

Name:		Centre/Index Number:		Class:	
--------------	--	-----------------------------	--	---------------	--



DUNMAN HIGH SCHOOL

Preliminary Examination

Year 6

H2 CHEMISTRY

Paper 4 Practical

9729/04

24 August 2021
2 hours 30 minutes

Candidates answer on the Question Paper.

READ THESE INSTRUCTIONS FIRST

Write your centre number, index number, name and class at the top of this page.

Give details of the practical shift and laboratory where appropriate, in the boxes provided.

Write in dark blue or black pen.

You may use an HB pencil for any diagrams or graphs.

Do not use staples, paper clips, glue or correction fluid.

Answer **all** questions in the spaces provided on the Question Paper.

The use of an approved scientific calculator is expected, where appropriate.

You may lose marks if you do not show your working or if you do not use appropriate units.

Qualitative Analysis Notes are printed on pages 23 and 24.

At the end of the examination, fasten all your work securely together.

The number of marks is given in brackets [] at the end of each question or part question.

Shift
Laboratory

For Examiner's Use	
1	22
2	5
3	15
4	13
Total	55

This question paper consists of **18** printed pages.

Answer **all** questions in the spaces provided.

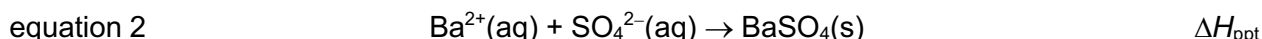
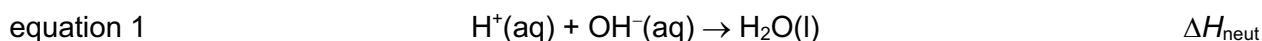
1 Determination of a value for the enthalpy change of precipitation, ΔH_{ppt} , of barium sulfate

FA 1 is 1.00 mol dm⁻³ sodium hydroxide, NaOH.

FA 2 is a saturated solution of barium hydroxide, Ba(OH)₂.

FA 3 is 1.00 mol dm⁻³ sulfuric acid, H₂SO₄.

The addition of **FA 2** to **FA 3** results in neutralisation and precipitation reactions occurring as shown in equations 1 and 2 respectively.



In the first experiment, you will determine the maximum temperature change as a result of both reactions occurring in a Styrofoam cup.

In the second experiment, you will measure the temperature of the resulting mixture after each addition of a fixed volume of **FA 1** solution to the remaining H⁺ ions in the Styrofoam cup. You will analyse your results graphically in order to determine an accurate value for the temperature change of the mixture, caused by the neutralisation reaction shown in equation 1.

You will use this value to calculate the heat change for the second experiment and hence determine a value for the enthalpy change of neutralisation which will then be used to determine a value for the enthalpy change of precipitation for BaSO₄.

(a) (i) Experiment 1

1. Use a measuring cylinder to add 20.0 cm³ of **FA 3** into a Styrofoam cup. Place this cup inside a second Styrofoam cup, which is placed in a 250 cm³ glass beaker. Stir and measure the temperature of this solution, $T_{1,\text{initial}}$. Record this temperature in Table 1.1.
2. Use another measuring cylinder to add 5.0 cm³ of **FA 2** into the Styrofoam cup. Using the thermometer, stir the mixture and record its maximum temperature, $T_{1,\text{max}}$, in Table 1.1. You should expect a small rise in temperature.

Retain the contents in the Styrofoam cup for experiment 2 in (b)(i).

Complete Table 1.1 by calculating the value for $\Delta T_{1,\text{max}}$.

Table 1.1

$T_{1,\text{initial}} / ^\circ\text{C}$	30.0
$T_{1,\text{max}} / ^\circ\text{C}$	31.0
$\Delta T_{1,\text{max}} / ^\circ\text{C}$	+1.0

[1]

(ii) Calculate the heat change, q_1 , for experiment 1.

You should assume that the specific heat capacity of the solution is 4.18 J g⁻¹ K⁻¹, and that the density of the solution is 1.00 g cm⁻³.

[1]

$$q_1 = mc\Delta T = (25)(4.18)(\Delta T_{1,\max}) = w \text{ J (3sf)}$$

 $q_1 = \dots\dots\dots$
Marker's Comments

This question was answered well by the majority. Some students identified the total volume of solution wrongly.

