Surname	StemUp Model Solution	Centre Number	Candidate Number
First name(s)			2

GCE A LEVEL

A410U30-1

S23-A410U30-1

FRIDAY, 23 JUNE 2023 – MORNING

CHEMISTRY – A level component 3

Chemistry in Practice

1 hour 15 minutes

For Examiner's use only		
Question	**Maximum Mark**	**Mark Awarded**
1.	**10**	
2.	**17**	
3.	**6**	
4.	**10**	
5.	**17**	
Total	**60**	

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01

ADDITIONAL MATERIALS

- A calculator
- **Data Booklet** supplied by WJEC.

INSTRUCTIONS TO CANDIDATES

Use black ink or black ball-point pen. Do not use gel pen or correction fluid.

You may use a pencil for graphs and diagrams only.

Write your name, centre number and candidate number in the spaces at the top of this page.

Answer **all** questions.

Write your answers in the spaces provided in this booklet. If you run out of space, use the additional page(s) at the back of the booklet, taking care to number the question(s) correctly.

INFORMATION FOR CANDIDATES

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

The maximum mark for this paper is 60.

Your answers must be relevant and must make full use of the information given to be awarded full marks for a question.

The assessment of the quality of extended response (QER) will take place in **Q3**.

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Answer **all** questions.

1. The hydration of anhydrous sodium thiosulfate is represented by the following equation.

$$Na_2S_2O_3(s) + 5H_2O(l) \longrightarrow Na_2S_2O_3.5H_2O(s) \qquad \Delta H_1$$

It is not possible to measure the enthalpy change for this reaction directly. However, it is possible to measure the enthalpy changes of solution when hydrated sodium thiosulfate and anhydrous sodium thiosulfate are in turn dissolved in water.

$$Na_2S_2O_3.5H_2O(s) + aq \longrightarrow Na_2S_2O_3(aq) \qquad \Delta H_2$$

$$Na_2S_2O_3(s) + aq \longrightarrow Na_2S_2O_3(aq) \qquad \Delta H_3$$

These enthalpy values can then be used to calculate ΔH_1 using Hess's Law.

Part I: Enthalpy change of solution of hydrated sodium thiosulfate, $Na_2S_2O_3.5H_2O$ (ΔH_2)

An experiment was carried out to determine ΔH_2.

- 50.0 cm^3 of water was added to an insulated polystyrene cup and the temperature was recorded every 30 seconds.
- At exactly 2 minutes, 12.38 g of hydrated sodium thiosulfate was added to the water in the cup and stirred.
- The temperature was recorded every 30 seconds for a further 8 minutes.
- The resulting temperature / time plot is shown opposite.

(a) Draw appropriate lines on the graph and use these to determine the maximum temperature change, ΔT. [2]

