult to measure.

[1]

(c) The chemist needed to determine which wavelength of light was most absorbed by a solution of MnO_4^-(aq). The clean, dry cuvette was filled with 3.0×10^{-4} mol dm^{-3} MnO_4^- and different wavelengths of light were passed through the solution. A graph of the results was plotted.



Use the graph to estimate the wavelength of light that is most absorbed by the MnO_4^- solution.

wavelength of light most absorbed = 529.5 nm [1]

(d) The spectrophotometer was then set to the wavelength that is most absorbed by the MnO_4^-(aq) solution.

The chemist measured the absorbance of solutions of known concentrations of MnO_4^-(aq). The results are shown in the table.

concentration of MnO_4^-(aq) / mol dm^{-3}	absorbance
3.00×10^{-4}	0.748
2.70×10^{-4}	0.680
2.40×10^{-4}	0.610
2.10×10^{-4}	0.530
1.80×10^{-4}	0.440
1.50×10^{-4}	0.378
1.20×10^{-4}	0.315
0.90×10^{-4}	0.230
0.60×10^{-4}	0.150

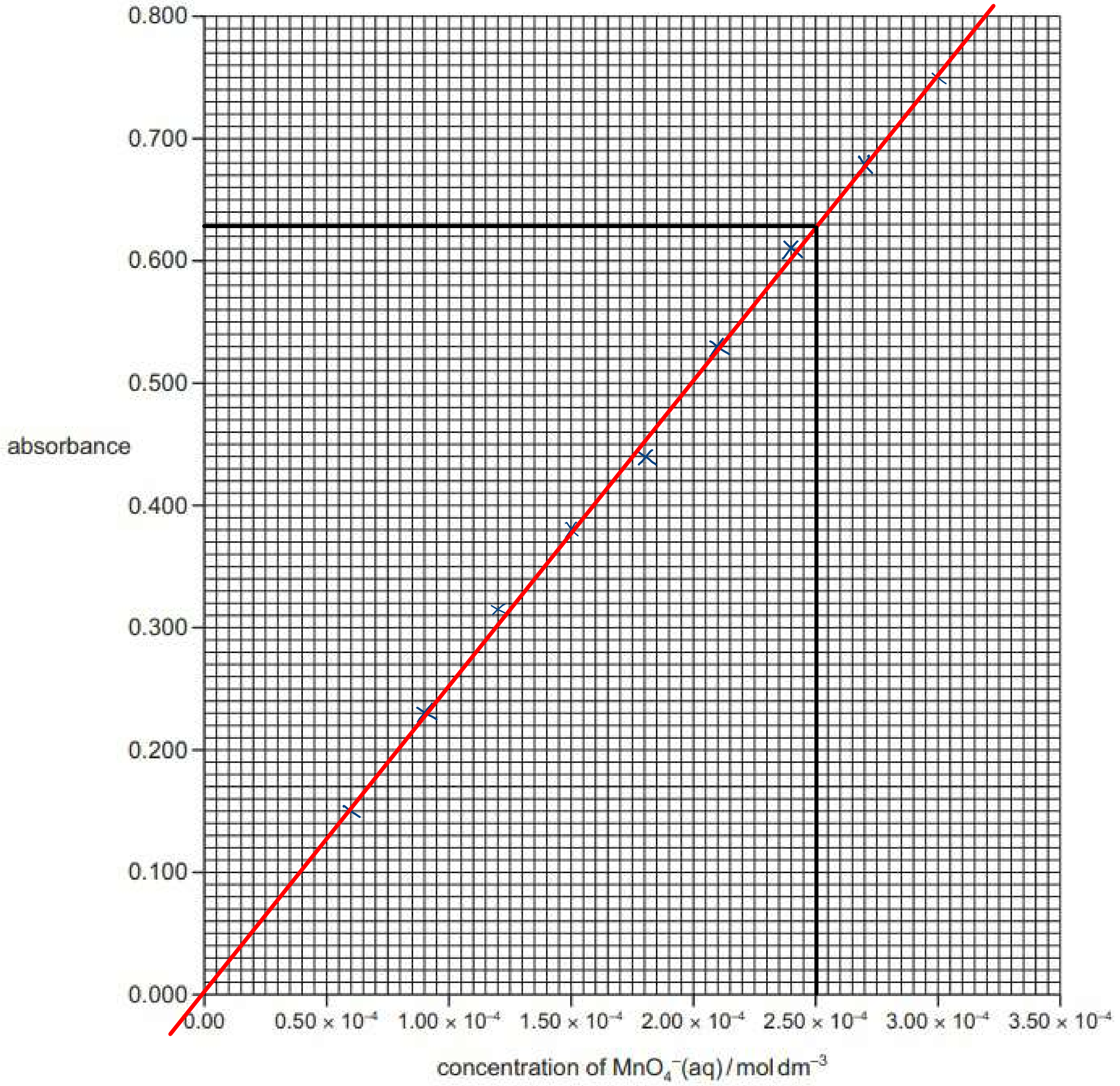
(i) Plot a graph on the grid on page 11 to show the relationship between the absorbance and the concentration of MnO_4^-(aq).
Use a cross (×) to plot each data point. Draw a line of best fit. [2]