- (iii) Calculate the percentage error associated with the value of $\Delta T_{1,\max}$ and hence comment on the accuracy of the heat change, q_1 , calculated in (a)(ii).

[1]

$$\% \text{ error} = \pm \frac{0.1 + 0.1}{\Delta T_{1,\max}} \times 100 = \pm \text{ ______ } \%$$

Since there is a high percentage error associated with the value of $\Delta T_{1,\max}$, the heat change, q_1 , calculated using $\Delta T_{1,\max}$ is not accurate.

(b) (i) Experiment 2

1. Fill a burette with **FA 1**.
2. Record, in Table 1.2, the initial temperature of the mixture in the Styrofoam cup from experiment 1.
3. Run 5.00 cm³ of **FA 1** from the burette into the Styrofoam cup. Using the thermometer, stir the mixture continuously until it reaches its maximum temperature. Record this temperature, $T_{2,\max}$, in Table 1.2.
4. Repeat step 3 until a total volume of 50.00 cm³ of **FA 1** is added.

Table 1.2

total volume of FA 1 added after each addition, $V_{\text{FA 1}}$ / cm ³	$T_{2,\max}$ / °C
0.00	30.8
5.00	32.8
10.00	34.2
15.00	35.4
20.00	36.2
25.00	37.0
30.00	37.8
35.00	38.4
40.00	38.4
45.00	37.8
50.00	37.2

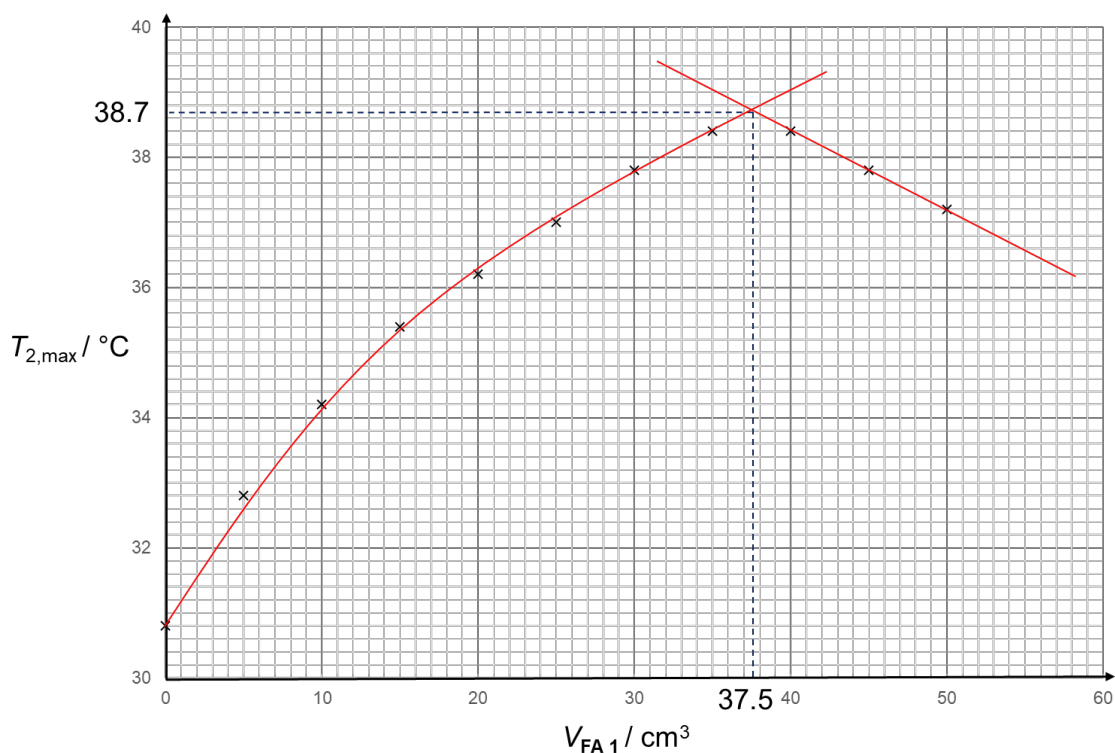
[2]

(ii) Plot a graph of $T_{2,\max}$ on the y-axis against $V_{\text{FA } 1}$ on the x-axis.

Draw two best-fit lines,

- the first is a smooth **curve** taking into account all of the points before,
- the second is a **straight line** taking into account all of the points after the temperature of the mixture has started to drop.