ΔT = 9.7 °C

Temperature before mixing = 19.8 °C

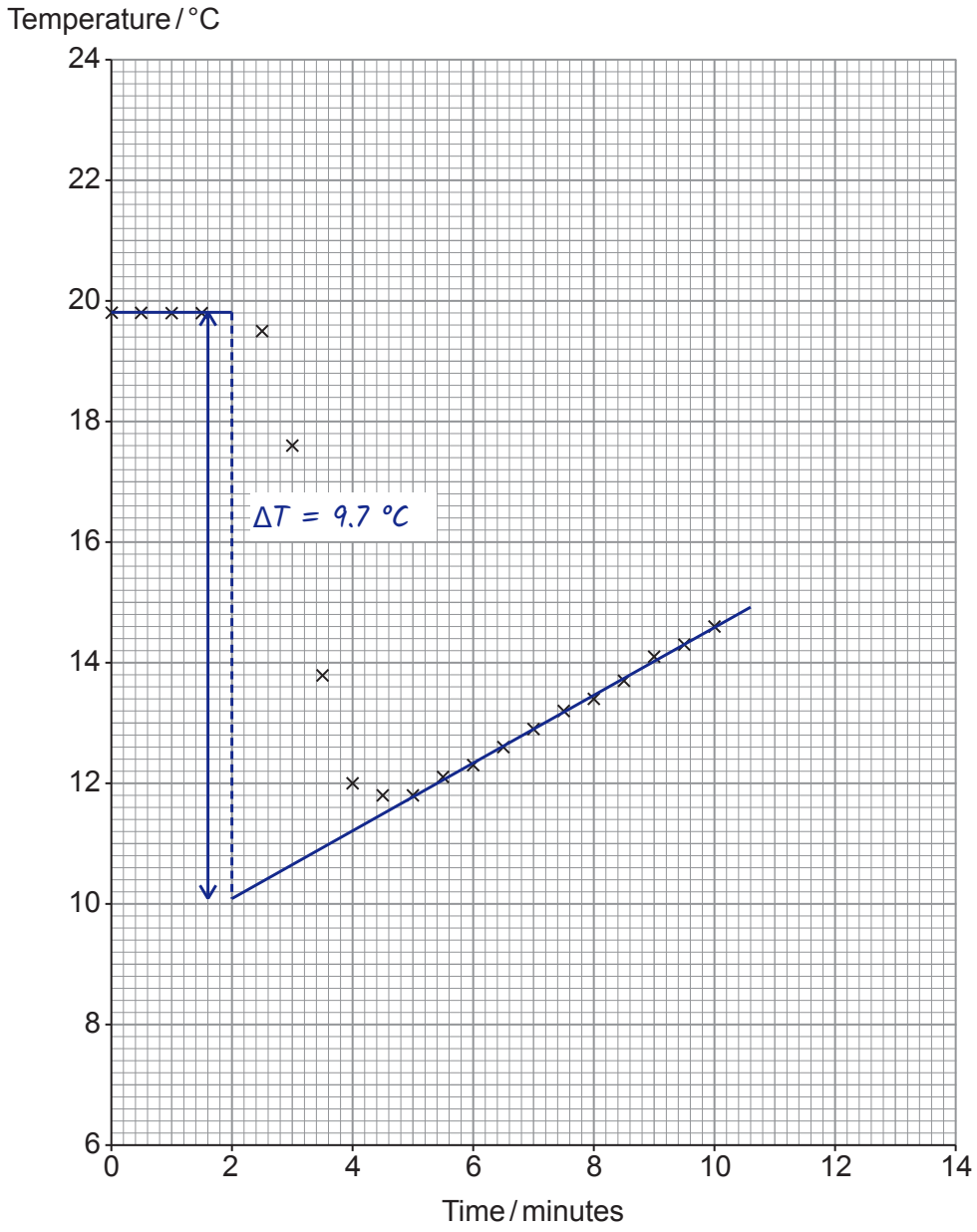
Extrapolated temperature at 2 min = 10.1 °C

ΔT = 19.8 − 10.1 = 9.7 °C

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(b) Calculate the molar enthalpy change of solution of hydrated sodium thiosulfate, ΔH_2. [2]

$n(Na_2S_2O_3.5H_2O) = 12.38 / 248.3 = 0.0499$ mol

$Q = mc\Delta T = 50.0 \times 4.18 \times 9.7 = 2027$ J $= 2.027$ kJ

$\Delta H_2 = +2.027 / 0.0499$ (endothermic, temperature fell)

ΔH_2 = ... +40.6 ..... kJ mol^{-1}

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(c) Give **one** assumption made when finding the value of ΔH_2 from the experimental results. [1]

No heat is gained from the surroundings (the cup is a perfect insulator).

(d) Using your result in (a), deduce the maximum temperature change expected if twice the volume of water is used with the same mass of hydrated sodium thiosulfate. [1]

4.85 °C

Part II: Enthalpy change of solution of anhydrous sodium thiosulfate, $Na_2S_2O_3$ (ΔH_3)

To determine ΔH_3, a student measured 50.0 cm^3 of water into an insulated polystyrene cup and added a known mass of anhydrous sodium thiosulfate whilst stirring. The temperature recorded increased from 20.2°C to a maximum of 23.6°C.

(e) If the value obtained for the molar enthalpy change of solution of anhydrous sodium thiosulfate, ΔH_3, was –14.2 kJ mol^{-1}, calculate the mass of anhydrous sodium thiosulfate added. [2]

$Q = mc\Delta T = 50.0 \times 4.18 \times (23.6 - 20.2) = 50.0 \times 4.18 \times 3.4 = 710.6$ J

$n = Q / \Delta H = 710.6 / 14\,200 = 0.0500$ mol

$M_r (Na_2S_2O_3) = 158.2$

mass = 0.0500×158.2

Mass = 7.91 g

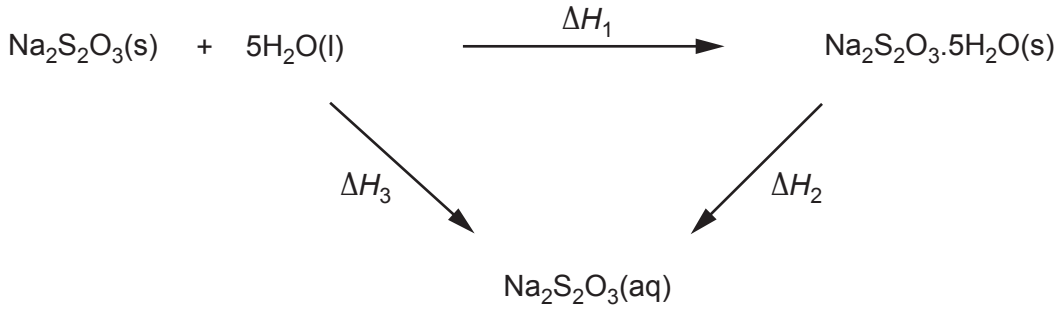
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Part III: Enthalpy change of hydration of anhydrous sodium thiosulfate (ΔH_1)

