Catholic Junior College

JC2 Preliminary Examinations

Higher 2

CANDIDATE NAME	
CLASS	2T

CHEMISTRY

9729/04

Paper 4 Practical

23 August 2023
2 hours 30 minutes

Candidates answer on the Question Paper.
Additional Materials: As listed in the Confidential Instructions

READ THESE INSTRUCTIONS FIRST

Write your name and class in the boxes above.
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 21 and 22.

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	**/ 11**
2	**/ 18**
3	**/ 16**
4	**/ 10**
Total	**/ 55**

MARK SCHEME

This document consists of **21** printed pages and **1** blank page.

[Turn over

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

1 Qualitative analysis of an inorganic solid mixture FA 8 is $MnCO_3 + Al(NO_3)_3$

FA 8 is a solid mixture that contains two cations and two anions listed in the Qualitative Analysis Notes on pages 21 and 22.

Unless otherwise stated, the volumes given below are approximate and should be estimated rather than measured.

1. Transfer the **FA 8** provided into a boiling tube and add in 3 cm depth of deionised water.
2. Using a glass rod, stir the mixture well for one minute until no more of the solid mixture can dissolve.
3. Filter the mixture into another clean boiling tube. **While waiting, proceed to Question 2 first**.
4. When the filtration is completed, place the filter funnel containing the residue on a conical flask and **wash the residue** with a little deionised water.

Keep both filtrate and residue for tests (a) and (b) below.

Test and identify any gases evolved.

At each stage of any test, you are to record details of the following:

- colour changes seen
- the formation of any precipitate
- the solubility of such precipitates in an excess of the reagent added

If there is no observable change, write **no observable change**.

Record your observations in Tables 1.1 and 1.2 below.

No additional tests for ions present should be attempted.

(a) Tests on the filtrate.

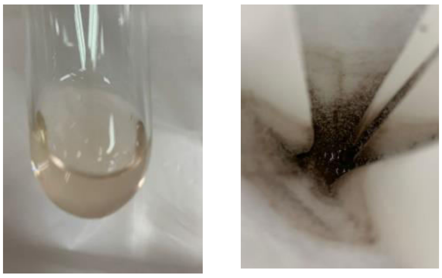
Table 1.1

	test	*observations*
(i)	Test the filtrate using Universal indicator paper. Hence conclude pH of the filtrate.	• **Universal indicator paper turns yellow.** **pH 3 to 5**

(ii)	To 1 cm depth of the filtrate in a test-tube, add aqueous sodium hydroxide dropwise, with shaking, until no further change is observed.	- **Off-white / light brown ppt formed; insoluble in excess NaOH**

(b) **Tests on the residue.**

Table 1.2

	test	*observations*
(i)	Place the funnel containing the residue (on the filter paper) into a clean boiling tube. Pour about 5 cm^3 of dilute nitric acid onto the residue. Collect 1 cm depth of the solution in the boiling tube. Remove the funnel and return it to the conical flask.	- **Effervescence observed; gas gives white ppt with limewater. $CO_2(g)$ liberated.** - **Colourless filtrate and dark brown residue obtained.** 

[Turn over

(ii)	To this solution in the boiling tube, add aqueous sodium hydroxide dropwise, with shaking, until no further change is observed. Keep this resultant mixture for test **(iii)** below.	• **Off-white / beige / pale or light brown/cream ppt insoluble in excess NaOH(aq).**  • **Ppt rapidly turns brown on contact with air / darkens on standing.** 
(iii)	To 1 cm depth of the mixture from **(ii)**, carefully add 1 cm depth of aqueous hydrogen peroxide.	• **ppt turns brown / darker brown / brown-black.** • **Brisk effervescence: gas re-ignite a glowing splint; O_2 gas liberated.**  

[6]

(c) **(i)** One of the cations in **FA 8** is responsible for the nature of the filtrate in test **1(a)(i)**. It is formed from a Period 3 element.

This cation present in **FA 8** is: **Al^{3+}** [1]

(ii) Write an equation to account for the nature of the filtrate in test **1(a)(i)**.

$$[Al(H_2O)_6]^{3+} + H_2O \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} + H_3O^+$$ [1]

(d) **(i)** The second cation present in **FA 8** is **Mn^{2+}**

evidence **In test 1b(ii), addition of NaOH (aq) results in formation of a off-white ppt that was insoluble in excess NaOH(aq). The ppt rapidly turned brown upon contact with air, indicating presence of Mn^{2+}.**

……………………………………………………………………………………………

……………………………………………………………………………………………[1]

(ii) **One** anion present in **FA 8** is CO_3^{2-}

evidence In test 1b(i), addition of nitric acid results in brisk effervescence of CO_2 gas which gave a white ppt with $Ca(OH)_2$, indicating presence of CO_3^{2-}.

