VICTORIA JUNIOR COLLEGE
JC2 PRELIMINARY EXAMINATION
Higher 2

CANDIDATE NAME .....

CT GROUP .....

CHEMISTRY 9729/04

Practical **30 August 2018**

2 hours 30 minutes

Additional Materials: As listed in the instructions below

READ THESE INSTRUCTIONS FIRST

Write your name and CT group on all the work you hand in.
Give details of the practical shift and laboratory where appropriate, in the boxes provided.
Write in dark blue or black pen.
You may use a 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 14 and 15.
Periodic Table is printed on page 16.

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	
2	
3	
4	
Total	

This document consists of **16** printed pages.

1 You are to determine the concentration, in g dm^{-3}, of sodium ethanedioate in a mixture of sodium ethanedioate and ethanedioic acid.

This experiment involves two titrations.
In titration one, you will carry out a titration to find the total amount of ethanedioate ion, $C_2O_4^{2-}$.
In titration two, you will use the information provided to find the amount of ethanedioic acid, $H_2C_2O_4$.
Finally, you will use the values found in the two titrations to calculate the concentration, in g dm^{-3}, of sodium ethanedioate in **FA 1**.

FA 1 is a mixture of aqueous sodium ethanedioate, $Na_2C_2O_4$, and ethanedioic acid, $H_2C_2O_4$.
FA 2 is approximately 2 mol dm^{-3} sulfuric acid, H_2SO_4.
FA 3 is 0.0200 mol dm^{-3} potassium manganate(VII), $KMnO_4$.

Titration One

(a) 1. By using a burette, measure between 42.50 cm^3 of **FA 1** into the 250 cm^3 graduated (volumetric) flask.
2. Record your burette readings and the volume of **FA 1** added to the flask in the space below.

Final burette reading / cm^3	42.50
Initial burette reading / cm^3	0.00
Volume of **FA 1** used / cm^3	42.50

[1]

3. Make up the contents of the flask to the 250 cm^3 mark with deionised water. Place the stopper in the flask and mix the contents thoroughly by slowly inverting the flask a number of times. Label this solution **FA 4**.
4. Fill a second burette with **FA 3**.
5. Pipette 25.0 cm^3 of **FA 4** from the graduated flask into a conical flask.
6. Use a measuring cylinder to add 25 cm^3 of **FA 2** to the conical flask.
7. Place the conical flask on a tripod and gauze and heat to about 65 °C.
8. If the neck of the flask is too hot to hold safely, use a folded paper towel to hold the flask.
9. Titrate the mixture in the conical flask with **FA 3** until a permanent pale pink colour is obtained. This is the end-point.
10. **If a brown colour appears during the titration**, reheat the flask to 65 °C. The brown colour should disappear and the titration can be completed as above. **If the brown colour does not disappear on reheating**, discard the solution and start the titration again.
11. Carry out as many titrations as you think necessary to obtain consistent results.
12. Record in an appropriate form all of your burette readings and the volume of **FA 3** added in each titration.

Final burette reading / cm^3	19.50	39.50
Initial burette reading / cm^3	0.00	20.00
Volume of **FA 3** used / cm^3	19.50	19.50

[5]

(b) From your titration results, obtain a suitable volume of **FA 3** to be used in your calculations. Show clearly how you obtained this volume.

Average volume of **FA 3** used
= (19.50 + 19.50) ÷ 2
= 19.50 cm^3

25.0 cm^3 of **FA 4** required ……19.50…… cm^3 of **FA 3**. [1]

Titration Two

(c) When 25.0 cm^3 of **FA 4** used in **(a)** is titrated with 0.100 mol dm^{-3} sodium hydroxide using phenolphthalein as the indicator, 15.50 cm^3 of sodium hydroxide is needed for complete reaction.

(i) Write an equation for the reaction between sodium hydroxide and ethanedioic acid.

$H_2C_2O_4 + 2NaOH \rightarrow Na_2C_2O_4 + 2H_2O$ [1]

(ii) Calculate the number of moles of sodium hydroxide required to react with 25.0 cm^3 of **FA 4**.

No. of moles of NaOH required
= $0.100 \times 15.50 \times 10^{-3}$
= 1.55×10^{-3} mol

moles of NaOH = 1.55×10^{-3} mol [1]

(iii) Hence, calculate the number of moles of ethanedioic acid in 25.0 cm^3 of **FA 4**.

