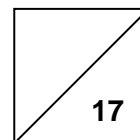


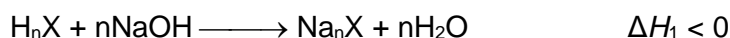
Name: (2T)

Date:

Catholic Junior College
H2 Chemistry (9729)**Experiment 22:** Determination of the basicity of H_nX and enthalpy change of reaction of H_nX with $NaOH$ **FA 1** is $0.50 \text{ mol dm}^{-3} H_nX$.**FA 2** is 1.10 mol dm^{-3} sodium hydroxide, $NaOH$.

You will perform a series of experiments where different volumes of **FA 1** and **FA 2** are mixed and the temperature change, ΔT , for each experiment is determined. You will then plot a graph of ΔT against the volume of **FA 1** used.

With the aid of this graph, you will determine the basicity of H_nX and the enthalpy change, ΔH_1 , of the exothermic reaction between H_nX and sodium hydroxide.

**(a) Procedure**

1. Use a 50 cm^3 measuring cylinder to transfer 10.0 cm^3 of **FA 1** into a polystyrene cup. Place this polystyrene cup in a 250 cm^3 beaker to prevent the cup from tipping over.
2. Use a thermometer to stir and measure the temperature of this solution. Tilt the cup if necessary to ensure that the bulb of the thermometer is fully immersed. The temperature, T_1 , is the initial temperature of the solution.
3. Use another 50 cm^3 measuring cylinder to transfer 40.0 cm^3 of **FA 2** into the polystyrene cup containing **FA 1**.
4. Using the thermometer, stir the mixture continuously until it reaches its **highest** temperature. Record this temperature, T_2 .
5. Discard the contents of the polystyrene cup. Wash and dry it carefully.
6. Repeat steps 1 to 5 using 15.0 cm^3 , 20.0 cm^3 , 25.0 cm^3 , 30.0 cm^3 and 35.0 cm^3 of **FA 1**, each time using an appropriate volume of **FA 2** so that the total volume of each mixture is 50.0 cm^3 .
7. Record all measurements of volume, temperature and temperature change, ΔT , in an appropriate format in the space provided below.

Results

Volume of FA 1 / cm^3	10.0	15.0	20.0	25.0	30.0	35.0
Volume of FA 2 / cm^3	40.0	35.0	30.0	25.0	20.0	15.0
$T_2 / ^\circ\text{C}$	32.7	34.0	35.2	35.0	33.8	32.6
$T_1 / ^\circ\text{C}$	28.6	28.6	28.6	28.6	28.6	28.6
$\Delta T / ^\circ\text{C}$	4.1	5.4	6.6	6.4	5.2	4.0