……………………………………………………………………………………………

……………………………………………………………………………………………

[1]

(iii) Suggest what type of reaction takes place when hydrogen peroxide is added in test **1(b)(iii)**.

Redox ……………………………………………………………………………………[1]

[Total: 11]

2 **Determination of the molar enthalpy changes of two reaction by indirect methods.**

Sodium carbonate, Na_2CO_3 reacts with water and carbon dioxide to form sodium hydrogencarbonate, $NaHCO_3$, according to equation 1.

equation 1 $Na_2CO_3(s) + H_2O(l) + CO_2(g) \rightarrow 2NaHCO_3(s)$ ΔH_1

The enthalpy change of this reaction cannot be determined directly.

Both $Na_2CO_3(s)$ and $NaHCO_3(s)$ react with dilute sulfuric acid.

equation 2 $Na_2CO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$ ΔH_2
equation 3 $NaHCO_3(s) + ½\ H_2SO_4(aq) \rightarrow ½\ Na_2SO_4(aq) + H_2O(l) + CO_2(g)$ ΔH_3

In this experiment, you will determine the values for ΔH_2 and ΔH_3, and then use your results to calculate a value for ΔH_1.

You are provided with:
FA 1 is sodium carbonate, Na_2CO_3.
FA 2 is sodium hydrogen carbonate, $NaHCO_3$.
FA 3 is 1.00 mol dm^{-3} sulfuric acid, H_2SO_4 (**FA 3** will also be used in **Q3**).

(a) Method

Experiment 1: $Na_2CO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

1. Use a measuring cylinder to transfer 25 cm^3 of the acid, **FA 3**, into a polystyrene cup supported in a 250 cm^3 beaker. The acid is in excess.
2. Weigh the capped bottle containing **FA 1** and record the balance reading.
3. Place the thermometer in the acid and record its initial temperature.
4. Carefully tip all the **FA 1**, in small portions, into the acid and stir to dissolve.
5. Read and record the highest temperature reached.
6. Reweigh the emptied bottle and its cap and record the balance reading and the mass of **FA 1** used.

Experiment 2: $NaHCO_3(s) + ½\ H_2SO_4(aq) \rightarrow ½\ Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

1. Replace the wet polystyrene cup with a clean, dry polystyrene cup.
2. Repeat Steps 1 to 4 in **Experiment 1** but use **FA 2** in place of **FA 1**.
3. Read and record the lowest temperature reached.
4. Reweigh the emptied bottle and its cap and record the balance reading and the mass of **FA 2** used.

Results

Record **all** weighings and temperature readings to an appropriate level of precision in the table below.

Table 2.1

	Experiment 1	**Experiment 2**
mass of capped bottle + **solid** / g	**7.599**	**10.625**
mass of capped bottle + residual **solid** / g	**5.607**	**8.129**
mass of **solid** added / g	**1.992**	**2.496**
initial temperature, T_i, of **FA 3** / °C	**28.0**	**28.0**
final temperature, T_f / °C	**34.0**	**20.3**
temperature change, ΔT / °C	**6.0**	**–7.7**

[1] all mass recorded consistently to 2 or 3 d.p. + all temperature readings recorded to 1 d.p. + ΔT and mass of solid used correctly calculated.

[4] Accuracy:

Experiment 1: compare $\frac{\Delta T}{\text{mass of FA 1}}$ (to 3 s.f.) with trs'

(2) if difference, $\delta \le 0.30$ °C g^{-1}

(1) if difference, $0.30 < \delta \le 0.60$ °C g^{-1}

(0) if difference, $\delta > 0.60$ °C g^{-1}

Experiment 2: compare $\frac{\Delta T}{\text{mass of FA 2}}$ (to 3 s.f.) with trs'

(2) if difference, $\delta \le 0.30$ °C g^{-1}

(1) if difference, $0.30 < \delta \le 0.60$ °C g^{-1}

(0) if difference, $\delta > 0.60$ °C g^{-1}

[5]

(b) For the purpose of calculations, 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 1.0 g cm^{-3}.