(f) The Hess's cycle connecting the three enthalpy terms is given below. Use this to calculate the value of ΔH_1. [1]



$\Delta H_1 + \Delta H_2 = \Delta H_3$

$\Delta H_1 = \Delta H_3 - \Delta H_2 = -14.2 - (+40.6)$

ΔH_1 = −54.8 kJ mol^{-1}

(g) Suggest **one** reason why the enthalpy change for this reaction cannot be determined directly. [1]

It is not possible to prevent some of the anhydrous sodium thiosulfate dissolving in the water added.

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2. This question considers two different methods of preparing soluble salts.

Method 1: Solid and solution

In this method, an insoluble solid is added to a suitable solution to form a soluble salt.

One example is the reaction of excess copper(II) carbonate with hydrochloric acid to form copper(II) chloride.

$$CuCO_3(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + CO_2(g) + H_2O(l)$$

(a) State why this method uses excess copper(II) carbonate and not excess hydrochloric acid. [1]

Excess insoluble $CuCO_3$ can be separated by filtration, whereas excess acid cannot easily be removed and would contaminate the salt solution.

(b) Describe the practical steps for making pure, crystalline copper(II) chloride from solid copper(II) carbonate and hydrochloric acid. [3]

Add the $CuCO_3$ to the hydrochloric acid a little at a time, with stirring (warm gently).

Continue adding $CuCO_3$ until it is in excess: no more fizzing occurs and some solid remains undissolved.

Filter off the excess unreacted $CuCO_3$, keeping the filtrate.

Evaporate the filtrate to reduce its volume (until crystals start to form), then leave it to cool and crystallise. Filter off the crystals and dry them between filter papers.

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(c) (i) In an experiment to prepare copper(II) chloride, a student was given 60.0 cm^3 of 0.500 mol dm^{-3} hydrochloric acid.

Calculate the minimum mass of copper(II) carbonate (M_r 123.5) required for complete reaction. [2]

$n(HCl) = 0.500 \times 60.0 / 1000 = 0.0300$ mol

$n(CuCO_3) = 0.0300 / 2 = 0.0150$ mol

mass $= 0.0150 \times 123.5$

Mass = 1.85 g

(ii) Hydrated copper(II) chloride, $CuCl_2.xH_2O$, is formed as a green crystalline solid.

Calculate the value of x if 2.56 g of pure $CuCl_2.xH_2O$ was obtained. [2]

$n(CuCl_2.xH_2O) = n(CuCO_3) = 0.0150$ mol

$M_r = 2.56 / 0.0150 = 170.7$

mass of $H_2O = 170.7 - 134.5 = 36.2$

$x = 36.2 / 18.02 = 2.0$

x = 2

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Method 2: Solution and solution

In this method, exact quantities of both reactants are used so that the final solution only contains the required salt.

A method that allows this is titration. One example is the titration of aqueous sodium carbonate with hydrochloric acid to form sodium chloride.

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

Methyl orange can be used as an indicator in this titration.

A student proposes to make solid sodium chloride by the following method.

Step 1	Fill a burette to the 0.00 cm^3 mark with approximately 0.50 $mol\,dm^{-3}$ hydrochloric acid solution
Step 2	Carefully transfer 25.0 cm^3 of 0.250 $mol\,dm^{-3}$ sodium carbonate solution to a conical flask
Step 3	Add 2–3 drops of methyl orange indicator
Step 4	Add the hydrochloric acid to the conical flask until the endpoint is near
Step 5	Continue to add acid drop by drop, washing down the walls of the flask with deionised water, until the indicator just changes colour
Step 6	Evaporate the solution formed to leave solid sodium chloride

(d) (i) State why it is not necessary to fill the burette to exactly zero (step 1). [1]

The volume used is measured by difference (final reading – initial reading), so any starting reading can be used.

(ii) State why the walls of the flask should be washed down with deionised water before the indicator changes colour (step 5). [1]

Any drops of Na_2CO_3 or HCl that splashed onto the walls are washed back in so that they take part in the reaction.

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(iii) The teacher said that the method would not give pure sodium chloride.

State how the student should adapt the method in order to get pure solid sodium chloride. [1]

Repeat the titration without the indicator, using 25.0 cm^3 of Na_2CO_3 and adding exactly the titre volume of HCl found, then evaporate this solution.