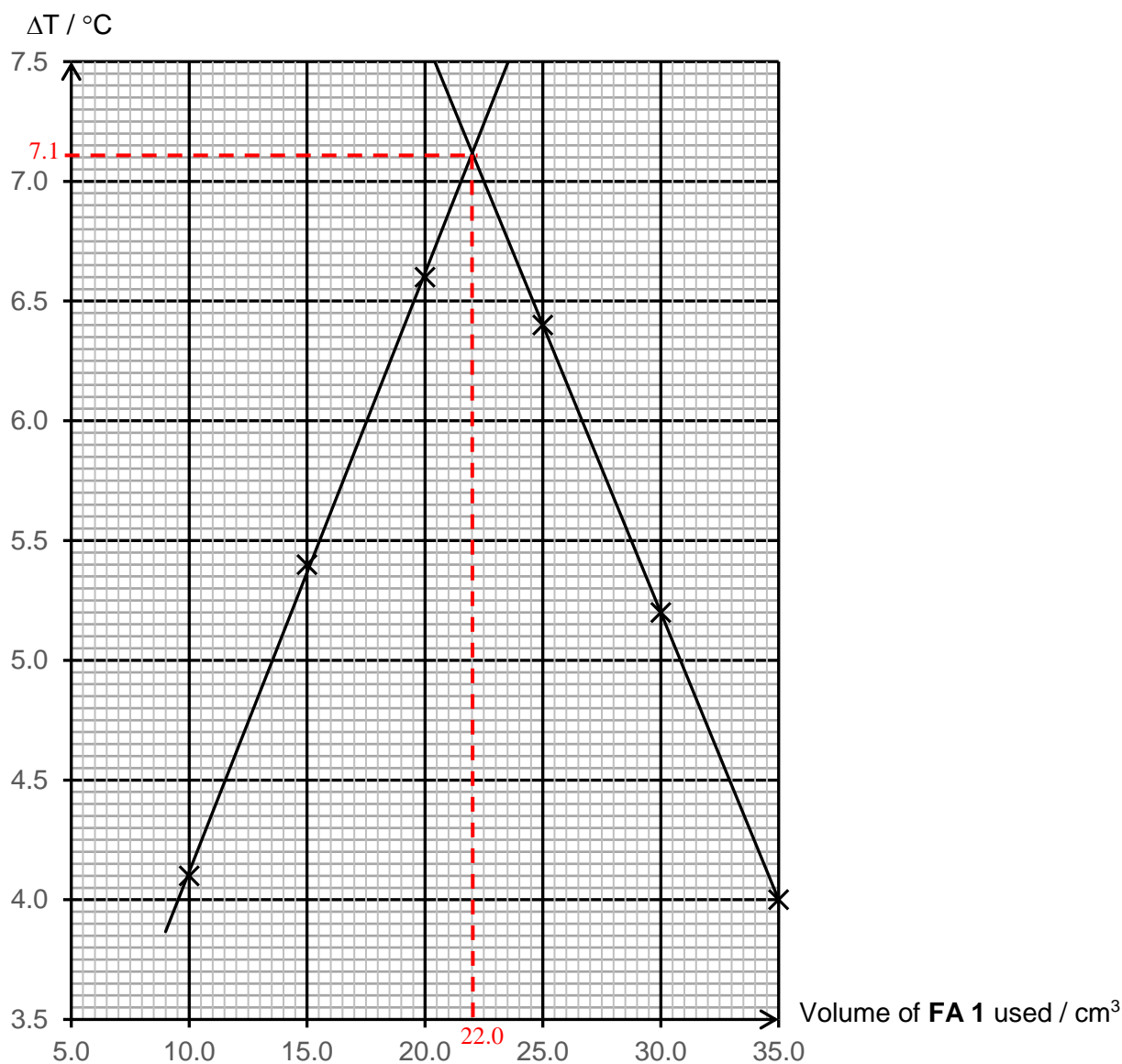
Appropriate headers and units [1]
 Appropriate volumes of **FA 2** to give overall volumes of 50 cm^3 [1]
 Recording of volumes to 1 d.p. and temperature to $0.2 ^\circ\text{C}$ [1]

[3]

- (b) Plot a graph of temperature change, ΔT , on the y-axis, against volume of **FA 1** added, on the x-axis on the grid provided below.

Draw two lines of best fit for the points plotted, and extrapolate both lines to find

- the maximum temperature change, ΔT_{\max} ,
- the volume of **FA 1** required to completely react with **FA 2**.



maximum temperature change of reaction mixture, $\Delta T_{\max} = \underline{+7.1\text{ }^{\circ}\text{C}} \text{ (1 d.p.)}$

volume of **FA 1** required for complete reaction, $V_{\text{rxn}} = \underline{22.0\text{ cm}^3} \text{ (1 d.p.)}$

[4]

Axis labels, appropriate scale chosen, graph covers $> \frac{1}{2}$ of graph paper in both directions [1]

All points plotted correctly [1]

Best fit lines through points in ^ shape [1]

ΔT_{\max} and V_{rxn} correctly obtained [1]

- (c) (i) Calculate the corresponding volume of **FA 2** required for complete reaction.

$$\text{Vol. of FA 2} = 50 - 22.0 = \underline{28.0 \text{ cm}^3} \text{ [1]}$$

$$\text{volume of FA 2} = \dots\dots\dots \underline{28.0 \text{ cm}^3} \dots\dots\dots [1]$$

- (ii) Calculate the number of moles of H_nX in **FA 1** and of NaOH in **FA 2** required for complete reaction.

$$\text{Amt of H}_n\text{X in FA 1} = 0.5 (22.0/1000) = \underline{0.0110 \text{ mol}} \text{ (3 s.f.)}$$

$$\text{Amt of NaOH in FA 2} = 1.10 (28.0/1000) = \underline{0.0308 \text{ mol}} \text{ (3 s.f.) [both, 1]}$$

$$\text{number of moles of H}_n\text{X in FA 1} = \dots\dots\dots \underline{0.0110 \text{ mol}} \dots\dots\dots$$

$$\text{number of moles of NaOH in FA 2} = \dots\dots\dots \underline{0.0308 \text{ mol}} \dots\dots\dots [1]$$

- (iii) (1) Use your answer to (c)(ii) to calculate the mole ratio of NaOH to H_nX required for complete reaction, to 3 significant figures.