(i) Use relevant results from **Table 2.1** to calculate the heat energy produced when **FA 1** was added to the acid in **Experiment 1**. Hence, calculate the enthalpy change, in kJ mol^{-1}, when 1 mol of **FA 1**, Na_2CO_3, reacts with the acid. Include the sign of ΔH_2 in your answer.
[A_r: C, 12.0; O, 16.0; Na, 23.0].

heat evolved = mc ΔT

= 25 × 4.18 × 6.0

= 627 J

heat energy produced =……**627**……. J

M_r of Na_2CO_3 = 2(23.0) + 12.0 + 3(16.0) = 106.0

mol of Na_2CO_3 = $\frac{1.992}{106.0}$ = 0.0188 mol

$$\therefore\ \Delta H_2 = -\frac{627}{0.0188}\ \text{J mol}^{-1}$$

$$= -33350\ \text{J mol}^{-1}$$

$$= -33.4\ \text{kJ mol}^{-1}$$

enthalpy change, ΔH_2 = ...**–33.4**...... kJ mol^{-1} [3]

(ii) Use relevant results from **Table 2.1** to calculate the heat energy absorbed when **FA 2** was added to the acid in **Experiment 2**. Hence, calculate the enthalpy change, in kJ mol^{-1}, when 1 mol of **FA 2**, $NaHCO_3$, reacts with the acid. Include the sign of ΔH_3 in your answer.
[A_r: H, 1.0; C, 12.0; O, 16.0; Na, 23.0]

heat absorbed = mc ΔT
= 25 × 4.18 × 7.7
= 804.6 J

heat energy absorbed =...**804.6**...... J

M_r of $NaHCO_3$ = 23.0 + 1.0 + 12.0 + 3(16.0) = 84.0

mol of $NaHCO_3$ = $\frac{2.496}{84.0}$ = 0.0297 mol

$$\therefore\ \Delta H_3 = +\frac{804.6}{0.0297}\text{ J mol}^{-1} = +27090\text{ J mol}^{-1} = +27.1\text{ kJ mol}^{-1}$$

enthalpy change, ΔH_3 = ...**+27.1**...... kJ mol^{-1} [3]

(iii) Using your answers to **(b)(i)** and **(b)(ii)** and the equations for **Experiment 1** and **Experiment 2**, construct an energy cycle to determine a value for the enthalpy change for the reaction:

$$\begin{tikzcd}[column sep=small]
Na_2CO_3(s) + H_2O(l) + CO_2(g) \arrow[rr, "\Delta H_1 = ?"] \arrow[dr, "+ H_2SO_4(aq)"', "\Delta H_2"] & & 2NaHCO_3(s) \arrow[dl, "+ H_2SO_4(aq)", "2\Delta H_3"'] & \Delta H_1 \\
& Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g) & &
\end{tikzcd}$$

By Hess' Law,

$$\Delta H_1 = \Delta H_2 - 2\Delta H_3 = (-33.4) - 2(+27.1) = -87.6\text{ kJ mol}^{-1}$$

–87.6

enthalpy change, ΔH_1 = ... kJ mol^{-1} [2]

(c) An additional experiment must be carried out using **FA 2**, solid sodium hydrogencarbonate, $NaHCO_3$, and deionised water to determine the enthalpy change of solution of $NaHCO_3$, ΔH_4.

(i) Carry out the additional experiment using between 2.0 to 2.5 g of **FA 2** and 50 cm^3 of deionised water. Record your results, in a suitable form, in the space below.

Mass of weighing bottle / g	4.08
mass of weighing bottle + **FA 2** / g	6.43
mass of weighing bottle + residual **FA 2** / g	4.15
mass of **FA 2** added / g	2.28
initial temperature of water / °C	30.4
lowest temperature reached / °C	28
temperature fall, ΔT / °C	-2.4

[1]

(ii) Use your results in **c(i)** to calculate a value for the enthalpy change of solution of $NaHCO_3$, ΔH_4. Include the sign of ΔH_4 in your answer.

heat absorbed = mc ΔT

= $50 \times 4.18 \times 2.4$

= 502 J

M_r of $NaHCO_3$ = 23.0 + 1.0 + 12.0 + 3(16.0) = 84.0

mol of $NaHCO_3$ = $\frac{2.28}{84.0}$ = 0.0271 mol

$\therefore\ \Delta H_4 = +\frac{502}{0.0271}$ J mol^{-1}

= +18480 J mol^{-1}

= +$\underline{18.5\text{ kJ mol}^{-1}}$

enthalpy change, ΔH_4 = ...**+18.5**.... kJ mol^{-1} [2]