Extrapolate (extend) both lines until they intersect (cut) each other.



[3]

(iii) From your graph, read $V_{\text{FA } 1}$ and $T_{2,\max}$ of the intersection point.

Record these values in the spaces provided.

Deduce the maximum temperature change, $\Delta T_{2,\max}$.

$V_{\text{FA } 1} = \dots\dots\dots \text{cm}^3$

$T_{2,\max} = \dots\dots\dots ^\circ\text{C}$

$\Delta T_{2,\max} = \dots\dots\dots ^\circ\text{C}$

$$V_{\text{FA } 1} = 37.50 \text{ cm}^3$$

$$T_{2,\max} = 38.7 ^\circ\text{C}$$

$$\Delta T_{2,\max} = 38.7 - 30.8 = +7.9 ^\circ\text{C}$$

[2]

- (iv) Calculate the heat change, q_2 , for experiment 2 using the values you deduced in (b)(iii).

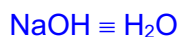
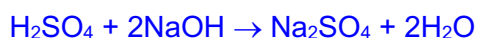
You should assume that the specific heat capacity of the solution is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$, and that the density of the solution is 1.00 g cm^{-3} .

$$q_2 = \dots\dots\dots$$

$$q_2 = mc\Delta T = (25 + V_{\max})(4.18)(\Delta T_{2,\max}) = x \text{ J (3sf)}$$

[1]

- (v) Hence determine a value for the enthalpy change, for the neutralisation reaction shown in equation 1, ΔH_{neut} .



$$\text{Moles of H}_2\text{O formed} = \text{moles of NaOH added} = V_{\max} \times 10^{-3} \times 1 = y \text{ mol}$$

$$\Delta H_{\text{neut}} = -\frac{x}{y} \times 10^{-3} = -z \text{ kJ mol}^{-1}$$

$$\Delta H_{\text{neut}} = \dots\dots\dots$$

[2]

- (c) (i) The **FA 2** solution was prepared as follows.

An excess of solid $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ was added to 100 cm^3 of water and the mixture was allowed to stand. The undissolved solid was then removed by filtration to obtain the saturated solution of barium hydroxide.

Calculate the concentration of barium and hydroxide ions in the **FA 2** solution given that the solubility of $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ in water is 56 g dm^{-3} .

$$[M_r: \text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}, 315.3]$$

[1]

$$\text{solubility of Ba(OH)}_2 \cdot 8\text{H}_2\text{O} = \frac{56}{315.3} = 0.17761 \text{ mol dm}^{-3}$$

$$[\text{Ba}^{2+}] \text{ in FA 2} = 0.178 \text{ mol dm}^{-3}$$

$$[\text{OH}^-] \text{ in FA 2} = 0.17761 \times 2 = 0.355 \text{ mol dm}^{-3}$$

$$[\text{Ba}^{2+}] = \dots\dots\dots$$

$$[\text{OH}^-] = \dots\dots\dots$$

- (ii) Hence calculate the heat evolved from the neutralisation reaction between **FA 2** and **FA 3** in experiment 1.

[1]

$$\text{moles of water formed} = \text{moles of OH}^- \text{ neutralised}$$

$$= 0.35522 \times 0.005 = 0.0017761 \text{ mol}$$

$$\text{Heat evolved from neutralisation} = 0.0017761 \times |\Delta H_{\text{neut}}| \times 10^3 = r \text{ J}$$

- (iii) Use your answers in (a)(ii) and (c)(ii) to determine the heat change due to the precipitation of BaSO_4 in experiment 1. Hence determine a value for the enthalpy change for the precipitation reaction shown in equation 2, ΔH_{ppt} .

[3]

Heat evolved from precipitation = q_1 – heat evolved from neutralisation = s J

Heat change of solution due to the precipitation of BaSO_4 = $+s$ J

moles of BaSO_4 = moles of Ba^{2+} = $0.17761 \times 0.005 = 8.8804 \times 10^{-4}$ mol

$$\Delta H_{\text{ppt}} = -\frac{\text{heat evolved from precipitation}}{8.8804 \times 10^{-4}} \times 10^{-3} = -p \text{ kJ mol}^{-1}$$

Heat change due to the precipitation of BaSO_4 =

$\Delta H_{\text{ppt}} = \dots\dots\dots$

- (iv) Explain, in terms of the chemistry involved, the sign of ΔH_{ppt} . Hence explain why precipitation of BaSO_4 occurs spontaneously under the experimental conditions.

