OCR

Tuesday 17 June 2014 – Afternoon

A2 GCE CHEMISTRY A

F325/01 Equilibria, Energetics and Elements

Candidates answer on the Question Paper.

OCR supplied materials:
- *Data Sheet for Chemistry A* (inserted)

Other materials required:
- Scientific calculator

Duration: 2 hours

Candidate forename	Max	Candidate surname	

Centre number						Candidate number				

INSTRUCTIONS TO CANDIDATES

- The Insert will be found inside this document.
- Write your name, centre number and candidate number in the boxes above. Please write clearly and in capital letters.
- Use black ink. HB pencil may be used for graphs and diagrams only.
- Answer **all** the questions.
- Read each question carefully. Make sure you know what you have to do before starting your answer.
- Write your answer to each question in the space provided. If additional space is required, you should use the lined pages at the end of this booklet. The question number(s) must be clearly shown.
- Do **not** write in the bar codes.

INFORMATION FOR CANDIDATES

- The number of marks is given in brackets [ ] at the end of each question or part question.
- ✎ Where you see this icon you will be awarded marks for the quality of written communication in your answer.

  This means, for example, you should:
  - ensure that text is legible and that spelling, punctuation and grammar are accurate so that meaning is clear;
  - organise information clearly and coherently, using specialist vocabulary when appropriate.
- You may use a scientific calculator.
- A copy of the *Data Sheet for Chemistry A* is provided as an insert with this question paper.
- You are advised to show all the steps in any calculations.
- The total number of marks for this paper is **100**.
- This document consists of **24** pages. Any blank pages are indicated.

Turn over

Answer **all** the questions.

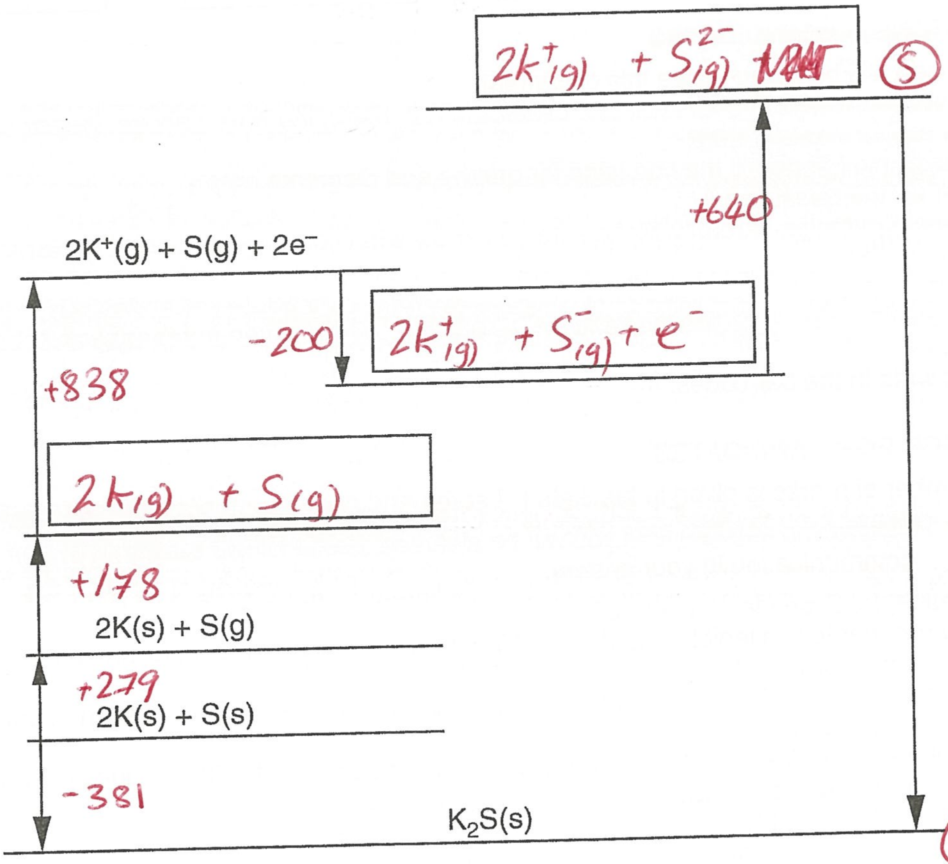
1 Born–Haber cycles can be used to calculate enthalpy changes indirectly.