(iii) Hence use your answer in **b(iii)** and **(c)(ii)**, determine the enthalpy change for the following reaction:

$$Na_2CO_3(s) + H_2O(l) + CO_2(g) \rightarrow 2NaHCO_3(aq) \qquad \Delta H_5$$

$$\begin{tikzcd}
Na_2CO_3(s) + H_2O(l) + CO_2(g) \arrow[rr, "\Delta H_5=?"] \arrow[dr, "\Delta H_1"'] & & 2NaHCO_3(aq) \\
& 2NaHCO_3(s) \arrow[ur, "2\Delta H_4"'] &
\end{tikzcd}$$

By Hess' Law,

$\Delta H_5 = \Delta H_1 + 2\Delta H_4$

= (–87.6) + 2(+18.5)

= $\underline{-50.6\text{ kJ mol}^{-1}}$

enthalpy change, ΔH_5 = −50.6 kJ mol^{-1}

[1]

(d) Calculate the percentage error incurred when using the thermometer to measure the temperature change in **Experiment 1**.

% error incurred for temperature reading = $\left(\frac{2\times0.1}{6.0}\right)\times100\%$ = 3.33 %

[1]

[Total: 18]

3 Investigation of reaction between propanone and iodine

FA 3 is 1.00 mol dm^{-3} sulfuric acid, H_2SO_4.
FA 4 is 1.00 mol dm^{-3} propanone, CH_3COCH_3.
FA 5 is an aqueous solution of iodine, I_2.
FA 6 is 0.0100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.
FA 7 is 0.50 mol dm^{-3} sodium hydrogen carbonate, $NaHCO_3$.

You are also provided with starch indicator.

The reaction between propanone and iodine in aqueous solution is shown in **reaction 1**. This reaction is first order with respect to both CH_3COCH_3 and H^+ ions.

reaction 1 $CH_3COCH_3(aq) + I_2(aq) \rightarrow CH_3COCH_2I(aq) + HI(aq)$

You are to investigate how the rate of reaction changes with $[I_2]$. To do this, you will prepare a reaction mixture containing **FA 3**, **FA 4** and **FA 5**. At timed intervals, you will withdraw **five aliquots** (portions) of the reaction mixture and quench the reaction with excess **FA 7**.

You will then titrate the remaining iodine in the resulting solutions against sodium thiosulfate (**reaction 2**).

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

Your titre values at the different times will indicate the concentration of I_2 remaining in the reaction mixture at those different times. Hence, the rate of reaction between CH_3COCH_3 and I_2 at the different times can be determined by graphical analysis of your results.

You should read all of the instructions on this page and the method on the next page before you start this experiment.

Recording your results

In an appropriate format in the space provided on **page 13** under the heading, **Results**, in **(b)**, prepare a table to record, for each of your aliquots, the

- transfer time in minutes and seconds,
- titration results (initial and final burette readings; and volume of **FA 6** added),
- time, which is the transfer time converted to minutes, to one decimal place (e.g. a transfer time of 2 min 27 s becomes 2 min + 27/60 min = 2.5 min).

Make certain that your recorded results show the precision of your working.

[Turn over

(a) Preparation and titration of the reaction mixture

Notes: You will perform each titration **once** only. Great care must be taken that you do not overshoot the end-point.

You should **keep the reaction mixture stoppered** except when removing aliquots. Once you have started the stopwatch, it must continue for the duration of the experiment. You must **not** stop it until you have finished this experiment.

You should aim **not** to exceed a maximum reaction time of 25 minutes for this experiment.

1. Using a 10 cm^3 measuring cylinder, add about 10 cm^3 of **FA 7** to each of the labelled boiling tubes, **1** to **5**.
2. Fill a burette with **FA 6**.
3. Use appropriate measuring cylinders to add to the conical flask labelled **reaction mixture**
   - 25.0 cm^3 of **FA 3**,
   - 25.0 cm^3 of **FA 4**.
4. Place 50.0 cm^3 of **FA 5** into a 50 cm^3 measuring cylinder.
5. Pour **FA 5** into the conical flask labelled **reaction mixture**. Start the stopwatch at the instant of mixing. **Insert the stopper** and swirl the mixture thoroughly to mix its contents.
6. At approximately 4 minutes, transfer a 10.0 cm^3 aliquot (portion) of the reaction mixture into the boiling tube labelled **1** using a pipette and shake the mixture. Read and record the transfer time (in minutes and seconds, to the nearest second) when half of the reaction mixture has emptied from the pipette.
7. Transfer all the contents of boiling tube **1** into a second conical flask. Rinse this boiling tube with deionised water and add the washings to the conical flask.
8. **Immediately** titrate the I_2 in the second conical flask with **FA 6** until the solution turns pale yellow. Then add about 1 cm^3 of starch indicator. The solution will turn blue-black. Continue to titrate until the blue-black colour **just** disappears at the end-point. Record the final burette reading and the volume of **FA 6** added.
9. Empty the contents of this conical flask into the waste bottle. Wash this conical flask thoroughly with water.