Energy is released on formation of ionic bonds between the oppositely charged ions, Ba^{2+} and SO_4^{2-} , in the solid ionic compound, BaSO_4 .

$\Delta S < 0$. There is a decrease in disorder of the system as the ions in the solid ionic lattice structure are more orderly than the aqueous ions dispersed in the solution.

$\Delta G_{\text{ppt}} = \Delta H_{\text{ppt}} - T\Delta S_{\text{ppt}}$. Although $-T\Delta S_{\text{ppt}} > 0$, magnitude of $T\Delta S_{\text{ppt}}$ is smaller than that of ΔH_{ppt} such that $\Delta G_{\text{ppt}} < 0$ and the precipitation occurs spontaneously.

[2]

- (d) A student performed experiment 2 using 2.00 mol dm^{-3} sodium hydroxide from the bench reagents instead of **FA 1**.

Explain how the volume of 2.00 mol dm^{-3} sodium hydroxide required for complete neutralisation of the mixture in the Styrofoam cup would differ from that of **FA 1**.

[2]

The volume of 2.00 mol dm^{-3} sodium hydroxide required for complete neutralisation will be half that of **FA 1**.

The $[\text{OH}^-]$ in 2.00 mol dm^{-3} sodium hydroxide is twice that of **FA 1**. Half the volume of 2.00 mol dm^{-3} sodium hydroxide is required to provide the same moles of OH^- for neutralisation of H^+ in the mixture.

[Total: 22]

2 Investigation of pH of solution

FA 4 is an organic compound with the molecular formula, $C_2H_4O_2$.

Perform the tests described in Table 2.1. Some of the observations have been completed for you. There is no need to carry out those tests. Record your observations in Table 2.1.

Unless otherwise stated, the volumes given below are approximate and should be estimated rather than measured. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 2.1

		test	observations
(a)	(i)	Add 4 cm depth of deionised water to a test-tube and test it using Universal Indicator (UI) paper. Add 2 drops of FA 1 to the test-tube and test the solution using Universal Indicator paper.	UI paper turns light green pH 5 UI paper turns dark blue pH 12
	(ii)	Add 3 cm depth of FA 4 to a test-tube. Add 1 cm depth of aqueous sodium hydroxide to the test-tube and test the solution using Universal Indicator paper. Add 2 drops of FA 1 to the test-tube and test the solution using Universal Indicator paper.	UI paper turns light orange pH 4 UI paper turns light orange pH 4
	(iii)	Add 2 cm depth of FA 4 to a test-tube. Add half a spatula of solid sodium carbonate to the test-tube.	effervescence observed gas evolved formed white ppt with limewater the gas is carbon dioxide

[2]

- (b) (i)** Suggest the identity of **FA 4**. Give evidence from your observations in **(a)** to support your conclusion.

identity

evidence

[1]

FA 4 = ethanoic acid

FA 4 contains a carboxylic acid functional group because acid-carbonate reaction occurred in (a)(iii) to produce CO_2 gas.

- (ii) Explain, in terms of the chemistry involved, the difference in observations made upon the addition of **FA 1** to the test-tubes in (a)(i) and (a)(ii).

[2]

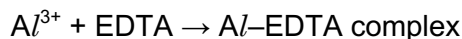
In (a)(i), addition of FA 1 containing OH^- to water produced a basic solution and thus there was a large increase in pH.

In (a)(ii), addition of NaOH(aq) partially neutralised FA 4 to give an acidic buffer solution containing CH_3COOH and $\text{CH}_3\text{COO}^-\text{Na}^+$ which resisted changes in pH when a small amount of FA 1 containing OH^- was added. The pH of the solution remains relatively constant.

[Total: 5]

3 Determination of amount of Mg^{2+} and Al^{3+} in an antacid tablet

To determine the amount of Mg^{2+} and Al^{3+} in an antacid tablet, a complexometric back titration involving ethylenediaminetetraacetic acid, EDTA, can be used. Both Mg^{2+} and Al^{3+} form a complex with EDTA in a 1:1 mole ratio.