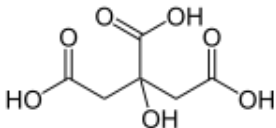
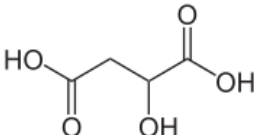
$$\begin{aligned} \text{Ratio of NaOH to H}_n\text{X} \\ &= 0.0308 : 0.0110 \\ &= \underline{2.80 : 1} \end{aligned}$$

$$\text{mole ratio} = \dots\dots\dots \underline{2.80 : 1} \dots\dots\dots$$

- (2) Hence, state the value of n (to the nearest whole number) in H_nX .

$$\text{value of } n = \dots\dots\dots \underline{n = 3} \text{ [Both, 1]} \dots\dots\dots [1]$$

(iv) Information about some acids is provided below.

acid	structural formula	basicity	pK_{a1}	pK_{a2}	pK_{a3}
citric acid		tribasic	3.13	4.76	6.40
phosphoric acid	H_3PO_4	tribasic	2.15	7.20	12.35
ethanedioic acid	$HOOC-COOH$	dibasic	1.23	4.19	--
malic acid		dibasic	1.90	6.07	--

H_nX is one of the above acids.

Using information from the above parts and that the pK_a value is estimated to be about 2-3, deduce the identity of H_nX , explaining how you arrived at your conclusion.

identity of H_nX phosphoric acid [1]

explanation From (c)(iii), H_nX is a tribasic acid i.e. H_nX is either citric acid
 or phosphoric acid.

..... Since the pK_a of H_nX is estimated to be 2-3. [1]

..... Hence, H_nX is phosphoric acid. [2]

(v) Calculate the enthalpy change of reaction per mole of H_nX , ΔH_1 , in kJ mol^{-1} , for the reaction between H_nX and NaOH .

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

$$q = mc\Delta T = 50(4.18)(7.1) = 1483.9 \text{ J [1]}$$

$$\Delta H = -1483.9 / 0.0110 = -134.9 \text{ kJ mol}^{-1} = \underline{-135 \text{ kJ mol}^{-1} \text{ (3sf) [1]}}$$

$$\Delta H_1 = \underline{-135 \text{ kJ mol}^{-1}} \dots\dots\dots [2]$$

- (d) Suggest how the value of ΔH_1 would be different if H_nX were a strong acid. Explain your answer.

The value of ΔH_1 would be more negative (or more exothermic). If H_nX were a strong acid, it would have spontaneously dissociated completely in aqueous solution. During the neutralisation reaction with NaOH, energy is not needed to break any O-H bond in H_nX and hence more heat energy is released. [1]

[1]

- (e) By considering the pK_a 's of the acid you have identified in (c)(iv), account for the mole ratio of NaOH to H_nX calculated to 3 significant figures in (c)(iii)(1).

From (c)(iii), the mole ratio of H_nX to NaOH is 2.80 : 1, implying that the third proton of only some of the acid in solution is released in the reaction i.e. it is difficult to release the third proton from H_3PO_4 . [1, owtte]

[1]

- (f) Predict and explain the effect on ΔT_{max} if solutions of lower concentrations were used in the procedure in (a) instead of FA 1 and FA 2.

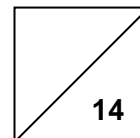
For the same volumes of solutions, because they are more dilute, there are fewer number of moles of reactants/less moles of H_2O formed. Hence less energy will be released and ΔT_{max} will be smaller. [1]

[1]

[Total: 17]

Name: (2T)

Date:

Catholic Junior College
H2 Chemistry (9729)**Experiment 23:** Investigation of some inorganic and organic reactions**FA 5** is a solid sample which contains a divalent cation, M^{2+} , and CH_3COO^- .

Carry out the following tests. Carefully record your observations in the tables provided.