10. Repeat Step **6** to **9** four more times at about 8 minutes, 12 minutes, 16 minutes and 20 minutes, transferring the aliquots into the boiling tubes labelled **2** to **5**.

(b) **Results**

Transfer time	Final burette reading /cm^3	Initial burette reading /cm^3	Volume of FA 6 added /cm^3	t /min
4 min 12 s	16.80	0.00	16.80	4.2
8 min 30 s	15.10	0.20	14.90	8.5
12 min 5 s	28.70	15.10	13.60	12.1
16 min 0 s	40.70	28.70	12.00	16.0
20 min 18 s	40.15	29.90	10.25	20.3

[5]

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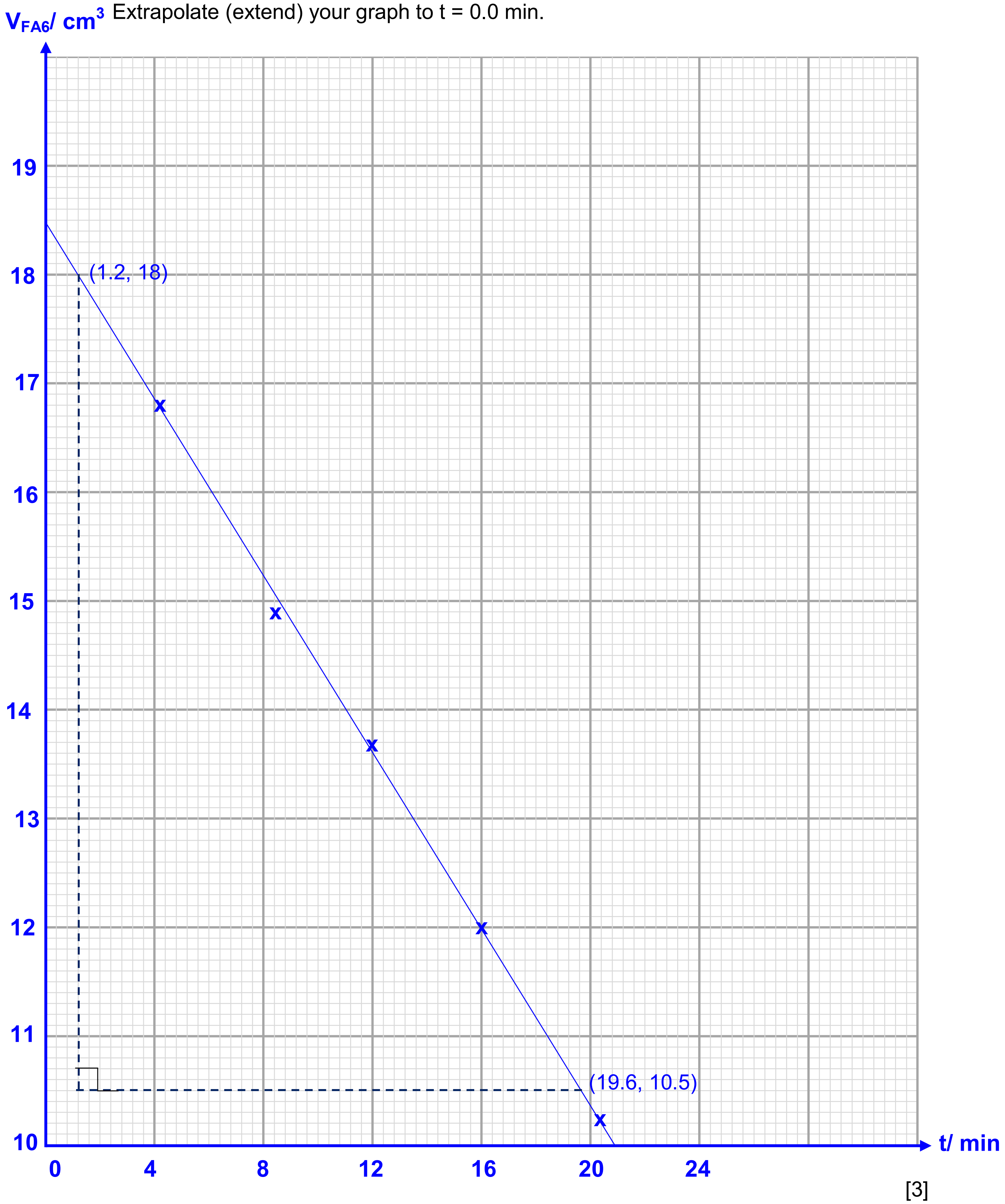
(c)(i) On the grid below, plot a graph of the volume of sodium thiosulfate, **FA 6**, on the y-axis, against time, t, on the x-axis.