No. of moles of $H_2C_2O_4$ present in 25.0 cm^3
= ½ × no. of moles of NaOH required
= ½ × 1.55×10^{-3}
= 7.75×10^{-4} mol

moles of $H_2C_2O_4$ in 25.0 cm^3 of **FA 4** = 7.75×10^{-4} mol [1]

(d) **(i)** Use your answer from **(b)** to calculate the number of moles of potassium manganate(VII), **FA 3**, required to react with 25.0 cm^3 of **FA 4** in **Titration One**.

No. of moles of $KMnO_4$ required
= $0.0200 \times 19.50 \times 10^{-3}$
= 3.90×10^{-4} mol

moles of $KMnO_4$ = 3.90×10^{-4} mol [1]

(ii) The equation for the reaction between acidified manganate(VII) ions and ethanedioate ions is shown below.

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

Use your answer from **(d)(i)** to calculate the total number of moles of ethanedioate ions in 25.0 cm^3 of **FA 4**.

Total no. of moles of $C_2O_4^{2-}$ present in 25.0 cm^3
= 5/2 × no. of moles of MnO_4^- required
= 5/2 × 3.90×10^{-4}
= 9.75×10^{-4} mol

total moles of $C_2O_4^{2-}$ in 25.0 cm^3 of **FA 4** = 9.75×10^{-4} mol [1]

(iii) Use your answers from **(c)(iii)** and **(d)(ii)** to calculate the number of moles of ethanedioate ions which came from the sodium ethanedioate in 25.0 cm^3 of **FA 4**.

No. of moles of $C_2O_4^{2-}$ from $Na_2C_2O_4$ in 25.0 cm^3
= $(9.75 \times 10^{-4}) - (7.75 \times 10^{-4})$
= 2.00×10^{-4} mol

moles of $C_2O_4^{2-}$ from $Na_2C_2O_4$ in 25.0 cm^3 of **FA 4** = 2.00×10^{-4} mol [1]

(iv) Hence, calculate the concentration, in g dm^{-3}, of sodium ethanedioate in **FA 1**.

No. of moles of $Na_2C_2O_4$ in 250 cm^3 of solution of **FA 4**
= $2.00 \times 10^{-4} \times \frac{250}{25.0}$
= 2.00×10^{-3} mol

Mass of $Na_2C_2O_4$ in 42.70 cm^3 of **FA 1**
= $2.00 \times 10^{-3} \times (2 \times 23.0 + 2 \times 12.0 + 4 \times 16.0)$
= 0.268 g

Concentration, in g dm^{-3}, of $Na_2C_2O_4$ in **FA 1**
= $0.268 \div \frac{42.50}{1000}$
= 6.31 g dm^{-3}

Concentration of sodium ethanedioate is 6.31 g dm^{-3} [2]

(e) Explain why the decolourisation of **FA 3** in **Titration One** is initially slow but becomes faster as the reaction proceeds.

Sufficient auto-catalyst, $Mn^{2+}(aq)$ ions, will be formed to catalyse the reaction to occur faster. [1]

[Total: 16]

2 This question concerns the solubility of **FA 5**, potassium bromate(V), $KBrO_3$, in water.

The **solubility** of a substance in water is defined as: the mass of substance that will dissolve in and just saturate 100 g of water at a particular temperature. When a solution is saturated, the dissolved solid is in equilibrium with undissolved solid. When a solution of potassium bromate(V) is cooled, it becomes saturated when crystals form in the solution.

You are to investigate how the solubility of **FA 5** in water varies with temperature.

You are provided with the following materials
- weighing bottle labelled **FA 5**, containing potassium bromate(V), $KBrO_3$, and
- deionised water.

Read through the instructions before starting any practical work.

(a) 1. Prepare a hot water bath by filling a 250 cm^3 beaker half full of water and heat it over the Bunsen burner until almost boiling. Turn off the Bunsen burner.
2. Weigh an empty boiling tube.
3. Add the contents of the weighing bottle labelled **FA 5** to the boiling tube.
4. Reweigh the boiling tube and its contents.
5. Record, in an appropriate form below, your weighings and the mass of **FA 5** used.