Test and identify any gases evolved. If there is no observable change, write **no observable change**.**Table 3.1**

		test	observations
(a)	(i)	<p>Add one full spatula of FA 5 into a test-tube.</p> <p>Add 5 cm depth of deionised water to dissolve the solid.</p> <p>Leave the test-tube to stand to allow any undissolved solid to settle to the bottom of the test-tube.</p> <p>Decant the solution into a clean test-tube. This solution is FA 6.</p>	<p>A blue solution was obtained. [1]</p> <p>[1]</p>
	(ii)	<p>Add 1 cm depth of FA 6 to a test-tube.</p> <p>Add aqueous ammonia slowly, with shaking, until no further change is seen.</p> <p>Add dilute nitric acid dropwise until in excess.</p>	<p>Blue ppt formed.</p> <p>Ppt soluble in excess NH_3, forming a deep blue solution. [1]</p> <p>Blue ppt reformed, soluble in excess HNO_3 to form a pale blue solution. [1]</p> <p>[2]</p>
	(iii)	<p>Add 0.5 cm depth of FA 6 to a boiling tube.</p> <p>Add 1 cm depth of dilute nitric acid slowly.</p> <p>Add 4 full spatulas of solid sodium chloride and shake the mixture thoroughly until no further change is seen.</p> <p>Heat the mixture.</p>	<p>The blue solution turned <u>pale blue</u> [1] (and a sour smell was observed – this is ethanoic acid!)</p> <p>On adding $NaCl$, the pale blue solution turned <u>green</u>. [1]</p> <p>On heating, the green solution turned <u>yellow</u>. [1]</p> <p>[3]</p>

	<p>(iv) To a quarter of a spatula of FA 5 in a test-tube, add 10 drops of ethanol and 3 drops of concentrated sulfuric acid.</p> <p>Heat the mixture.</p>	<p>On heating, a sweet smell was observed. [1, accept other reasonable descriptions of the smell]</p> <p>[1]</p>
--	--	--

- (v) Identify the cation, M^{2+} , present in **FA 6**. Use evidence from your observations in (a)(ii) to support your deduction.

cation Cu^{2+}

evidence In test (a)(i), **FA 5** reacted with $NH_3(aq)$ to form a blue ppt. of $Cu(OH)_2$

which was soluble in excess $NH_3(aq)$ to give a dark blue solution

containing $[Cu(NH_3)_4]^{2+}$. [1]

- (vi) Write a balanced equation for the reaction that occurred in (a)(iv).

equation $CH_3CO_2H + CH_3CH_2OH \rightarrow CH_3CO_2CH_2CH_3 + H_2O$ [1]

Do note that formation of ester is not used as a distinguishing test because smell is subjective. The experiment in (a)(iv) is conducted to enable you to experience the esterification reaction.

Read all the instructions for test (b) carefully before you start this experiment. You are to perform test (b) only once. Record your observations in Table 3.2.

You are reminded to wear safety goggles as a safety precaution.

You will be making use of the experimental set-up shown in Figure 1 below.

NOTE:

To prevent suck back during heating, remove the test-tube containing deionised water from the delivery tube first. Then, remove the test-tube containing **FA 5** from the Bunsen flame.

After heating, immediately return the test-tube labelled **FA 5**, with the delivery tube attached, to the test-tube rack to cool down. DO NOT remove the rubber bung with delivery tube from the hot test-tube.

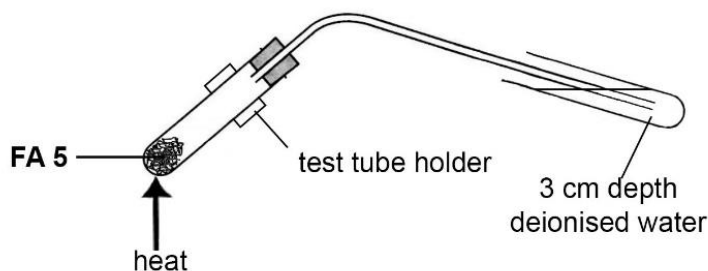


Figure 1

Table 3.2

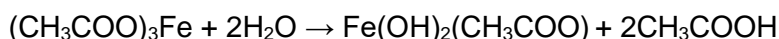
	test	observations
(b)	<p>Prepare the experimental set-up as shown in Figure 1 above, by following the steps below.</p> <ul style="list-style-type: none"> • Add 1 full spatula of FA 5 into a clean and dry test-tube. • Securely attach a rubber bung with a delivery tube to the test-tube containing FA 5. • Place a separate test-tube with 3 cm depth of deionised water, as shown in Figure 1. <p>Heat FA 5 for approximately 1 minute to collect any vapour evolved into the deionised water.</p> <p><u>Do not heat FA 5 beyond 1 minute.</u></p> <p>To prevent suck back, see the note above.</p> <p>Stir this solution in the other test-tube thoroughly. This solution is FA 7. Keep FA 7 for test (c).</p>	<p>Water vapour condensed on the cooler parts of the test tube as water droplets.</p> <p>Brown solid dust was dispersed in the test tube containing FA 5.</p> <p>Colour of solid turned from <u>blue to pink</u>. [1]</p> <p>[1]</p>

Carry out the following test on **FA 7** and record your observations in Table 3.3.