Draw the most appropriate best-fit graph taking into account all of your plotted points. Extrapolate (extend) your graph to t = 0.0 min.



[3]

[Turn over

(ii) Deduce the order of reaction with respect to the $[I_2]$ in reaction 1. Explain your answer.

Order……**Order of reaction is zero with respect to I_2**……

explanation……**Graph is a straight line with constant gradient / rate of reaction is independent of $[I_2]$ / $[I_2]$ decreases linearly with time / $[I_2]$ decreases at a constant rate.**……

…… [1]

(iii) Given that the reaction is first order with respect to both CH_3COCH_3 and H^+ ions and using your answer from **(c)(ii)**, write the rate equation for **reaction 1**.

Rate = $k\,[CH_3COCH_3][H^+]$

…… [1]

(iv) The reaction between propanone and bromine proceeds by a similar mechanism as the reaction between propanone and iodine. How would the rate of reaction between propanone and bromine be compared with that of propanone and iodine? Explain your answer.

The rate would be the same as I_2 (and hence Br_2) does not appear in the rate equation.

…… [1]

(d) The **initial rate** of change of concentration of iodine, $[I_2]$, can be determined from the gradient of the graph at t = 0.0 min.

(i) Calculate the gradient of the graph at t = 0.0 min, showing clearly how you did this.

$$\text{Gradient} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{10.5 - 18}{19.6 - 1.2}$$

$$= -0.408 \text{ cm}^3 \text{ min}^{-1} \text{ (3 s.f)}$$

Gradient = ………………………….. cm^3 min^{-1}

[1]

(ii) Use your answer in **(d)(i)** to determine the rate of change of the amount of $S_2O_3^{2-}$ ions required in mol min^{-1}.

rate of change = (d)(i) × $\frac{0.010}{1000}$

= − 0.408 × $\frac{0.010}{1000}$

= − 4.08 × 10^{-6} mol min^{-1} (3 s.f.)

Rate of change of the amount of $S_2O_3^{2-}$ required = **−4.08 × 10^{-6}** mol min^{-1} [1]

(iii) Use your answer in **d(ii)** to determine change in amount of I_2 at t = 0.0 min in mol min^{-1}.

$I_2 \equiv 2S_2O_3^{2-}$

rate change of I_2 = $\frac{1}{2}$ × (d)(ii)

= $\frac{1}{2}$ × - 4.08 × 10^{-6}

= − 2.04 × 10^{-6} mol min^{-1} (3 s.f.)

Change in amount of I_2 = **−2.04 × 10^{-6}** mol min^{-1} [1]

(iv) Using your answer in **d(iii)** and considering that a 10.0 cm^3 aliquot (portion) of the reaction mixture was used, determine the rate of change of I_2 at t = 0.0 min in mol dm^{-3} min^{-1}

rate of change of [I_2] = (d)(iii) ÷ $\frac{10}{1000}$ ***divide by volume of 10cm^3 aliquot***

= −2.04 x 10^{-6} ÷ $\frac{10}{1000}$

≈ − 2.04 x 10^{-4} mol dm^{-3} min^{-1} (3 s.f)

Rate of change of [I_2] at t = 0.0 min = **− 2.04 x 10^{-4}** mol dm^{-3} min^{-1} [1]

(e) In Step 6, the aliquot is transferred into a boiling tube containing $NaHCO_3$. Explain why $NaHCO_3$ must be added and what effect failing to do it would have on the titre values.

$NaHCO_3$ must be added to react with the acid catalyst to quench the reaction / acts as a quenching agent. If aliquots are not quenched, the titre values would be lower than expected.

[Turn over

…………………………………………………………………………………………………… [1]

[Total: 16]

4 Planning: Measuring the activation energy for a reaction

Vitamin C, also known as ascorbic acid ($C_6H_8O_6$) is essential to health. However, in the presence of heat, ascorbic acid is easily oxidised to L-dehydroascorbic acid ($C_6H_6O_6$) as shown in **reaction 1** below and it no longer serves its purpose.

reaction 1 $C_6H_8O_6 \rightarrow C_6H_6O_6 + 2H^+ + 2e^-$

The rate equation for this reaction is

$$\text{rate} = k\,[C_6H_8O_6]$$

where *k* is the rate constant.