Mass of boiling tube + **FA 5** / g	32.360
Mass of boiling tube / g	30.299
Mass of **FA 5** used / g	2.061

6. Use the 10 cm^3 measuring cylinder to transfer 8.0 cm^3 of deionised water to the weighed boiling tube containing **FA 5**.
7. Use the clamp as a holder for the boiling-tube. Take care not to break the tube by clamping it too tightly.
8. Warm the tube carefully in the water bath, while stirring the contents with a thermometer, until all the solid has dissolved. (Take care that you do not break the thermometer bulb or the tube while stirring.)
9. Remove the tube from the water bath and attach the clamp to a retort stand.
10. Let the tube cool and continue to stir gently with the thermometer.
11. Watch the solution carefully. Note and record (**on the next page**) the temperature at which you **first** notice crystals forming in the solution.
12. If you are uncertain about the temperature when crystals first form, warm the tube again for a few moments and repeat the cooling.
13. As soon as you have recorded the temperature, add a further 2.0 cm^3 of deionised water to the tube using the 10 cm^3 measuring cylinder.
14. Warm the tube in the water bath to re-dissolve the solid and cool as before.
15. Note and record (**on the next page**) the temperature at which crystals now form in the solution. This will be lower than the temperature obtained with 8.0 cm^3 of water.
16. Repeat the addition of 2.0 cm^3 of deionised water, the heating and the cooling, until you have **four** readings in total.
17. In an appropriate form in the space below, record the following.
    - the total volume of deionised water in the boiling-tube
    - the temperature at which crystals first appeared for each solution

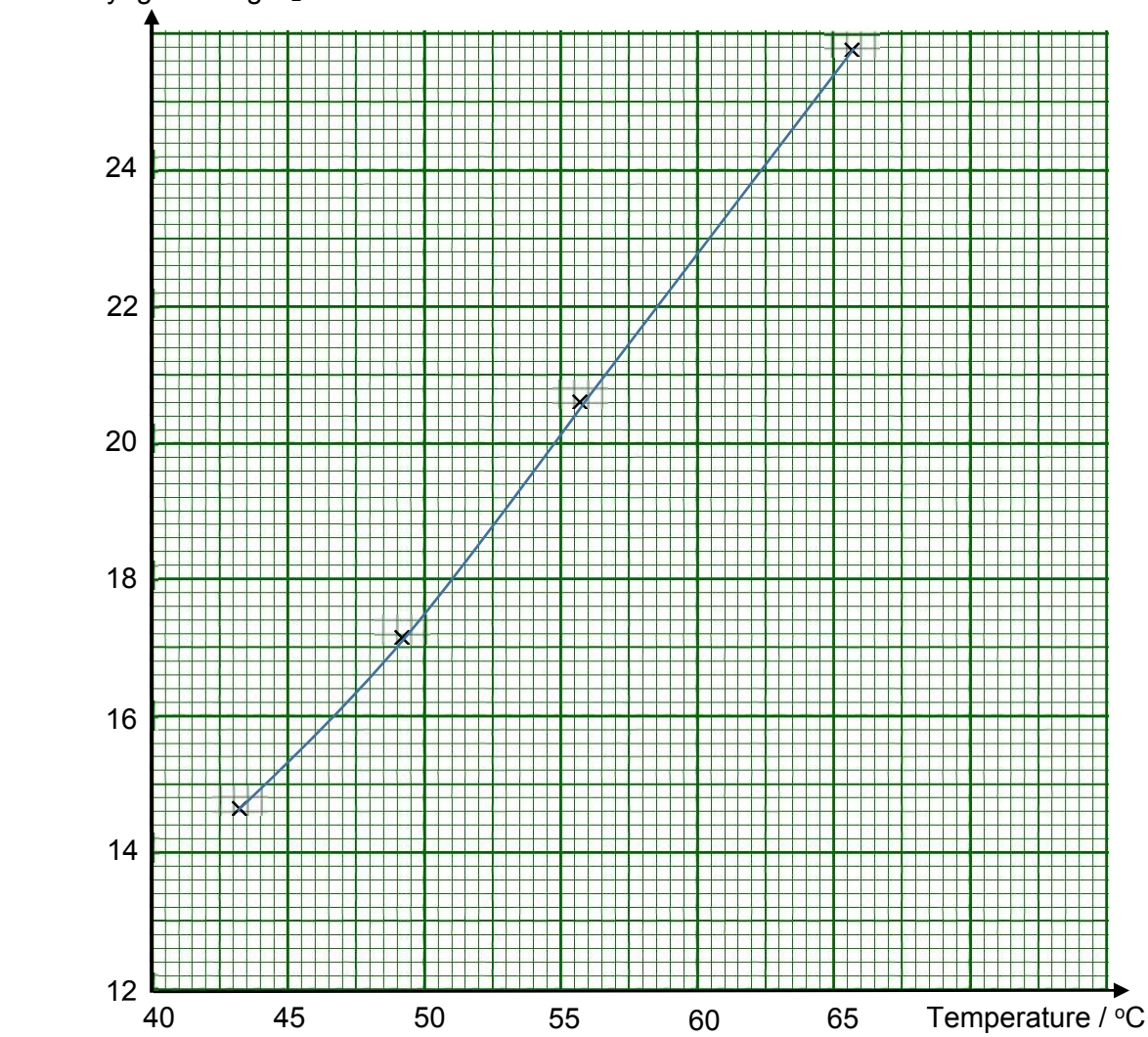
- the **solubility** (in grams of solid per 100 g of water) which can be calculated using the following formula