(iv) Suggest **one** further improvement to the method. [1]

Repeat the titration to obtain concordant titres (within 0.10 cm^3) and use their mean.

(e) The student then decides to prepare a sample of solid potassium sulfate by titration, using the improved method.

He transfers 25.0 cm^3 of 0.250 mol dm^{-3} potassium carbonate solution to a conical flask and titrates against approximately 0.50 mol dm^{-3} sulfuric acid.

(i) Give the equation for the reaction of potassium carbonate with sulfuric acid. [1]

$K_2CO_3 + H_2SO_4 \rightarrow K_2SO_4 + CO_2 + H_2O$

(ii) Explain why, in the preparation of solid potassium sulfate, the percentage error in the burette reading will be considerably greater than in the preparation of sodium chloride.

You should refer to the equations for the preparation of both salts in your answer. [2]

Mole ratio Na_2CO_3 : HCl is 1 : 2, whereas K_2CO_3 : H_2SO_4 is 1 : 1 (and both acids are about 0.50 mol dm^{-3}).

So the volume of H_2SO_4 used is half the volume of HCl, therefore the percentage error in the burette reading is doubled.

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(iii) 25.0 cm^3 of the 0.250 mol dm^{-3} potassium carbonate solution was used in the preparation of potassium sulfate.

Calculate the volume, in cm^3, of carbon dioxide gas formed at 298 K and 1 atm.

[2]

$n(K_2CO_3) = 0.250 \times 25.0 / 1000 = 0.00625$ mol

$n(CO_2) = n(K_2CO_3) = 0.00625$ mol

volume = $0.00625 \times 24.5 = 0.153$ dm^3

Volume = 153 cm^3

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3. Transition metals display basic or amphoteric behaviour.

Describe how you would use sulfuric acid and sodium hydroxide solution to determine whether iron(II) oxide and chromium(III) oxide are basic or amphoteric.

Include observations and equations in your answer. [6 QER]

Iron(II) oxide: add the black FeO powder to dilute sulfuric acid and warm. It dissolves to give a pale green solution of Fe^{2+} ions: $FeO + H_2SO_4 \rightarrow FeSO_4 + H_2O$.

Add aqueous NaOH dropwise to this solution: a green precipitate forms, $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$. In excess NaOH(aq) the green precipitate does not dissolve (FeO also does not react with NaOH solution). Fe^{2+} reacts only with acids, so iron(II) oxide is a basic oxide.

Chromium(III) oxide: add the green Cr_2O_3 to sulfuric acid and warm. It dissolves to form a violet-blue / green solution of Cr^{3+} ions: $Cr_2O_3 + 3H_2SO_4 \rightarrow Cr_2(SO_4)_3 + 3H_2O$.

Add aqueous NaOH dropwise: a grey-green precipitate forms, $Cr^{3+}(aq) + 3OH^-(aq) \rightarrow Cr(OH)_3(s)$. In excess NaOH(aq) the precipitate dissolves to give a deep green solution: $Cr(OH)_3(s) + 3OH^-(aq) \rightarrow [Cr(OH)_6]^{3-}(aq)$.

Cr^{3+} reacts with both acids and bases, so chromium(III) oxide is an amphoteric oxide.

6

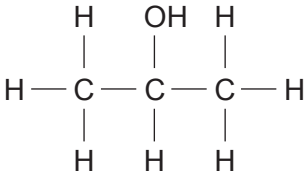
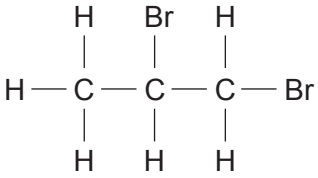
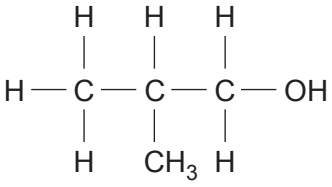
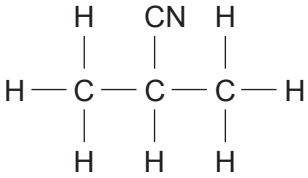
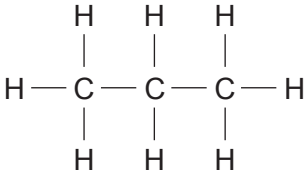
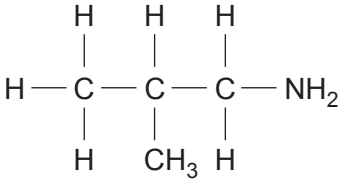
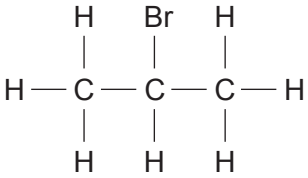
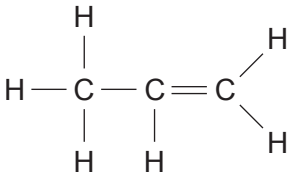
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4. You are required to show how **each** of the eight compounds (**A–H**) below can be formed from any one of the other seven compounds.