The rate of oxidation of ascorbic acid can be determined by monitoring the amount of ascorbic acid left in the sample at various timed intervals. The oxidation of ascorbic acid becomes significant only at temperatures **40 °C and above**.

Portions of the reaction mixture are

- removed at regular timed intervals,
- quenched by putting it in an ice bath,
- titrated against a standard solution of aqueous iodine, I_2 as shown in **reaction 2**

reaction 2 $C_6H_8O_6 + I_2 \rightarrow C_6H_6O_6 + 2H^+ + 2I^-$

The activation energy, E_a, of the reaction can be determined from the equation.

$$\ln k = -\frac{E_a}{R}\left(\frac{1}{T}\right) + c$$

R is the molar gas constant.
T is the reaction temperature in kelvin.
k is the rate constant at a chosen temperature.
c is a constant for the reaction.

k can be determined from the initial rate of each experiment at different temperatures, *T*.

$$k = \frac{\text{Initial rate}}{[C_6H_8O_6] \text{ at t=0 min}}$$

(a) Plan an investigation, to determine the effect of temperature, *T*, on the rate of oxidation of ascorbic acid using the titrimetric method to monitor the rate as described above. The activation energy, E_a, is to be graphically determined.

Plotting ln *k* against $\frac{1}{T}$ gives a straight line of best fit. The gradient of this line is $-\frac{E_a}{R}$, where *R* is the molar gas constant.

You may assume that you are provided with

- five Vitamin C tablets (each tablet contains 500mg of ascorbic acid),
- deionised water,
- aqueous iodine, I_2,
- starch solution,
- ice,
- hot water bath,
- the equipment normally found in a school or college laboratory.

In order to prepare the reaction mixture at temperature, *T* °C, a total of 75 cm^3 of deionised water is added to a conical flask. Ensure that the temperature of the deionised water has reached *T* °C before adding one Vitamin C tablet into the conical flask. The conical flask is then swirled to ensure the entire tablet has been dissolved.

Your plan should include brief details of:

- the reactants and conditions that you would use,
- the apparatus you would use,
- the procedure you would follow and the measurements you would make at different timings and temperatures,
- how you would determine the initial rate for each experiment by plotting an appropriate graph.

In your plan, it is **not** required to refer to concentrations or to perform calculations.

1. Fill a 50.00 cm^3 burette with aqueous I_2.
2. Using a second burette, add **75.0 cm^3** of water into a 250 cm^3 conical flask labelled 'Reaction Mixture'.
3. Place the conical flask of water in a **water bath** of 40 °C. Measure the temperature of the conical flask of deionised water using a **thermometer**. Once the flask of deionised water has reached 40 °C, add 1 Vitamin C tablet into the flask and swirl the flask to ensure all Vitamin C has been dissolved. **Start the stopwatch once the tablet has been dissolved**.
4. After 5 minutes, use a **pipette** to transfer **10.0 cm^3** aliquot of reaction mixture into a 250 cm^3 **conical flask**.
5. Place the conical flask in the **ice bath** to quench the reaction.
6. Add 5 drops of **starch indicator** into the conical flask and immediately titrate the reaction mixture with I_2. The end point is reached when the solution in the conical flask turns from **colourless to first permanent blue-black** colour is seen.
7. Repeat steps 4 to 6 and until a total of **four aliquots** have been titrated and their results recorded. Step 4 is repeated at appropriate timings of **10, 15, 20, and 25 minutes**.
8. Repeat the experiment (from steps 2 to 7) for another **four** reaction mixtures at **different temperatures** (50 °C, 60 °C, 70 °C, 80 °C) by placing the 'Reaction Mixture' prepared in step 2 in a water bath at the chosen temperature.
9. To determine the **initial rate** for each experiment

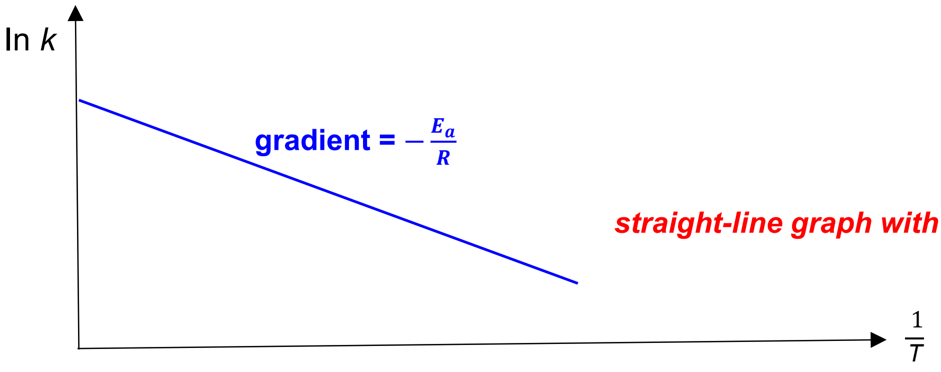
For each experiment, plot a graph of the **volume of I_2** added on the y-axis **against time** on the x-axis.