$$\text{solubility} = \frac{100}{\text{volume of water}} \times \text{mass of } \textbf{FA 5} \text{ dissolved}$$

Volume of water/ cm^3	Temperature at first appearance of crystals/ °C	Solubility/ g per 100 g H_2O
8.0	65.5	25.8
10.0	55.5	20.6
12.0	49.0	17.2
14.0	43.0	14.7

[6]

(b) On the grid, plot a graph of solubility against temperature and draw an appropriate line through the points. Do **not** start at zero on either axis. You will need to be able to find the solubility of **FA 5** at 55 °C.



[3]

(c) Use your solubility curve in **(b)** to answer the following.

(i) Explain if dissolving potassium bromate(V), $KBrO_3$, is an exothermic or endothermic process.

$$KBrO_3(s) \rightleftharpoons K^+(aq) + BrO_3^-(aq)$$

The graph has a positive gradient. Hence, solubility of $KBrO_3$ increases with temperature. Dissolving $KBrO_3$ is an endothermic process.

[1]

(ii) Calculate the solubility of potassium bromate(V), $KBrO_3$, at 55 °C in mol dm^{-3}.
[Given that the M_r of $KBrO_3$ is 167.0.]

From the graph, solubility of $KBrO_3$ is 20.2 g per 100 g of H_2O.

Solubiity in mol dm^{-3} = $\frac{20.2}{167.0} \times \frac{1000}{100}$ = 1.21 mol dm^{-3}

[1]

(iii) Hence, calculate the K_{sp} of potassium bromate(V) at this temperature, giving the units.

$K_{sp} = [K^+][BrO_3^-] = (1.21)^2 = 1.46$ mol^2 dm^{-6} (ecf)

[2]

(d) Student **A** claims that both the solubility and solubility product of potassium bromate(V) will decrease with the addition of some solid potassium nitrate at a particular temperature. Comment on the student's claim.

On addition of potassium nitrate, concentration of K^+ increases and the position of equilibrium shift left. Solubility of $KBrO_3$ decreases.

However, solubility product is only dependent on temperature/independent of temperature and hence remains unchanged.

[1]

(e) A literature value for the solubility of potassium bromate(V) is 13.1 g per 100 g of water at 40 °C. Student **A** followed the instructions in **(a)** and obtained a solubility value for $KBrO_3$ to be 15.0 g per 100 g of water at the same temperature. Calculate the magnitude of the percentage experimental error for student **A**'s measurement.

Experimental error = $\left| \frac{13.1 - 15.0}{13.1} \right|$ x 100% = 14.5%

[1]

[Total:15]

3 Planning

FA 6 is a powdered mixture of mica and iron(III) oxide, Fe_2O_3, which is used as ingredients in mineral makeup such as eye shadow or blusher. Iron(III) oxide adds a red colour to makeup while mica gives makeup a light reflecting quality.