(ii) State the relationship between absorbance and concentration of MnO_4^-(aq). Explain your answer with reference to particles.

absorbance is directly proportional to concentration of MnO_4^-
The more there particles there are, the more light is absorbed.
[2]

(iii) Do you consider the results obtained to be reliable? Explain your answer.

Yes, all points are on or close to the line
[1]



(e) **(i)** The chemist used the MnO_4^-(aq) solution of concentration 3.00×10^{-4} mol dm^{-3} to prepare the solutions in the table on page 10.

Calculate the volume of 3.00×10^{-4} mol dm^{-3} MnO_4^-(aq) solution and the volume of distilled water required to prepare a 25.00 cm^3 solution of 2.70×10^{-4} mol dm^{-3} MnO_4^-(aq).
Give your answers to **two decimal places**.

$c_1V_1 = c_2V_2$

$(3\times10^{-4})V_1 = (2.70\times10^{-4})\times25\times10^{-3}$

$V_1 = 0.0225\ dm^3 \times 1000 = 22.5\ cm^3$

volume of 3.00×10^{-4} mol dm^{-3} MnO_4^-(aq) solution = ...22.5.... cm^3

volume of distilled water = ...2.5.... cm^3
[1]

(ii) The volumes of the two solutions given in **(e)(i)** could be measured using the same type of apparatus.

Name a suitable piece of apparatus which could be used to measure these volumes.

burette .... [1]

The chemist dissolved a known mass of steel, containing manganese, in acid. The manganese was then oxidised to manganate(VII) ions, MnO_4^-, using a very strong oxidising agent. The resulting solution was made up to 100.0 cm^3 in a volumetric flask.

(f) A small sample of the solution of MnO_4^-(aq) prepared from the steel sample was placed into a clean, dry cuvette and its absorbance measured using the spectrophotometer.

(i) The absorbance of the MnO_4^-(aq) solution was 0.630.

Use the graph you have drawn in **(d)(i)** to determine the concentration of MnO_4^-(aq) in this solution.
Give your answer to **three significant figures**.

concentration of MnO_4^-(aq) = ...2.50×10^{-4}.... mol dm^{-3} [1]

(ii) Calculate the mass of manganese present in the steel sample. Show your working.
[A_r: Mn, 54.9]

$n = c\times V = 2.5\times10^{-4} \times 100\times10^{-3} = 2.5\times10^{-5}$

$mass = mol \times M_r/A_r$

$= 2.5\times10^{-5} \times 54.9$

$= 1.3725\times10^{-3}$

mass of manganese = ...1.37×10^{-3}.... g [1]

(g) The steel sample that the chemist used had a mass of 1.209 g.

Use the mass of manganese you calculated in **(f)(ii)** to calculate the percentage of manganese by mass that was present in the steel sample.
(If you were unable to calculate an answer to **(f)(ii)** you may use 0.00143 g as the mass of manganese. This is not the correct answer.)

$\frac{1.37\times10^{-3}}{1.209} \times 100 = 0.1133\ \%$

percentage of manganese in the steel sample = 0.113 % [1]

(h) Another way of analysing the manganese content of the steel sample is by titration. The steel sample is prepared in the same way as previously. It is dissolved in acid and then oxidised using a very strong oxidising agent. The $MnO_4^-(aq)$ ions produced are titrated with a solution of iron(II) ions. The equation for this reaction is shown.