[8]

The initial rate is given by the **gradient of the tangent to the curve at t = 0 min**.

[Turn over

(b) Sketch the graph you would expect to obtain from **4(a)** on the axes below.



straight-line graph with negative gradient

[1]

(c) Describe how you would use your graph to determine the value of E_a.

From the gradient of the graph, $\frac{-E_a}{R}$ can be determined.

gradient of straight-line graph = $\left(\frac{y_1 - y_2}{x_1 - x_2}\right) = -\frac{E_a}{R}$

$\therefore E_a = -\left(\frac{y_1 - y_2}{x_1 - x_2}\right) \times 8.31$ J mol^{-1}

[1]

[Total: 10]

Qualitative Analysis Notes

[*ppt. = precipitate*]

(a) Reactions of aqueous cations

cation	***reaction with***	
	NaOH(aq)	NH_3(aq)
aluminium, Al^{3+}(aq)	white ppt. soluble in excess	white ppt. insoluble in excess
ammonium, NH_4^+(aq)	ammonia produced on heating	–
barium, Ba^{2+}(aq)	no ppt. (if reagents are pure)	no ppt.
calcium, Ca^{2+}(aq)	white ppt. with high [Ca^{2+}(aq)]	no ppt.
chromium(III), Cr^{3+}(aq)	grey-green ppt. soluble in excess giving dark green solution	grey-green ppt. insoluble in excess
copper(II), Cu^{2+}(aq)	pale blue ppt. insoluble in excess	blue ppt. soluble in excess giving dark blue solution
iron(II), Fe^{2+}(aq)	green ppt. turning brown on contact with air insoluble in excess	green ppt. turning brown on contact with air insoluble in excess
iron(III), Fe^{3+}(aq)	red-brown ppt. insoluble in excess	red-brown ppt. insoluble in excess
magnesium, Mg^{2+}(aq)	white ppt. insoluble in excess	white ppt. insoluble in excess
manganese(II), Mn^{2+}(aq)	off-white ppt. rapidly turning brown on contact with air insoluble in excess	off-white ppt. rapidly turning brown on contact with air insoluble in excess
zinc, Zn^{2+}(aq)	white ppt. soluble in excess	white ppt. soluble in excess

[Turn over

(b) Reactions of aqueous anions

ion	*reaction*
carbonate, CO_3^{2-}	CO_2 liberated by dilute acids
chloride, Cl^-(aq)	gives white ppt. with Ag^+(aq) (soluble in NH_3(aq))
bromide, Br^-(aq)	gives pale cream ppt. with Ag^+(aq) (partially soluble in NH_3(aq))
iodide, I^-(aq)	gives yellow ppt. with Ag^+(aq) (insoluble in NH_3(aq))
nitrate, NO_3^-(aq)	NH_3 liberated on heating with OH^-(aq) and Al foil
nitrite, NO_2^-(aq)	NH_3 liberated on heating with OH^-(aq) and Al foil; NO liberated by dilute acids (colourless NO → (pale) brown NO_2 in air)
sulfate, SO_4^{2-}(aq)	gives white ppt. with Ba^{2+}(aq) (insoluble in excess dilute strong acids)
sulfite, SO_3^{2-}(aq)	SO_2 liberated with dilute acids; gives white ppt. with Ba^{2+}(aq) (soluble in dilute strong acids)

(c) Tests for gases

gas	*test and test result*
ammonia, NH_3	turns damp red litmus paper blue
carbon dioxide, CO_2	gives a white ppt. with limewater (ppt. dissolves with excess CO_2)
chlorine, Cl_2	bleaches damp litmus paper
hydrogen, H_2	"pops" with a lighted splint
oxygen, O_2	relights a glowing splint
sulfur dioxide, SO_2	turns aqueous acidified potassium manganate(VII) from purple to colourless

(d) Colour of halogens

halogen	*colour of element*	*colour in aqueous solution*	*colour in hexane*
chlorine, Cl_2	greenish yellow gas	pale yellow	pale yellow
bromine, Br_2	reddish brown gas / liquid	orange	orange-red
iodine, I_2	black solid / purple gas	brown	purple