A known amount of excess EDTA is first added to a sample containing Mg^{2+} and Al^{3+} . The resultant mixture contains both the complex and unreacted EDTA. This resultant mixture is then titrated with zinc sulfate to determine the amount of unreacted EDTA. Zn^{2+} also reacts with EDTA in a 1:1 mole ratio.



The dye indicator used in this titration is Eriochrome Black T, which will turn from blue to violet at the end-point.

When the titration is conducted at pH 5, only the Al-EDTA complex forms, while at pH 10, both the Mg-EDTA and Al-EDTA complexes form.

You will only be performing the titrations at pH 10, to determine the total amount of Mg^{2+} and Al^{3+} in a sample.

FA 5 is a solution prepared containing Mg^{2+} and Al^{3+}

FA 6 is a buffer solution at pH 10, containing Na_2CO_3 and NaHCO_3

FA 7 is $0.0100 \text{ mol dm}^{-3}$ EDTA

FA 8 is $0.0100 \text{ mol dm}^{-3}$ ZnSO_4

Solution T is Eriochrome Black T indicator

As EDTA is harmful to the environment, **FA 7** should be disposed in the waste bottle. You should also wear gloves throughout the experiment.

(a) Pre-titration: Determination of the colour at end-point for titration

1. Using a dropping pipette, add the following solutions into a clean boiling tube:

- 6 drops of **FA 5**
- 15 drops of **FA 6**
- 20 drops of **FA 7**
- 1 drops of **Solution T**

The colour of the solution in the test tube should be blue. If the colour of the solution is not blue, add a few more drops of **FA 6** and **FA 7** into the boiling tube.

2. Using a dropping pipette, add **FA 8** dropwise into the test tube, with shaking, until one drop of **FA 8** causes the blue colouration of the solution to fade to yield a violet colour.
3. Keep this solution as a reference for the colour at the end-point of the titration in (b).

(b) (i) Determination of the amount of Mg^{2+} and Al^{3+} in the sample

1. Fill a burette with **FA 8**.
2. Using a pipette, transfer 10.0 cm^3 of **FA 5** into a 250 cm^3 conical flask.
3. Using appropriate measuring cylinders, add 20.0 cm^3 of **FA 6**, followed by 35.0 cm^3 of **FA 7** into the conical flask.
4. Swirl the conical flask to ensure a homogeneous solution. It is normal for the solution to appear cloudy.
5. Heat the conical flask over the Bunsen burner until the temperature of the solution reaches 65°C .
6. Remove the conical flask from the flame. If the neck of the flask is too hot to hold safely, use a folded paper towel to hold the flask.
7. Add 5 drops of **Solution T** into the conical flask. The solution should be blue at this point.
8. Run **FA 8** from the burette into the conical flask.
9. The end-point is reached when the blue colour fades to yield the violet colour of the solution in the boiling tube from **(a)**.
10. Record your titration results, to an appropriate level of precision, in Table 3.1.

Titration results**Table 3.1**

	1	2
Final burette reading / cm^3	14.90	29.80
Initial burette reading / cm^3	0.00	15.00
Volume of FA 8 used / cm^3	14.90	14.80

[3]

- (ii)** From your titrations, obtain a suitable volume of **FA 8**, $V_{\text{FA 8}}$, to be used in your calculations. Show clearly how you obtained this volume.

[1]

$$V_{\text{FA 8}} = \frac{14.90 + 14.80}{2} = 14.85 \text{ cm}^3$$

- (c) (i)** Using your answer in **(b)(ii)**, calculate the amount of unreacted EDTA present in the conical flask after step 3.

[1]



$$\begin{aligned} \text{No. of moles of unreacted EDTA} &= \frac{14.85}{1000} \times 0.0100 \\ &= 1.485 \times 10^{-4} \text{ mol} = 1.49 \times 10^{-4} \text{ mol (to 3 s.f.)} \end{aligned}$$

- (ii) Calculate the amount of EDTA that formed a complex with Mg^{2+} and Al^{3+} , and hence, determine the total concentration of Mg^{2+} and Al^{3+} in **FA 5**. [3]

$$\text{Total moles of EDTA added} = \frac{35.00}{1000} \times 0.0100 = 3.50 \times 10^{-4} \text{ mol}$$

$$\text{No. of moles of EDTA reacted} = (3.50 - 1.485) \times 10^{-4} \text{ mol} = 2.015 \times 10^{-4} \text{ mol}$$