The reaction between Fe_2O_3, and ethanedioic acid, $H_2C_2O_4$, is exothermic as shown.

$$Fe_2O_3(s) + 3H_2C_2O_4(aq) \rightarrow Fe_2(C_2O_4)_3(aq) + 3H_2O(l) \qquad \Delta H < 0$$

In a thermometric titration, the end–point is reached when the maximum temperature change occurs. A thermometric titration between Fe_2O_3 and $H_2C_2O_4$ can be carried out to find out the

percentage by mass of iron(III) oxide in **FA 6**. The temperature of the reaction mixture is monitored when a certain mass of **FA 6** is added into a fixed volume of aqueous $H_2C_2O_4$ solution. The experiment is then repeated using different masses of **FA 6**. You may assume mica remains unchanged in this experiment.

The data obtained is plotted and two best–fit graph lines are drawn. One line is drawn using data before the end–point and the second line using the remaining data. These lines are then extrapolated until they intersect.

(a) Using the information given above, you are required to write a plan for a thermometric titration between **FA 6** and aqueous $H_2C_2O_4$ solution.

You are provided with

- 500 cm^3 of 1.00 mol dm^{-3} of $H_2C_2O_4$,
- 30 g of solid **FA 6** containing approximately 90% by mass of Fe_2O_3,
- styrofoam cup,
- thermometer, and
- apparatus commonly found in a college laboratory

In your plan you should include details of

- justification of specific quantities of reactants that you would use,
- the apparatus you would use and the procedure you would follow,
- a sketch of the graph you would expect to obtain, with the end–point clearly labelled, and
- how the data obtained from the graph would be used to calculate the actual percentage by mass of Fe_2O_3 in **FA 6**

(i) <u>Justification of quantities of reactants used for the experiments</u>

Assume 50 cm^3 of 1.00 mol dm^{-3} $H_2C_2O_4$ is used.
Amount of $H_2C_2O_4$ used in 50 cm^3 of $H_2C_2O_4$ = (50/1000) x 1.00
= 0.0500 mol

Mass of Fe_2O_3 required to react completely with 50 cm^3 of $H_2C_2O_4$
= (0.05/3) x (55.8x2 + 16.0x3)
= 2.66 g

Approximate mass of FA 6 required to react with 50 cm^3 of $H_2C_2O_4$
= (2.66/90) x 100
= 2.96 g

Experiment	Volume of $H_2C_2O_4$/cm^3	Mass of FA 6 /g	ΔT / ^oC
1	50.0	1.000	
2	50.0	1.500	
3	50.0	2.000	
4	50.0	2.500	
5	50.0	3.000	
6	50.0	3.500	
7	50.0	4.000	
8	50.0	4.500	
9	50.0	5.000	

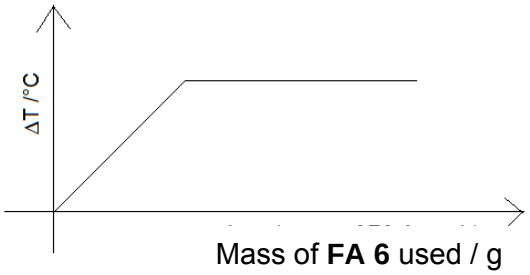
(ii) <u>Procedure</u>

1. Use a 50 cm^3 burette to introduce 50.00 cm^3 of 1 mol dm^{-3} of $H_2C_2O_4^-$ into a dry styrofoam cup.

2. Place the thermometer into the solution and record initial temperature of the solution.
3. Weigh accurately 1.000 g of FA 6 and put it into the Styrofoam cup.
4. Stir gently with the thermometer. Record the highest temperature rise.
5. Reweigh the emptied weighing bottle to get the actual mass of FA 6 that has been used.
6. Repeat steps 1 to 5 for experiment 2 to experiment 9.

(iii) Sketch of graph

1. Plot a graph of ΔT versus actual mass of FA 6 added.
2. Correct shape and indicate the intersection point will give us the mass of FA 6 needed to completely react with $H_2C_2O_4$. This is when end–point has been reached.



(iv) Calculations

Percentage mass of Fe_2O_3 in FA 6
= (2.66/Mass of FA 6 obtained from the graph) x 100%

[10]

(b) A student suggested that using a burette to measure the 25.0 cm^3 of acid would give a more accurate result than using a pipette. The percentage error of a 25.0 cm^3 pipette is 0.24%. Is the student correct? Explain your answer.