Table 3.3

		test	observations
(c)	(i)	Add 1 cm depth of FA 7 to a boiling tube. Add an equal volume of neutral iron(III) chloride. Carefully boil the mixture.	The solution turned from yellow to orange-brown. A red-brown ppt was formed. [1]

(ii) Upon boiling the mixture, the following reaction occurred.



State the type of reaction that occurred in this reaction.

...Acid-Base reaction [1].....[1]

When **FA 5** undergoes further thermal decomposition, it produces a volatile compound, **FA 8**, which can be distilled out for analysis.

You are given the results of some tests that were carried out on **FA 8**, in Table 3.4.

Table 3.4

		test	observations
(d)	(i)	Add 1 cm depth of aqueous sodium hydroxide and 1 cm depth of aqueous iodine to a test-tube. Add 1 cm depth of FA 8 .	Yellow ppt was formed.
	(ii)	To 1 cm depth of FA 8 in a test-tube, add 5 drops of 2,4-dinitrophenylhydrazine.	Orange ppt was formed.

(iii) **FA 8**, the smallest member in its homologous series, is not oxidised by KMnO_4 . Use evidence from the observations in (d)(i) and (ii) to identify **FA 8**.

FA 8 ... CH_3COCH_3 [1]..... [1]

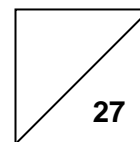
(iv) Upon strong heating, 1 mole of **FA 5** decomposes to give equimolar amounts of its solid oxide, an acidic gas and **FA 8**. Write an equation for the thermal decomposition of 1 mole of **FA 5**.

$\text{Cu}(\text{CH}_3\text{COO})_2(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{CO}_2(\text{g}) + \text{CH}_3\text{COCH}_3(\text{g})$ [1].....[1]

[Total: 14]

Name: (2T)

Date:

Catholic Junior College
H2 Chemistry (9729)

Experiment 24: Determination of the dependence of the rate of reaction between glucose and acidified potassium manganate(VII) on temperature

Read through the whole method before starting any practical work. Where appropriate, prepare a table for your results in the space provided.

Show your working and appropriate significant figures in the final answer to **each** step of your calculations.

Glucose, $C_6H_{12}O_6$, is a sugar that can act as a reducing agent. You will investigate how an increase in temperature affects the rate of the redox reaction between glucose and acidified potassium manganate(VII).

FA 1 is $0.020 \text{ mol dm}^{-3}$ acidified potassium manganate(VII), $KMnO_4$.

FA 2 is 1.0 mol dm^{-3} sulfuric acid, H_2SO_4 .

FA 3 is an aqueous solution containing 32.8 g dm^{-3} glucose, $C_6H_{12}O_6$.

FA 4 is a solid mixture containing an unknown percentage of glucose.

You will measure the time it takes for the purple colour to disappear. Your table of results on the following page should include the rate of reaction for each experiment.

(a) Method**Experiment 1**

1. Fill the burette with **FA 1**.
2. Add 5.00 cm^3 of **FA 1** into the 250 cm^3 beaker.
3. Use the 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of **FA 2** into the beaker containing **FA 1**.
4. Use the same measuring cylinder to transfer 50.0 cm^3 of distilled water into the same beaker.
5. Place the beaker on the tripod and heat its contents to between 75°C and 80°C .
6. While the solution in the beaker is being heated, pour 25.0 cm^3 of **FA 3** into the 25 cm^3 measuring cylinder.
7. When the temperature of the contents of the beaker has reached between 75°C and 80°C , turn off the Bunsen burner and **carefully** hold the top of the hot beaker with either a cloth or paper towel and place it onto the white tile.
8. Record the temperature of the solution in the beaker.
9. Add the 25.0 cm^3 of **FA 3** and **immediately** start timing.
10. Stir the contents of the beaker once and stop timing as soon as the solution turns colourless. Record the time to the nearest second.
11. Record the temperature of the solution as soon as it is colourless.
12. Calculate and record the average temperature of the reaction mixture to one decimal place.
13. Empty, rinse and dry the beaker so it is ready for use in **Experiment 2**.