Since both Mg^{2+} and Al^{3+} form a complex with EDTA in a 1:1 mole ratio,

$$\text{Total moles of } \text{Mg}^{2+} \text{ and } \text{Al}^{3+} = 2.015 \times 10^{-4} \text{ mol}$$

$$\text{Total conc. of } \text{Mg}^{2+} \text{ and } \text{Al}^{3+} \text{ in FA 5} = 2.015 \times 10^{-4} \div \frac{10}{1000} = 0.0202 \text{ mol dm}^{-3} \text{ (3 sf)}$$

- (d) Briefly outline how you can determine the exact individual amounts of Mg^{2+} and Al^{3+} in **FA 5** by conducting a second titration involving ZnSO_4 and EDTA.

You may assume that only Mg^{2+} and Al^{3+} in **FA 5** react with EDTA. [2]

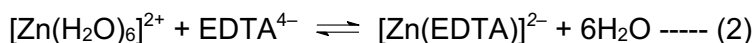
Repeat the same procedure but conduct the titration at pH 5. This will allow us to determine the amount of Al^{3+} in FA 5.

Subtract this amount from the total amount calculated in (c)(ii) to obtain the amount of Mg^{2+} in FA 5.

- (e) EDTA is better represented as H_4EDTA because it is a weak acid. The EDTA^{4-} anion can be formed from four successive deprotonation of H_4EDTA .



At high pH, Zn^{2+} readily forms a complex with EDTA^{4-} .



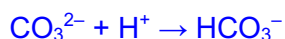
- (i) Explain the effect of increasing pH on the concentration of the $[\text{Zn}(\text{EDTA})]^{2-}$ complex. [2]

As pH increases, the concentration of H^+ in solution decreases.

By Le Chatelier's Principle, the position of equilibrium (1) shifts right to increase the concentration of H^+ . This results in an increase in concentration of EDTA^{4-} present in the solution.

As the concentration of EDTA^{4-} in solution increases, by Le Chatelier's Principle, position of equilibrium (2) shifts right to decrease the concentration of EDTA^{4-} . This results in an increase in the concentration of complex formed.

- (ii) The reaction mixture is kept at high pH due to the Na_2CO_3 and NaHCO_3 buffer. With the aid of an appropriate chemical equation, explain how this buffer maintains the high pH when a small amount of acid is added. [2]



The added H^+ is removed as HCO_3^- . Hence, $[\text{H}^+]$ is relatively constant and the pH is maintained at a high pH level.

- (iii) Given that the pK_a of HCO_3^- is 10.3, calculate the $\frac{[\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$ ratio in the $\text{Na}_2\text{CO}_3/\text{NaHCO}_3$ buffer solution at pH 10.

$$\text{pH} = pK_a + \log \frac{[\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$$

$$10 = 10.3 + \log \frac{[\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$$

$$\text{Hence, } \frac{[\text{CO}_3^{2-}]}{[\text{HCO}_3^-]} = 10^{-0.3} = 0.501$$

[Total: 15]

4 Planning

In the presence of glucose, acidified potassium manganate(VII) decolourises. This is because glucose acts as a reducing agent, reducing MnO_4^- to Mn^{2+} .

A series of experiments can be carried out at various temperatures to investigate the effect of temperature on the rate constant of the reaction. The time required for the purple reaction mixture to turn colourless will allow for the determination of the rate of reaction.

- (a) (i) State the effect of an increase in temperature on the time taken for the purple reaction mixture to turn colourless.

[1]

An increase in temperature will result in a faster rate, and hence will result in a shorter time for the solution to turn colourless.

- (ii) Using the concept of Collision Theory, explain your answer in (a)(i).

[2]

The increase in temperature results in an increase in average kinetic energy of all reactant particles. More reactant particles possess energy more than or equal to activation energy, resulting in an increase in the number of effective collisions per unit time. Since rate of reaction is proportional to the frequency of effective collisions, the rate of reaction increases, resulting in a shorter reaction time.

- (b) To investigate the effect of temperature on the rate constant, a series of experiments at different temperatures can be carried out.

You may assume that you are provided with:

- 0.1 mol dm^{-3} glucose solution
- 0.1 mol dm^{-3} aqueous potassium manganate(VII), KMnO_4 ,
- 2 mol dm^{-3} sulfuric acid, H_2SO_4 ,
- the apparatus normally found in a school laboratory.