Percentage error for burette = $\frac{2 \times 0.05}{25.0}$ x 100 = 0.40 %
No, the pipette is more accurate. This is because the percentage error of using the burette to measure the 25.0 cm^3 of acid is greater than the pipette.

[2]

[Total: 12]

4 You are given samples of eight aqueous solutions.

FB 1 containing bromide ions, Br^-

FB 2 containing bromate(I) ions, BrO^-

FB 3 containing iron(III) ions, Fe^{3+}

FB 4 containing hydrogen peroxide, H_2O_2

FB 5 containing aqueous iodine, I_2

FB 6 containing iodide ions, I^-

FB 7 containing potassium manganate(VII), $KMnO_4$

You are also given hexane, sulfuric acid and starch solution.

You will perform a series of tests to investigate if any redox reaction has occurred.

You will make deductions about the relative oxidising powers of different substances.

For example, under appropriate conditions, chlorine water will oxidise iodide ions to iodine.

$$Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$$

Deduction would be that chlorine is a stronger oxidising agent than iodine. This is represented by $Cl_2 > I_2$.

In another example, iodine will **not** oxidise bromide ions to bromine. So, the deduction would be that bromine is a stronger oxidising agent than iodine. This is represented by $Br_2 > I_2$.

(a) Perform the tests described in **Table 1**. Record your observations in the spaces provided. If there is no observable reaction, write **'no reaction'** in the observations column.

Your observations should allow you to deduce the relative oxidising powers of the substances involved.

State the relative oxidising power of the substances involved under the deductions column of **Table 1**. They should be written in the form e.g. $Cl_2 > I_2$.

[10]

Table 1

Tests		Observations	Deductions
1	Add 1 cm^3 of **FB 6** to a test–tube. Then add 8 drops of **FB 3**, followed by 5 drops of starch solution.	• Solution turned brown. • Solution turned blue–black with starch.	• $Fe^{3+} > I_2$
2	Add 1 cm^3 of **FB 3** to a test–tube. Then, add 1 cm^3 of **FB 1** followed by 1 cm^3 of hexane. Shake the mixture.	• No reaction.	• $Br_2 > Fe^{3+}$
3	Add 1 cm^3 of **FB 4** to a test–tube. Then add 8 drops of **FB 5**.	• No reaction.	• $O_2 > I_2$
4	Add 1 cm^3 of **FB 4** to a test–tube. Now add 1 cm^3 of dilute sulfuric acid. Then add 1 cm^3 of **FB 1** followed by 1 cm^3 of hexane. Shake the mixture.	• Aqueous layer turned yellow/orange. • Organic layer turned orange/red/brown	• $H_2O_2 > Br_2$
5	Add 1 cm^3 of **FB 1** to a test–tube. Now add 4 drops of dilute sulfuric acid. Then add 1 cm^3 of **FB 2**.	• Solution turned orange/red/yellow.	• $BrO^- > Br_2$
6	Add 1 cm^3 of **FB 4** to a test–tube. Then add 1 cm^3 of **FB 3**. Make observations for about 2 minutes before recording your results.	• Solution turned brown. • Effervescence (or bubbles) observed. • O_2 gas • relighted a glowing splint. • Solution remained yellow/brown.	• $Fe^{3+} > O_2$

7	Add 1 cm^3 of **FB 1** in a test–tube followed by 1 cm^3 of dilute sulfuric acid. Then add 8 drops of **FB 7** and 1 cm^3 of hexane. Shake the mixture.	• Aqueous layer turned yellow/orange/red. • Organic layer turned orange/brown/red.	• MnO_4^- > Br_2

(b) Using your answers in **Table 1**, arrange the following substances below in descending order of their oxidising power.

Br_2, BrO^-, Fe^{3+}, I_2, O_2

BrO^- > Br_2 > Fe^{3+} > O_2 > I_2

[1]

(c) What are the roles of **FB 3** in tests **1** and **6** respectively?

In test 1, it is acting as an oxidising agent (as it is reduced) while in test 6, it is a catalyst.