Experiment 2

1. Add 5.00 cm³ of **FA 1** into the 250 cm³ beaker.
2. Use the 50 cm³ measuring cylinder to transfer 50.0 cm³ of **FA 2** into the beaker containing **FA 1**.
3. Use the same measuring cylinder to transfer 50.0 cm³ of distilled water into the same beaker.
4. Place the beaker on the tripod and heat its contents to between 35 °C and 40 °C.
5. While the solution in the beaker is being heated, pour 25.0 cm³ of **FA 3** into the 25 cm³ measuring cylinder.
6. When the temperature of the contents of the beaker has reached between 35 °C and 40 °C, turn off the Bunsen burner and **carefully** hold the top of the hot beaker with either a cloth or paper towel and place it onto the white tile.
7. Record the temperature of the solution in the beaker.
8. Add the 25.0 cm³ of **FA 3** and **immediately** start timing.
9. Stir the contents of the beaker once and stop timing as soon as the solution turns colourless. Record the time to the nearest second.
10. Record the temperature of the solution as soon as it is colourless.
11. Calculate and record the average temperature of the reaction mixture to one decimal place.
12. Empty, rinse and dry the beaker so it is ready for use in **Experiment 3**.

Experiments 3, 4 and 5

1. Repeat the method for **Experiment 2** at three different temperatures.
2. Keep the temperature of the contents of the beaker between 35 °C and 40 °C.
3. Record all your results in your table.

Results:

The rate of reaction can be calculated as shown.

$$\text{rate} = \frac{1000}{\text{reaction time}}$$

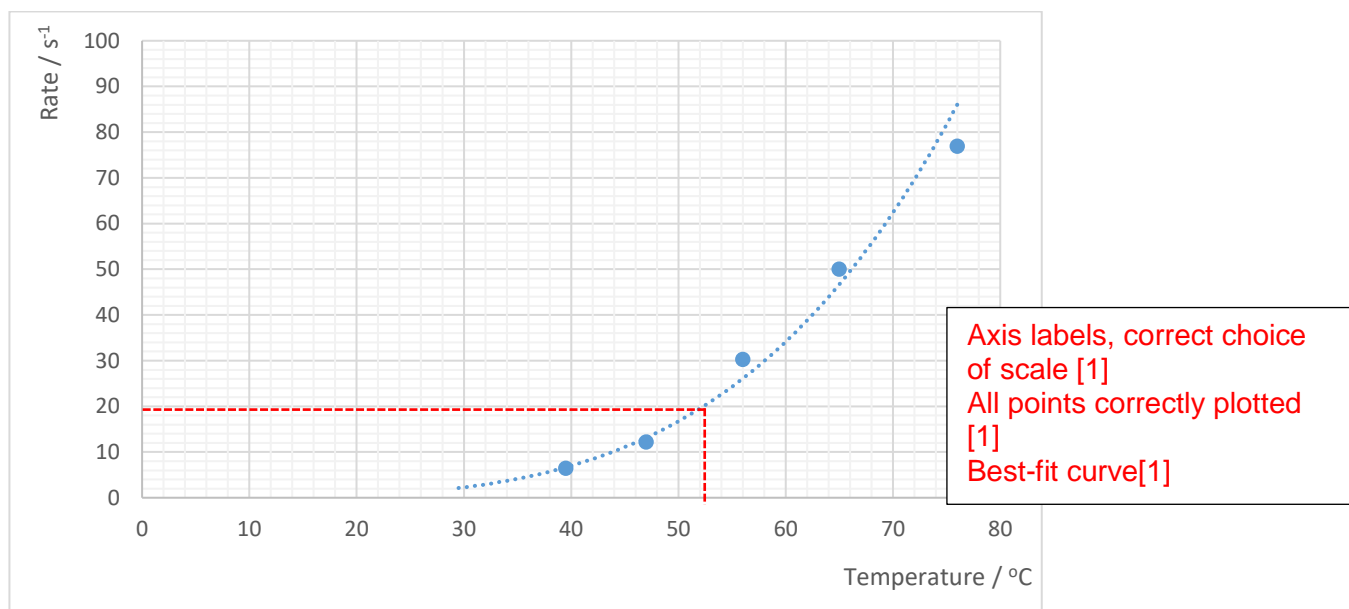
Calculate the rate of reaction, to 1 decimal place, for each experiment and include this in your table.