Complete the table opposite, giving the **letter** of the starting compound, the reagent(s) and any necessary reaction conditions.

You may use each starting compound once, more than once or not at all. [10]

A 	**B** 
C 	**D** 
E 	**F** 
G 	**H** 

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Starting compound	Product	Reagent(s)	Conditions
G	**A**	NaOH(aq)	aqueous solvent, heat under reflux
H	**B**	Br_2	liquid or aqueous bromine, room temperature
F	**C**	HNO_2 / nitric(III) acid ($NaNO_2$ + HCl)	room temperature
G	**D**	KCN	ethanol solvent, heat under reflux
H	**E**	H_2	150 °C, Ni catalyst
D	**F**	$LiAlH_4$	dry ethoxyethane solvent
H	**G**	HBr	room temperature
G	**H**	NaOH / KOH	ethanol solvent, heat under reflux

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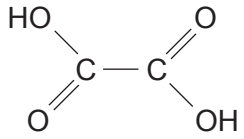
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5. Ethanedioic acid, $H_2C_2O_4$, is a dicarboxylic acid which occurs in many plants. The structure of ethanedioic acid is shown below.



Ethanedioic acid has two acidic hydrogens and therefore has two acid dissociation constants, K_{a1} and K_{a2}.

None of the second hydrogens dissociate until all of the first hydrogens have dissociated.

The acid dissociation constants have values of 5.62×10^{-2} and 5.25×10^{-5} mol dm^{-3}, but not necessarily in this order.

(a) Write equations to represent both stages in the dissociation of ethanedioic acid.

For both dissociations, give the value of the respective dissociation constant and explain your choice. [3]

Equation	Value of K_a / mol dm^{-3}
$HOOC-COOH \rightleftharpoons HOOC-COO^- + H^+$	5.62×10^{-2}
$HOOC-COO^- \rightleftharpoons {}^-OOC-COO^- + H^+$	5.25×10^{-5}

Explanation for choice of K_a value

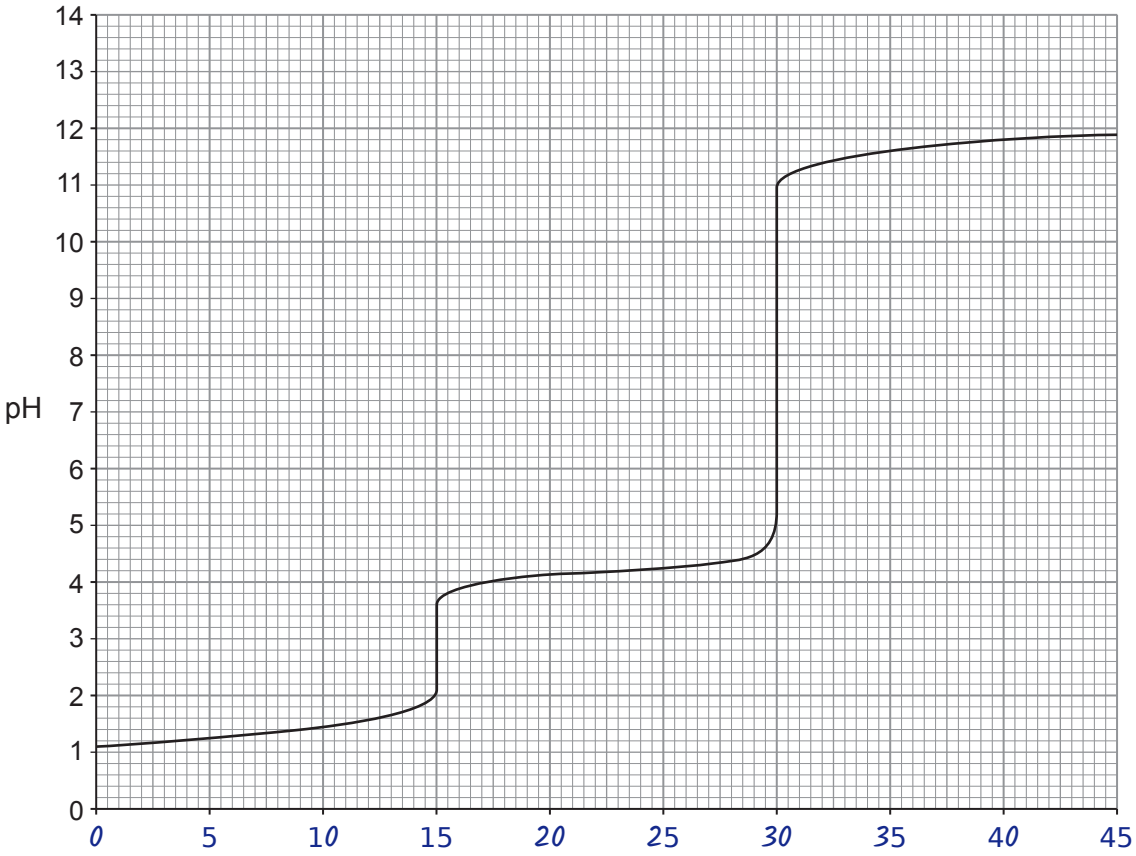
The first H^+ is removed from a neutral molecule, but the second H^+ is removed from a negatively charged ion, so it is attracted more strongly and is more difficult to remove. So K_{a2} is the smaller value, 5.25×10^{-5}, and $K_{a1} = 5.62 \times 10^{-2}$.