[1]

[Total: 12]

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

(b) Reactions of anions

anion	*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

The Periodic Table of Elements

Group																	
1	2											13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
3 Li lithium 6.9	4 Be beryllium 9.0		Key atomic number atomic symbol name relative atomic mass									5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2
11 Na sodium 23.0	12 Mg magnesium 24.3	3	4	5	6	7	8	9	10	11	12	13 Al aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium –	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3
55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –		114 Fl flerovium –		116 Lv livermorium –		

lanthanoids	57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.2	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0
actinoids	89 Ac actinium –	90 Th thorium 232.0	91 Pa protactinium 231.0	92 U uranium 238.0	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –

[Turn over

Apparatus and Chemicals for each candidate

Apparatus list

1. 2×50.00 cm^3 burette (to invert burette and put on retort stand)
2. 1×25.0 cm^3 pipette
3. 2×250 cm^3 conical flask
4. 2 x 150 cm^3 beaker
5. 1×250 cm^3 beaker
6. 1×250 cm^3 graduated flask
7. 1×10 cm^3 measuring cylinder
8. 1×50 cm^3 measuring cylinder
9. $1 \times$ glass rod
10. Plastic bag: $6 \times$ dropper, $1 \times$ long wooden splint, $2 \times$ paper towel, blue & red litmus papers, filter paper strips
11. Plastic bag: 5 test tubes
12. $1 \times$ dry boiling tube (pack with 5 test tubes)
13. $1 \times$ pipette filler
14. $1 \times$ retort stand and burette clamp
15. $1 \times$ clamp – to hold a boiling tube (to mount on retort stand)
16. $1 \times$ white tile
17. $2 \times$ filter funnel
18. $1 \times$ Bunsen burner
19. $1 \times$ wire gauge (heat proof mat)
20. $1 \times$ tripod stand
21. $1 \times$ lighter
22. $1 \times$ alcohol thermometer (-10 °C to 110 °C at 1 °C)
23. $1 \times$ wash bottle containing distilled water
24. $1 \times$ test tube rack
25. $1 \times$ test tube holder
26. $1 \times$ visualiser per lab
27. electronic weighing balance (5 per lab)
28. synchronise all the clocks in the 8 labs

Chemical List

29. **FA 1** is 0.250 mol dm^{-3} sodium ethanedioate, $Na_2C_2O_4$ (100 cm^3 in a vial labelled **FA 1**)
30. **FA 2** is approximately 2 mol dm^{-3} sulfuric acid, H_2SO_4 (100 cm^3 in a vial labelled **FA 2**)
31. **FA 3** is 0.0200 mol dm^{-3} potassium manganate(VII), $KMnO_4$ (120 cm^3 in a vial labelled **FA 3**)
32. **FA 5** is solid potassium bromate(V), $KBrO_3$ (1.90–2.10 g in a weighing bottle labelled **FA 5**)
33. One rack of bench reagents (only for show, no need to prepare fresh solutions)
34. 1 red tub per student for 3 shifts containing **FB 1** to **FB 8**, sulfuric acid, hexane and starch

Label	Capacity per student in capped reagent bottle
FB 1	10 cm^3 of 0.1 mol dm^{-3} potassium bromide, KBr, made by dissolving about 11.9 g of KBr in 1 dm^3 of deionised water
FB 2	10 cm^3 of 10% aqueous potassium bromate(I), KBrO
FB 3	5 cm^3 of 0.2 mol dm^{-3} ammonium iron(III) sulfate, $NH_4Fe(SO_4)_2 \cdot 12H_2O$, made by dissolving about 96.4 g of $NH_4Fe(SO_4)_2 \cdot 12H_2O$ in 1 dm^3 of deionised water
FB 4	10 cm^3 of 20 volume hydrogen peroxide, H_2O_2 (containing 200 cm^3 of freshly opened 100 volume hydrogen peroxide, made up to 1 dm^3 with deionised water)
FB 5	5 cm^3 of 0.01 mol dm^{-3} aqueous iodine, I_2, made by dissolving about 2.5 g of iodine and about 8 g of potassium, KI, in 1 dm^3 of deionised water
FB 6	5 cm^3 of 0.1 mol dm^{-3} potassium iodide, KI, made by dissolving about 16.6 g of KI in 1 dm^3 of deionised water
FB 7	5 cm^3 of 0.0200 mol dm^{-3} potassium manganate(VII) (same solution as **FA 3**)
sulfuric acid	one labelled reagent bottle containing 2 mol dm^{-3} H_2SO_4 (same solution as **FA 2**, one full reagent bottle)
hexane	hexane (one full reagent bottle)
starch	starch (one full reagent bottle)