(a) The table below shows enthalpy changes for a Born–Haber cycle involving potassium sulfide, K_2S.

	Enthalpy change / $kJ\,mol^{-1}$	
Formation of potassium sulfide, K_2S	–381	
1st electron affinity of sulfur	–200	
2nd electron affinity of sulfur	+640	
Atomisation of sulfur	+279	
1st ionisation energy of potassium	+419	×2= 838
Atomisation of potassium	+89	×2= 178

(i) The incomplete Born–Haber cycle below can be used to determine the lattice enthalpy of potassium sulfide.

In the boxes, write the species present at each stage in the cycle.
Include state symbols for the species.



[3]

(ii) Define, in words, the term *lattice enthalpy*.

FORMATION OF 1 MOLE OF COMPOUND ✓

FROM ITS GASEOUS IONS UNDER S. COND^{n}'S ✓

[2]

(iii) Using the Born–Haber cycle, calculate the lattice enthalpy of potassium sulfide.

$-640 + 200 - 838 - 178 - 279 - 381$ ✓

lattice enthalpy = −2116 kJ mol^{-1} **[2]**

(b) Several ionic radii are shown below.

Ion	Na^{+}	K^{+}	Rb^{+}	Cl^{-}	Br^{-}	I^{-}
Radius/pm	95	133	148	181	195	216
	1	2	3	1	2	3

Predict the order of melting points for NaBr, KI and RbCl from lowest to highest.

STRONGEST ATTRACTION (C. DENSITY)

WEAKEST ATTRACTION (C. DENSITY)

Explain your answer.

NaBr 1 2 ③ STRONGEST

KI 2 3 ⑤ WEAKEST

RbCl 3 1 ④

Lowest melting point KI

RbCl

Highest melting point NaBr ✓

Explanation

- SMALLER ION AND HIGHER CHARGE GIVE A HIGHER CHARGE DENSITY. GREATER ATTRACTION BETWEEN IONS OF HIGHER CHARGE DENSITY ✓
- ∴ MORE ENERGY REQUIRED TO OVERCOME ATTRACTION.

[3]

[Total: 10]

2 This question looks at different aspects of entropy.

(a) Three processes are given below.

For each process, state and explain whether the change would be accompanied by an increase or decrease in entropy.

(i) The freezing of water.

increase or decrease DECREASE

explanation ICE HAS LESS DISORDER THAN WATER

[1]

(ii) The reaction of calcium carbonate with hydrochloric acid.

increase or decrease INCREASE

explanation CO_2 GAS IS FORMED WHICH HAS GREATER DISORDER

[1]

(iii) The formation of $O_3(g)$ from $O_2(g)$. $3O_2 \rightarrow 2O_3$

increase or decrease DECREASE

explanation FEWER MOLES OF GAS MADE $3 \rightarrow 2$ WHICH IS LESS DISORDER

[1]

(b) The enthalpy and entropy changes of a reaction both have a negative sign.

Discuss how the feasibility of this reaction will change as the temperature increases.

$\Delta G = \Delta H - T\Delta S$

(−)VE TO BE SPONTANEOUS (−)VE (−)VE

AS T~~Δ~~ INCREASES, TΔS BECOMES MORE (−)VE ✓

TΔS WILL BECOME MORE (−)VE THAN ΔH

MEANING ΔG BECOMES LESS NEGATIVE **[2]**

OR MORE (+)VE ∴ LESS FEASIBLE ✓

(c) The metal tungsten is obtained on a large scale from its main ore, wolframite. Wolframite contains tungsten(VI) oxide, WO_3.