Experiment	$T_{\text{initial}} / ^\circ\text{C}$	$T_{\text{final}} / ^\circ\text{C}$	$T_{\text{average}} / ^\circ\text{C}$	time / s	rate / s ⁻¹
1	80.0	72.0	76.0	12	83.3
2	40.0	39.0	39.5	155	6.45
3	49.0	45.0	47.0	82	12.2
4	59.0	53.0	56.0	33	30.3
5	69.0	61.0	65.0	20	50.0

Appropriate headers and units [1]
 Initial temperatures chosen for expts 3-5 are at least 10 °C apart from each other [1]
 Average temperature calculated correctly [1]
 Rate calculated correctly [1]
 Average temperature & rate given to 1 d.p., time recorded to nearest second [1]

- (b) Plot a graph of rate (y-axis) against average temperature (x-axis) on the grid below. Select a scale on the x-axis to include an average temperature of 30.0 °C. Label any points you consider anomalous.

Draw the most appropriate best-fit curve taking into account all of your plotted points and extrapolate it to 30.0 °C.



[3]

- (c) Use your graph to calculate the time to the nearest second that the reaction would have taken if the average temperature had been 52.5 °C.

Show on the grid how you obtained your answer.

From the graph, at 52.5 °C,

$$\text{rate} = \frac{1000}{\text{reaction time}} = 20 \text{ s}^{-1} \quad [1, \text{correct rate obtained, showing working on the grid}]$$

$$\text{Hence, reaction time} = \frac{1000}{20} = 50 \text{ s} \quad [1]$$

$$\text{time} = 50 \text{ s}$$

[2]

- (d) Explain, by referring to your graph or your table of results, how the rate of reaction is affected by an increase in temperature.

The rate of reaction increases with increase in temperature. The gradient increases with temperature / graph is exponential / acceleration of rate with temperature increase. [1]

[1]

- (e) (i) Calculate the maximum percentage error in the reaction time recorded for **Experiment 1**. Assume the error of the timer is ± 1 s.

$$\begin{aligned}\text{Maximum percentage error} &= 1 / 12 \times 100 \\ &= 8.33\% \text{ [1]}\end{aligned}$$

maximum percentage error in **Experiment 1** = 8.33 % [1]

- (ii) You have carried out experiments at five different temperatures.

Identify an experiment, if any, you should have repeated. Give a reason for your answer.

Identifies no anomaly as all points are near line of best fit / identifies an experiment to repeat, e.g., temperature drop was too large, so average temperature obtained may not be accurate [1, accept reasonable answers]

[1]

- (iii) Suggest one way to improve the accuracy of the results for this investigation.

Use thermostatically controlled water bath to heat both reagents and keep them at constant temperature. [1]

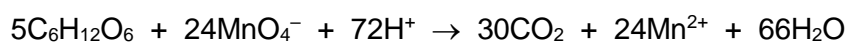
[1]

- (f) (i) Calculate the concentration of glucose in **FA 3** in mol dm^{-3} .

$$\text{Concentration of glucose} = \frac{32.8}{180.0} = 0.182 \text{ mol dm}^{-3} \quad [1]$$

$$\text{concentration of glucose in FA 3} = 0.182 \text{ mol dm}^{-3} \quad [1]$$

- (ii) Calculate the volume of $0.020 \text{ mol dm}^{-3}$ acidified KMnO_4 that would react with **all** the glucose present in 25.0 cm^3 of **FA 3**.



$$\begin{aligned} \text{Amount of glucose in } 25.0 \text{ cm}^3 &= 0.182 \times 25.0/1000 \\ &= 4.55 \times 10^{-3} \text{ mol} \quad [1] \end{aligned}$$

$$\begin{aligned} \text{Amount of KMnO}_4 \text{ reduced} &= 4.55 \times 10^{-3} \times \frac{24}{5} \\ &= 2.18 \times 10^{-2} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Volume of KMnO}_4 \text{ reduced} &= 2.18 \times 10^{-2}/0.020 \\ &= 1.09 \text{ dm}^3 \quad [1] \end{aligned}$$

$$\text{Volume of acidified KMnO}_4 \text{ required} = 1.09 \text{ dm}^3 \quad [2]$$

- (iii) Comment on the feasibility of conducting a redox titration between the given concentration of glucose and acidified KMnO_4 .