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(b) The titration curve shows how the pH changes as a solution of potassium hydroxide is added to 30.0 cm^3 of a 0.112 mol dm^{-3} ethanedioic acid solution.



(i) Show that the initial pH of the 0.112 mol dm^{-3} solution of ethanedioic acid is 1.1. [3]

$K_{a1} = [H^+]^2 / [(COOH)_2]$, so $[H^+]^2 = K_{a1} \times [(COOH)_2]$

$[H^+]^2 = 5.62 \times 10^{-2} \times 0.112 = 6.29 \times 10^{-3}$

$[H^+] = \sqrt{(6.29 \times 10^{-3})} = 0.0793$ mol dm^{-3}

$pH = -\log 0.0793 = 1.10 = 1.1$

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(ii) The potassium hydroxide solution had a concentration double that of the ethanedioic acid.

Complete the scale on the x-axis of the graph. Show your reasoning. [2]

n(acid) = 30.0 × 0.112 / 1000 = 0.00336 mol; [KOH] = 0.224 mol dm^{-3}

First proton 1 : 1, so n(KOH) = 0.00336 mol, V = 0.00336 / 0.224 = 15.0 cm^3 (first equivalence point); second equivalence point = 30.0 cm^3.

(c) Explain the terms **weak** and **dilute** as applied to acid solutions. [2]

Weak *only partially dissociates into H$^+$ ions in aqueous solution.*

Dilute *low concentration of acid molecules / H$^+$ ions in the solution.*

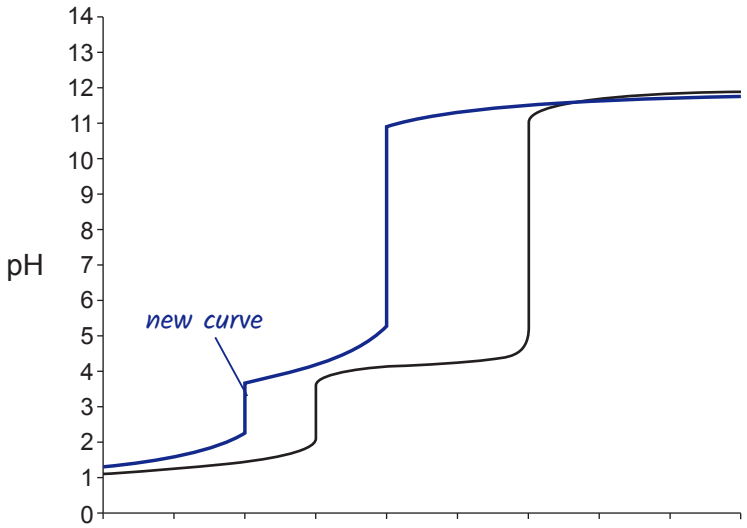
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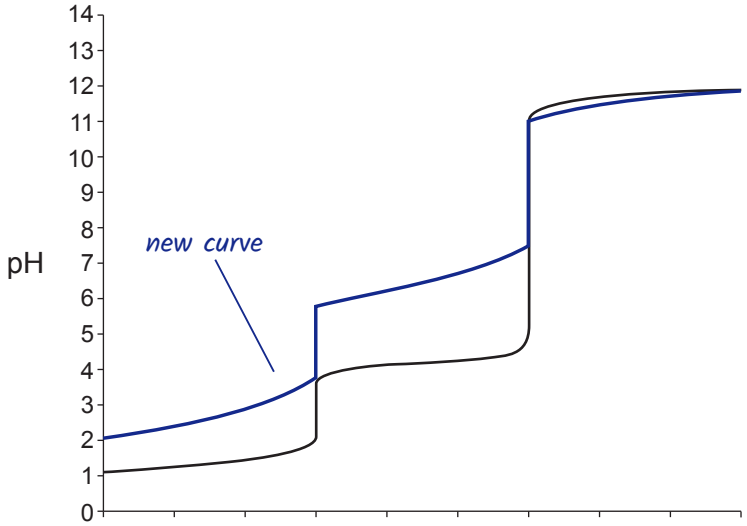
(d) (i) Sketch the titration curve obtained if the titration were repeated using 30.0 cm^3 of ethanedioic acid solution of **lower concentration**.