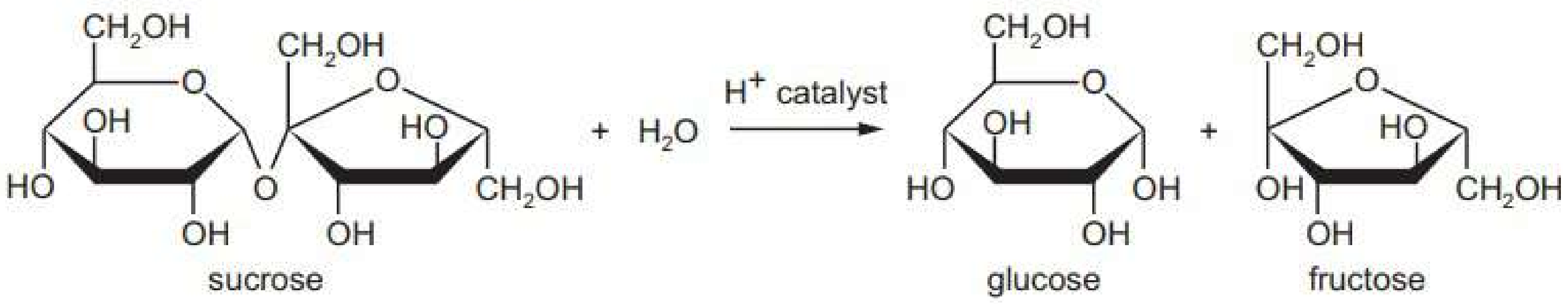
$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$$

Explain why it is essential to remove the strong oxidising agent used to prepare the solution of steel sample before carrying out the titration.

So that Fe^{2+} is not oxidised due to the strong oxidising agent [1]

[Total: 18]

2 Sucrose, $C_{12}H_{22}O_{11}$, is a naturally occurring sugar found in sugarcane and many fruits. It can be hydrolysed in acidic solution to give glucose and fructose. All three molecules are chiral and will rotate the plane of polarised light. The degree of rotation is known as the **optical rotation**, α.



In the presence of excess water, the reaction can be considered to be first order with respect to sucrose concentration.

The progress of the reaction can be monitored using a polarimeter, which measures the optical rotation, α, of the solution. The more concentrated the solution, the greater the optical rotation of the solution.

The concentration of sucrose at time t can be represented as $(\alpha - \alpha_{\text{final}})$, where α_{final} is the optical rotation of the solution after 6 hours.

The mathematical relationship is given by the following equation.

$$\log_{10}(\alpha - \alpha_{\text{final}}) = A - \frac{kt}{2.30}$$

A is a constant.
k is the rate constant.

(a) The experimentally determined values of optical rotation during the hydrolysis of sucrose at 298 K are recorded below.

Process the results to allow you to plot a graph of $\log_{10}(\alpha - \alpha_{\text{final}})$ against time, t.

Calculate $(\alpha - \alpha_{\text{final}})$ and record it to 1 decimal place.

Calculate $\log_{10}(\alpha - \alpha_{\text{final}})$ and record it to 2 decimal places.

time / s	optical rotation, α	$(\alpha - \alpha_{\text{final}})$	$\log_{10}(\alpha - \alpha_{\text{final}})$
0	39.9	51.9	1.72
300	29.1	41.1	1.61
600	21.3	33.3	1.52
900	15.5	27.5	1.44
1200	10.6	22.6	1.35
1500	6.2	18.2	1.26
1800	2.4	14.4	1.16
2100	–0.3	11.7	1.07
2400	–2.5	9.5	0.98
2700	–4.5	7.5	0.88