Table 4.1 shows the volumes of reactants used for experiment 1.

Once all the reactants have been added to a dry conical flask, the initial temperature of the reaction mixture, T_i , was measured and recorded. Immediately after the decolourisation of KMnO_4 , the final temperature of the mixture, T_f , was also measured and recorded. The average temperature of the reaction mixture, T_{ave} , was then determined.

Table 4.1 also shows the time required for the decolourisation of purple KMnO_4 for experiment 1.

Table 4.1

Experiment	Vol. of glucose solution / cm^3	Vol. of KMnO_4 / cm^3	Vol. of H_2SO_4 / cm^3	T_i / $^{\circ}\text{C}$	T_f / $^{\circ}\text{C}$	T_{ave} / $^{\circ}\text{C}$	Time taken for decolourisation / s
1	5.0	x	20.0	29.0	30.0	29.5	240
2							

- (i) Given that glucose and KMnO_4 react in a molar ratio of 5 : 24, state an appropriate value for x, the volume of KMnO_4 used in experiment 1.

[1]

24 cm³

To investigate the effect of temperature on the rate constant, experiment 2 can be conducted at a different temperature such that the time taken for the decolourisation of purple KMnO₄ will be more than 240 s.

- (ii) Fill Table 4.1 with the volumes of glucose, KMnO₄ and H₂SO₄ needed for experiment 2.

[1]

Use exactly the same volumes as experiment 1

- (iii) Explain your choice of reactant volumes used in (b)(ii).

[1]

By using the same volumes as experiment 1, the concentrations of reactants in the reaction mixture are kept constant. Hence, the only variable that has changed is the temperature.

- (iv) You are required to write a plan, describing how **experiment 2** can be carried out.

In your plan, you should have details of:

- the apparatus you would use,
- the measurements you would take,
- the procedure you would follow.

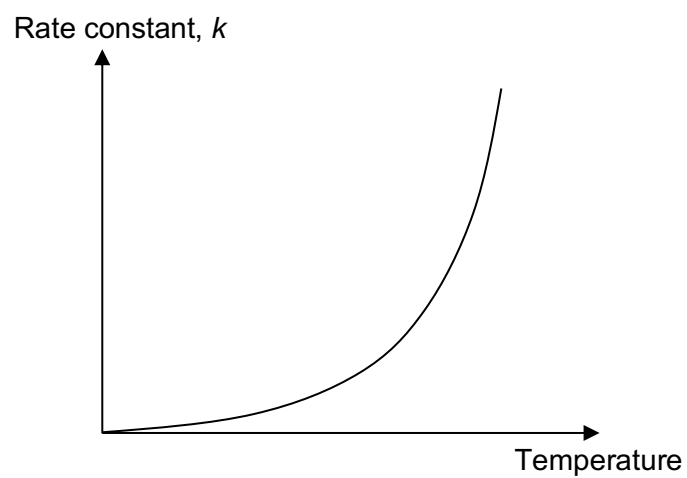
[6]

1. Using a 10 cm³ measuring cylinder, measure out 5.0 cm³ of the glucose solution.
2. Using another burette, add x cm³ of aqueous KMnO₄ into a dry 100 cm³ conical flask / beaker.
3. Using a 25 cm³ measuring cylinder, add 20.0 cm³ H₂SO₄ into the conical flask containing KMnO₄ and swirl the solution.
4. Place the conical flask in a thermostatic water bath set at 15 °C (or any temperature below 29.5 °C) and use a thermometer to measure the temperature of the solution in the conical flask.
5. Once the temperature of the solution reaches 15 °C, add the glucose solution from the measuring cylinder into the conical flask and start the stopwatch immediately.
6. Use the thermometer to measure the temperature of the reaction mixture once the stopwatch has been started. This is the initial temperature of the reaction mixture.
7. Stop the stopwatch once the reaction mixture turns colourless.
8. Record the time taken.
9. Use the thermometer to measure the temperature of the reaction mixture once it turns colourless. This is the final temperature of the reaction mixture.

- (v) The rate of reaction approximately doubles for every 10 °C increase in temperature.

Using the axes below, sketch a graph to show how the rate constant of the reaction would vary with temperature.

[1]



[Total: 13]