Tungsten is extracted from wolframite by reduction with hydrogen:

$$WO_3(s) + 3H_2(g) \rightarrow W(s) + 3H_2O(g) \qquad \Delta H = +115\,\text{kJ mol}^{-1}$$

Standard entropies are given in the table below.

3x 3x

Substance	WO_3(s)	H_2(g)	W(s)	H_2O(g)
$S^{\ominus}$/J K^{-1} mol^{-1}	76	131	33	189

393 567

÷ΔS BY 1000

298k

(i) Calculate the free energy change, ΔG, in kJ mol^{-1}, for this reaction at 25 °C.

Show your working.

$\Delta S = S_p - S_r \Rightarrow (567 + 33) - (76 + 393)$

$= 600 - 469 = +131$ JK^{-1}Mol^{-1} ✓

$(= 0.131)$

$\Delta G = \Delta H - T\Delta S$

$= 115 - (298 \times 0.131)$

$= +75.96$

✓

ΔG at 25 °C = ...+75.96.......... kJ mol^{-1} **[2]**

(ii) Calculate the minimum temperature, in K, at which this reaction becomes feasible.

Show your working.

$\Delta G = 0 \therefore T = \dfrac{\Delta H}{\Delta S}$ ✓

$= \dfrac{115}{0.131} = 878$K ✓

minimum temperature = ... K **[2]**

[Total: 9]

Turn over

3 Ethyne gas, C_2H_2, is manufactured in large quantities for a variety of uses.

Much of this ethyne is manufactured from methane as shown in the equation below.

$$2CH_4(g) \rightleftharpoons C_2H_2(g) + 3H_2(g) \qquad \Delta H = +377\,kJ\,mol^{-1}$$ ENDOTHERMIC

(a) Write an expression for K_c for this equilibrium.

$$Kc = \frac{[C_2H_2][H_2]^3}{[CH_4]^2}$$

[1]

(b) A research chemist investigates how to improve the synthesis of ethyne from methane at a high temperature.

- The chemist adds CH_4 to a 4.00 dm³ container. — TOTAL VOLUME
- The chemist heats the container and allows equilibrium to be reached at constant temperature. The total gas volume does not change.
- The equilibrium mixture contains 9.36×10^{-2} mol CH_4 and 0.168 mol C_2H_2.

MOLES EQ MOLES EQ

(i) Calculate the amount, in mol, of H_2 in the equilibrium mixture.

	$2CH_4$	$\rightleftharpoons$ C_2H_2	+ $3H_2$
START		0	0
REACTED			
EQM	9.36×10^{-2}	$0.168 \xrightarrow{\times 3}$	0.504

amount of H_2 = 0.504 mol **[1]**

(ii) Calculate the equilibrium constant, K_c, at this temperature, including units.

Give your answer to **three** significant figures.

$$[CH_4] = \frac{9.36\times10^{-2}}{4} = 2.34\times10^{-2} \qquad [C_2H_2] = \frac{0.168}{4} = 4.2\times10^{-2} \qquad [H_2] = \frac{0.504}{4} = 0.126 ✓$$

$$Kc = \frac{4.2\times10^{-2} \times (0.126)^3}{(2.34\times10^{-2})^2}$$

$$= 0.153$$

K_c = 0.153 ✓ units $mol^2\,dm^{-6}$ **[3]**

3SF

(iii) Calculate the amount, in mol, of CH_4 that the chemist originally added to the container.

$$2CH_4 \rightleftharpoons C_2H_2 + 3H_2$$

START	0.4296	0	0
REACTED	0.336 ← ×2		
EQM	9.36×10^{-2}	0.168	0.504

START = REACTED + AT EQM

$= 0.336 + 9.36 \times 10^{-2}$

$= 0.4296$

amount of CH_4 = 0.4296 mol **[1]**

(c) The chemist repeats the experiment three times.
In each experiment the chemist makes **one** change but uses the **same** initial amount of CH_4.