α_{final}	–12.0

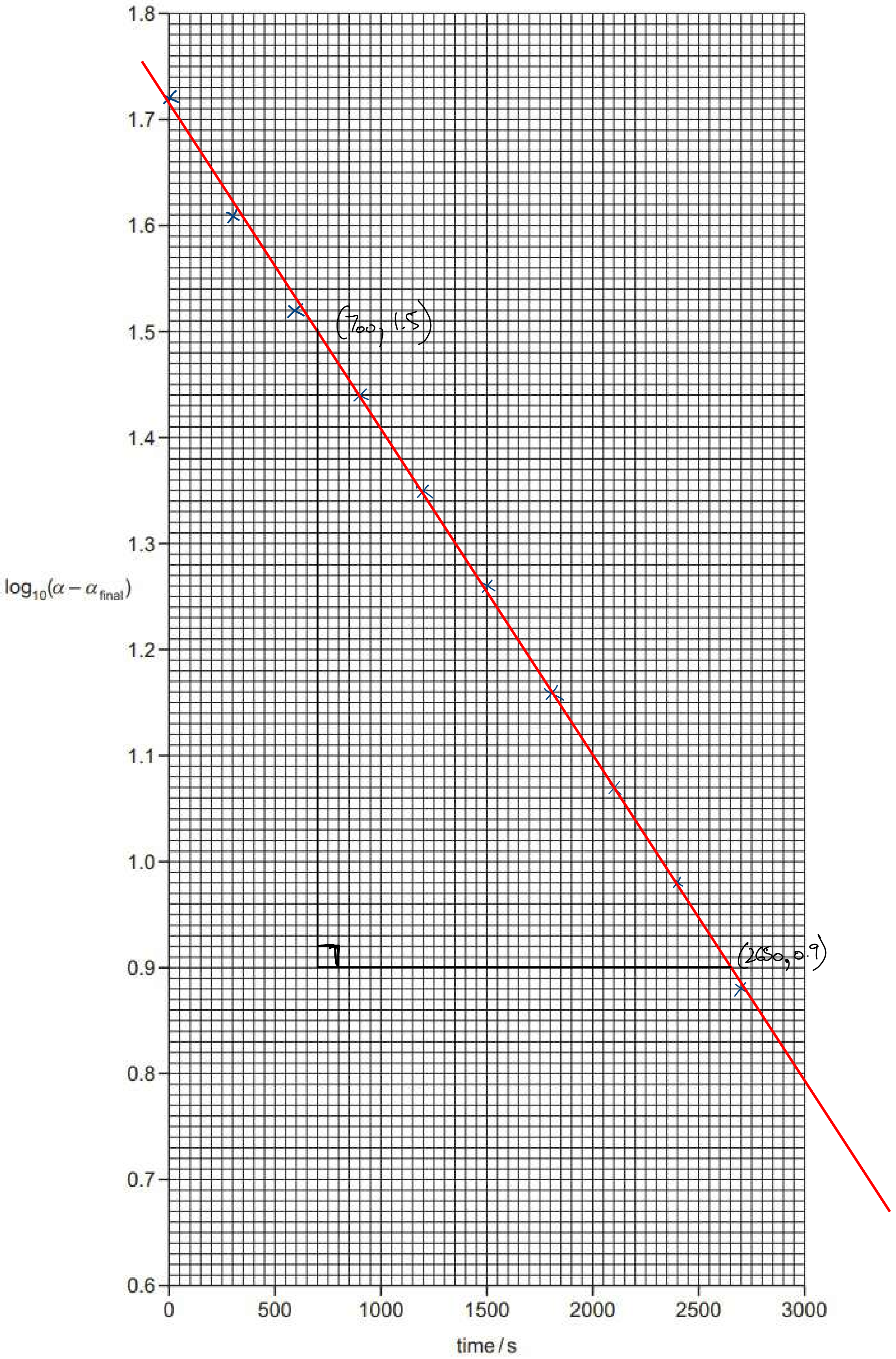
[2]

(b) **(i)** Plot a graph on the grid on page 9 to show how $\log_{10}(\alpha - \alpha_{\text{final}})$ varies with time, t. Use a cross (**x**) to plot each data point. Draw the line of best fit. [2]

(ii) State and explain whether the results and your graph confirm the relationship $\log_{10}(\alpha - \alpha_{\text{final}}) = A - \frac{kt}{2.30}$.

$$y = c + m \quad x$$

Yes, it is a straight line with intercept as expected. All of the points are either on the line or very close to it. [1]



(c) **(i)** Determine the gradient of the graph.
State the co-ordinates of both points you used for your calculation.
Record the value of the gradient to **three significant figures**.

co-ordinates 1 (700, 1.5) co-ordinates 2 (2650, 0.9)

$$\frac{0.9 - 1.5}{2650 - 700} = m$$

$$m = -3.08 \times 10^{-4}$$

gradient = -3.08×10^{-4} s^{-1} [2]

(ii) Use the gradient value to calculate a value for *k* in the expression shown.