It is not feasible as the volume required for acidified KMnO_4 is too large. [1]

[1]

(g) Determination of the percentage of glucose in FA 4 by titration (DRY PRACTICE)

1. Weigh accurately 1.5 g of **FA 4** in a weighing bottle. Record your weighing appropriately in the space below. If you use **TARE** facility of the balance, please indicate clearly in your recording.
2. Dissolve the solid in a beaker and quantitatively transfer into a 250 cm³ volumetric flask. Make up to the mark with distilled water. Label this **FA 4 solution**.
3. Pipette 25.0 cm³ of **FA 4 solution** into a conical flask.
4. Using a measuring cylinder add 50.0 cm³ of **FA 2** into the conical flask.
5. Place the conical flask on the tripod and heat its contents to between 75°C and 80°C.
6. When the temperature of the contents of the conical flask has reached between 75 °C and 80°C, turn off the Bunsen burner and **carefully** hold the neck of conical flask with either a cloth or paper towel.
7. Titrate with **FA 1** until a permanent colour change is observed.
8. Record the burette readings in the suitable format. Repeat titration to achieve consistent results.

Results:*Weighing of FA 4*

Mass of weighing bottle and FA 4 / g	1.825
Mass of empty weighing bottle / g	0.325
Mass of FA 4 used / g	1.500

*Tabulation with ALL readings recorded, clear headers and units for mass table [1]
3 d.p. correct subtraction and mass of solid within stated range (1.450 – 1.550 g) [1]*

Titration

Experiment	1	2
Final burette reading / cm ³	22.40	42.40
Initial burette reading / cm ³	0.00	20.00
Volume of FA 1 used / cm ³	22.40	22.40

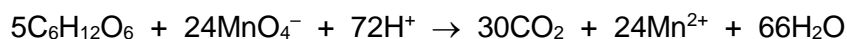
*Tabulation with ALL readings recorded, clear headers and units for titration table [1]
2 d.p. for titration reading [1]
At least two consistent readings ± 0.10 cm³ [1]*

- [5]
- (h) (i)** From your titrations, obtain a suitable volume of **FA 1** to be used in your calculations. Show clearly how you obtained this volume.

$$\begin{aligned}\text{Average volume of FA 1 used} &= (22.40 + 22.40) \div 2 \\ &= 22.40 \text{ cm}^3\end{aligned}$$

$$\text{volume of FA 1} = 22.40 \text{ cm}^3 \text{ [1] [1]}$$

- (ii) Calculate the amount of **FA 1** required for titration.



$$\begin{aligned}\text{Amount of FA 1 used} &= 22.40/1000 \times 0.020 \\ &= 4.48 \times 10^{-4} \text{ mol [1]}\end{aligned}$$

$$\text{amount of FA 1} = 4.48 \times 10^{-4} \text{ mol [1]}$$

- (iii) Hence, calculate the amount of glucose in **FA 4**.

$$\begin{aligned}\text{Amount of glucose in } 25.0 \text{ cm}^3 \text{ of FA 4 solution} &= 4.48 \times 10^{-4} \times \frac{5}{24} \\ &= 9.33 \times 10^{-5} \text{ mol} \\ \text{Amount of glucose in } 250 \text{ cm}^3 \text{ of FA 4 solution} &= 9.33 \times 10^{-5} \times \frac{250}{25.0} \\ &= 9.33 \times 10^{-4} \text{ mol [1]}\end{aligned}$$

$$\text{amount of glucose in FA 4} = 9.33 \times 10^{-4} \text{ mol [1]}$$

- (iv) Use your answer in (h)(iii) to determine the percentage of glucose in the mixture.
[Ar: C, 12.0; O, 16.0; H, 1.0]

$$\begin{aligned}\text{Mass of glucose in } 250 \text{ cm}^3 \text{ of FA 4 solution} &= 9.33 \times 10^{-4} \times 180.0 \\ &= 0.168 \text{ g}\end{aligned}$$

$$\begin{aligned}\text{Percentage of glucose in FA 4} &= 0.168/1.500 \times 100 \% \\ &= 11.2 \% [1]\end{aligned}$$

$$\text{percentage of glucose in the mixture} = 11.2 \% [1]$$

[Total: 27]