The concentration of the potassium hydroxide solution remains unchanged. [2]



(ii) Sketch the titration curve obtained if the titration were repeated using 30.0 cm^3 of a **different**, **weaker** dibasic acid, but of the **same concentration** as the ethanedioic acid.

Again, the concentration of the potassium hydroxide solution remains unchanged. [2]



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(e) (i) Outline how an indicator works. [1]

An indicator is a weak acid (or weak base), $HIn \rightleftharpoons H^+ + In^-$, where the protonated form (HIn) and the deprotonated form (In^-) have different colours.

(ii) Explain why two indicators are used in the titration of ethanedioic acid with potassium hydroxide solution. [2]

The two equivalence points (for K_{a1} and K_{a2}) occur at different pH values.

Each indicator's colour change range must lie within the vertical section of the curve at its equivalence point.

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END OF PAPER

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Question number	Additional page, if required. Write the question number(s) in the left-hand margin.

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CHEMISTRY – A level component 3
Data Booklet

Avogadro constant	N_A	$= 6.02 \times 10^{23}\ \text{mol}^{-1}$
molar gas constant	R	$= 8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$
molar gas volume at 273 K and 1 atm	V_m	$= 22.4\ \text{dm}^3\ \text{mol}^{-1}$
molar gas volume at 298 K and 1 atm	V_m	$= 24.5\ \text{dm}^3\ \text{mol}^{-1}$
Planck constant	h	$= 6.63 \times 10^{-34}\ \text{J s}$
speed of light	c	$= 3.00 \times 10^{8}\ \text{m s}^{-1}$
density of water	d	$= 1.00\ \text{g cm}^{-3}$
specific heat capacity of water	c	$= 4.18\ \text{J g}^{-1}\ \text{K}^{-1}$
ionic product of water at 298 K	K_w	$= 1.00 \times 10^{-14}\ \text{mol}^2\ \text{dm}^{-6}$
fundamental electronic charge	e	$= 1.60 \times 10^{-19}\ \text{C}$

temperature (K) = temperature (°C) + 273

$1\ \text{dm}^3 = 1000\ \text{cm}^3$
$1\ \text{m}^3 = 1000\ \text{dm}^3$
1 tonne = 1000 kg
$1\ \text{atm} = 1.01 \times 10^5\ \text{Pa}$

Multiple	Prefix	Symbol
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m

Multiple	Prefix	Symbol
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G

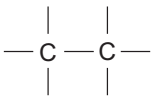
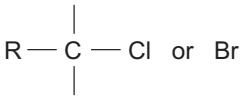
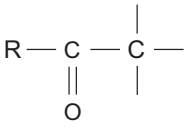
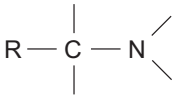
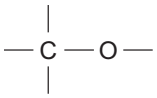
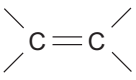
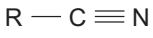
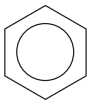
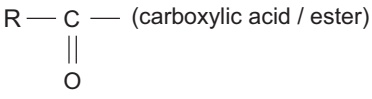
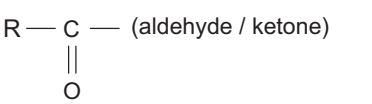
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Infrared absorption values

Bond	Wavenumber / cm^{-1}
C — Br	500 to 600
C — Cl	650 to 800
C — O	1000 to 1300
C = C	1620 to 1670
C = O	1650 to 1750
C ≡ N	2100 to 2250
C — H	2800 to 3100
O — H (carboxylic acid)	2500 to 3200 (very broad)
O — H (alcohol / phenol)	3200 to 3550 (broad)
N — H	3300 to 3500

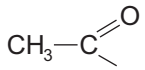
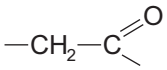
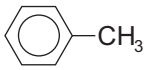
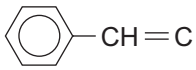
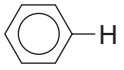
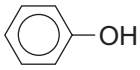
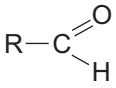
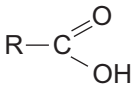
^{13}C NMR chemical shifts relative to TMS = 0

Type of carbon	Chemical shift, δ (ppm)
	5 to 40
	10 to 70
	20 to 50
	25 to 60
	50 to 90
	90 to 150
	110 to 125
	110 to 160
	160 to 185
	190 to 220