$$\log_{10}(\alpha - \alpha_{final}) = A - \frac{kt}{2.30}$$

$$\log_{10}(\alpha - \alpha_{final}) = \left(\frac{-K}{2.30}\right)t + A$$

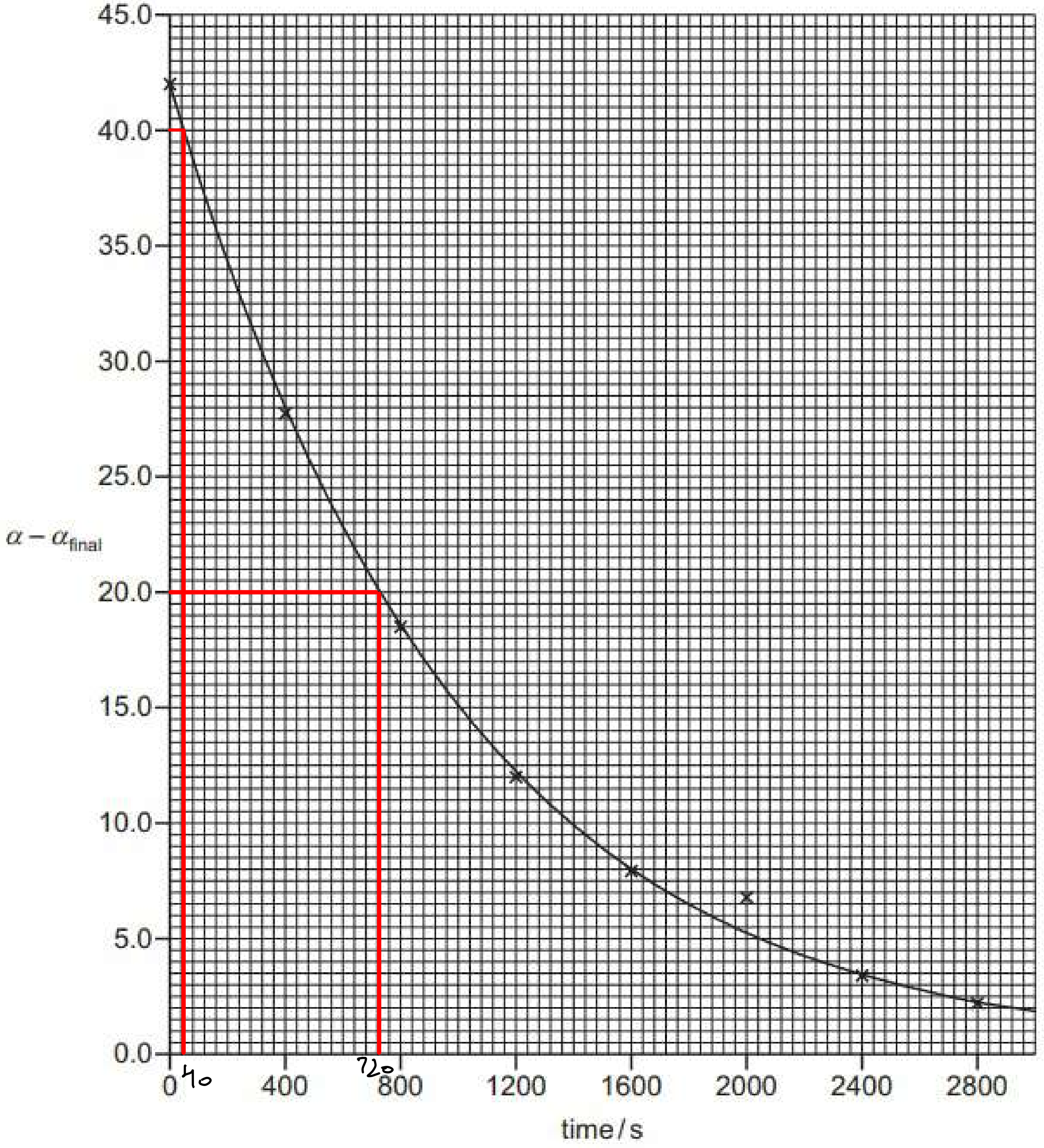
$$y = m \; x + c$$

$$\frac{+K}{2.30} = +3.08 \times 10^{-4}$$

$$K = 7.084 \times 10^{-4} \; s^{-1}$$

[2]

(d) The graph below shows the results obtained from a second hydrolysis of sucrose reaction performed at a different temperature.



(i) The point at time = 2000 s is considered to be anomalous.
Suggest what caused the anomaly.

The reading was taken earlier than 2000s

[1]

(ii) Use the graph to determine the half-life, $t_{½}$, of this reaction.
State the co-ordinates of both points you used in your calculation.

co-ordinates 1 (40, 40) co-ordinates 2 (720, 20)

720 − 40 = 680

half-life = 680 s [2]

(iii) For a first-order reaction, the following relationship exists.

$$\text{half-life, } t_{½} = \frac{0.693}{k'}$$

Use this relationship and your answer to **(ii)** to determine k', the rate constant for this second hydrolysis reaction.
If you have been unable to determine the half-life of the reaction in **(ii)**, you may use the value $t_{½}$ = 500 s, though this is not the correct answer.

$$\frac{0.693}{680} = K' = 1.02 \times 10^{-3}$$

k' = 1.02×10^{-3} s^{-1} [1]

(iv) State whether the temperature of the second reaction was higher or lower than that of the first.
Explain your answer with reference to the answers you obtained in **(c)(ii)** and **(d)(iii)**.
If you have been unable to calculate a value for k in **(c)(ii)**, you may use the value $k = 8.00 \times 10^{-4}$, though this is not the correct answer.

2nd reaction took place at higher temperature as its K is higher

[1]

(v) Would the value of the half-life change if the reaction were repeated with twice the initial concentration of sucrose? Give a reason for your answer.

No, half life is independant of concentration

[1]

2 Activated charcoal is a form of carbon with a very high surface area. It can be used to remove impurities from mixtures. It does this by a process called *adsorption*, where particles of the impurity bond (adsorb) to the activated charcoal surface.

A student wanted to determine the ability of activated charcoal to adsorb a blue dye (the impurity) from aqueous solution.

The equation that links the mass of activated charcoal with the amount of blue dye adsorbed is shown.

$$\log\left(\frac{D}{m}\right) = A + b\log[X]$$

$$y = c + m x$$

D = difference in concentration of dye (in $g\,dm^{-3}$) before and after adsorption
m = mass of activated charcoal (in g)
$[X]$ = final concentration of dye (in $g\,dm^{-3}$) after adsorption
A and b are constants

The student used the following procedure to investigate the ability of activated charcoal to adsorb a blue dye from an aqueous solution.

- Place a $50.0\,cm^3$ sample of a $25.00\,g\,dm^{-3}$ solution of blue dye in a conical flask.
- Add a weighed mass of activated charcoal to the flask.
- Stir the contents of the flask for three minutes and then leave for one hour.
- Filter the mixture.
- Determine the final concentration of the blue dye, $[X]$, by colorimetry.
- Repeat the procedure using different masses of activated charcoal.

(a) The final concentrations of blue dye after carrying out the procedure, $[X]$, are shown in the table.

(i) Process the results to complete the table.

Record your data to **two** decimal places.

mass of activated charcoal, m / g	initial concentration of blue dye / $g\,dm^{-3}$	final concentration of blue dye, $[X]$ / $g\,dm^{-3}$	difference in concentration of blue dye, D / $g\,dm^{-3}$	$\frac{D}{m}$	$\log\left(\frac{D}{m}\right)$	$\log[X]$
0.20	25.00	0.96	24.04	120.20	2.08	-0.02
0.25	25.00	0.69	24.31	97.24	1.99	-0.16
* 0.30	25.00	0.60	24.40	81.33	1.91	-0.22
0.35	25.00	0.41	24.59	70.26	1.85	-0.39
0.40	25.00	0.33	24.67	61.68	1.79	-0.48
0.45	25.00	0.27	24.73	54.96	1.74	-0.57
0.50	25.00	0.23	24.77	49.54	1.69	-0.64
0.55	25.00	0.20	24.80	45.09	1.65	-0.70
0.60	25.00	0.17	24.83	41.38	1.62	-0.77

[3]

(ii) By considering the data in the first three columns, state the effect of increasing the mass of activated charcoal, m, on the amount of adsorption that occurs.