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^{1}H NMR chemical shifts relative to TMS = 0

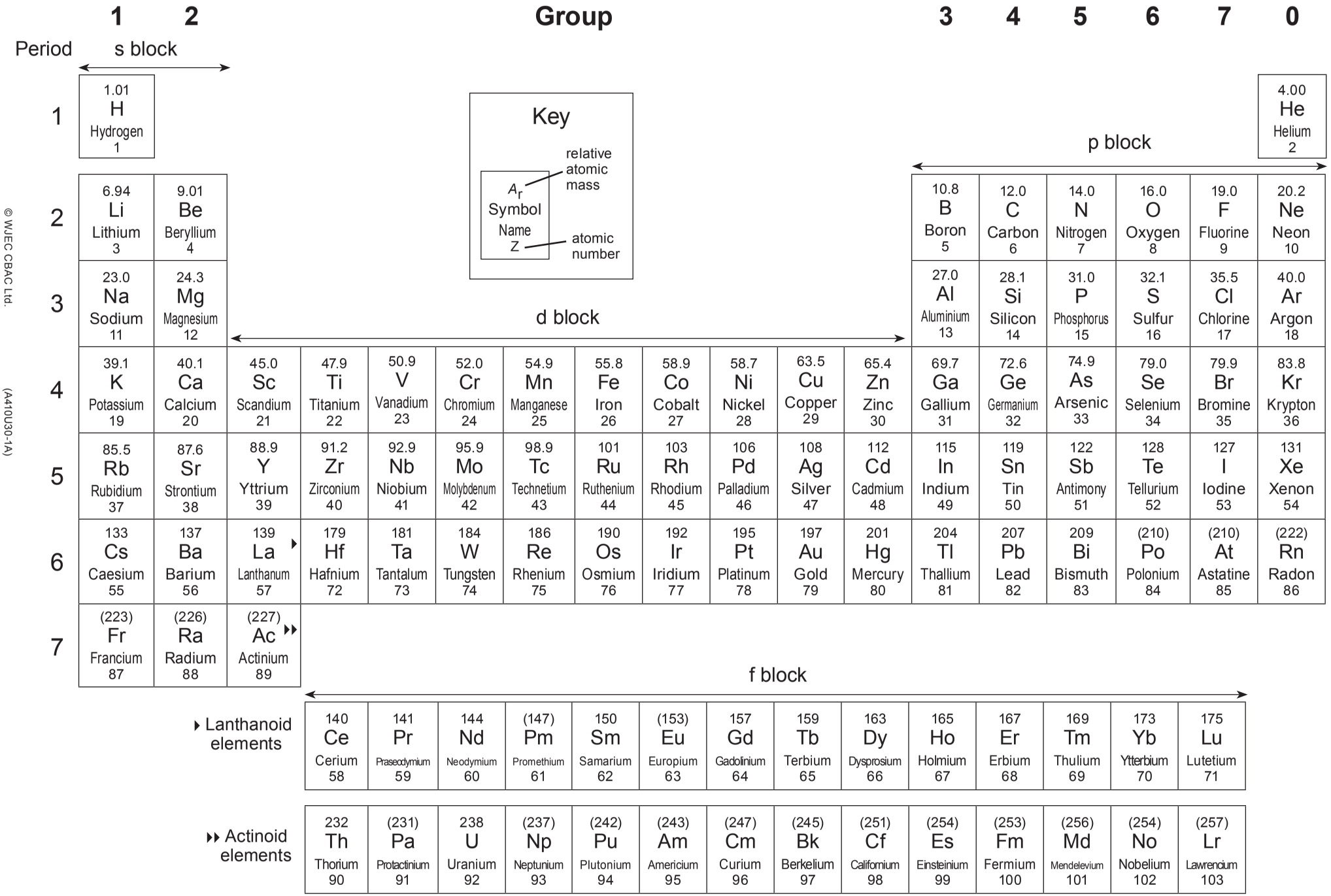
Type of proton	Chemical shift, δ (ppm)
$-CH_3$	0.1 to 2.0
$R-CH_3$	0.9
$R-CH_2-R$	1.3
$CH_3-C\equiv N$	2.0
	2.0 to 2.5
	2.0 to 3.0
	2.2 to 2.3
HC—Cl or HC—Br	3.1 to 4.3
HC—O	3.3 to 4.3
R—OH	4.5 *
—C=CH	4.5 to 6.3
—C=CH—CO	5.8 to 6.5
	6.5 to 7.5
	6.5 to 8.0
	7.0 *
	9.8 *
	11.0 *

*variable figure dependent on concentration and solvent

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THE PERIODIC TABLE



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