greater adsorption

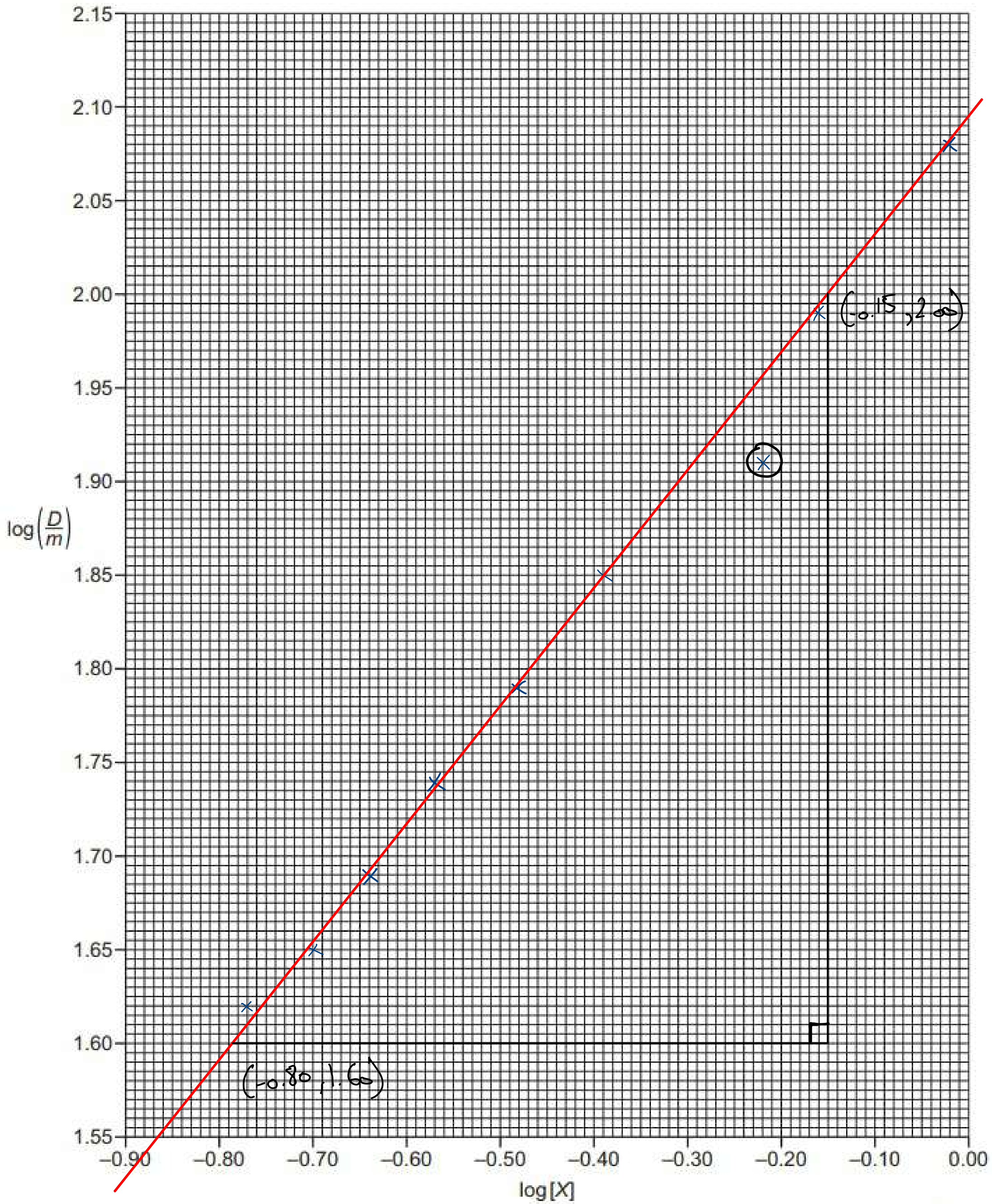
Explain this effect.

greater surface area available

[2]

(b) Plot a graph on the grid to show the relationship between $\log\left(\frac{D}{m}\right)$ and $\log[X]$.

Use a cross (×) to plot each data point. Draw the straight line of best fit.



[2]

(c) Circle the most anomalous point on the graph.

Suggest a reason why this anomaly may have occurred during the experimental procedure.

not enough stirring

surface area not high enough

not left for long enough time [2]

(d) **(i)** Use the graph to determine the gradient of the best-fit line. State the co-ordinates of both points you used in your calculation.

Determine the value of the constant *b*.

co-ordinates 1 (-0.80, 1.60) co-ordinates 2 (-0.15, 2.00)

$$\frac{2.00 - 1.60}{-0.15 + 0.80}$$

gradient = 0.615

b = 0.615 [2]

(ii) Use the graph to determine a value for *A*.

A = 2.095 [1]

[Total: 12]