Complete the table to show the predicted effect of each change compared with the original experiment.

Only use the words **greater**, **smaller** or **same**.

Change	K_c	**Equilibrium amount of $C_2H_2(g)$/mol**	**Initial rate**
The container is heated at constant pressure	GREATER (EQM → RHS)	GREATER	GREATER
A smaller container is used (↑P) ↑	SAME (ONLY T CHANGES Kc)	SMALLER (EQM FEWER MOLES OF GAS)	GREATER
A catalyst is added to CH_4 at the start	SAME (ONLY T CHANGES Kc)	SAME	GREATER

[3]

(d) In this manufacture of ethyne, hydrogen is also produced. To improve the atom economy of the process, it is important to make use of the hydrogen. For example, hydrogen can be used in the extraction of some metals from their ores.

State **two** other large-scale uses of the hydrogen.

1 FUEL CELLS

2 MARGARINE

[1]

[Total: 10]

4 A student carries out an initial rates investigation on the reaction below.

$$5I^-(aq) + IO_3^-(aq) + 6H^+(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$$

From the results, the student determines the rate equation for this reaction:

$$\text{rate} = k\,[I^-(aq)]^2\,[IO_3^-(aq)]\,[H^+(aq)]^2$$

(a) **(i)** What is the overall order of reaction?

5 **[1]**

(ii) A proposed mechanism for this reaction takes place in several steps.

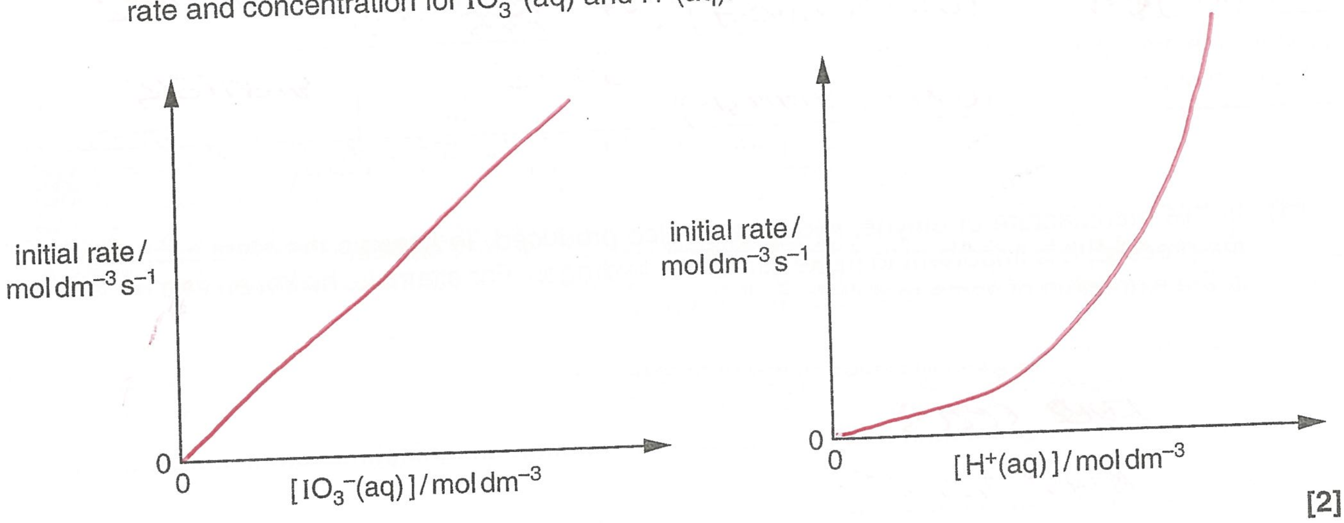
Suggest **two** reasons why it is unlikely that this reaction could take place in one step.

STOICHIOMETRY OF RATE EQ^{N} DOES NOT MATCH THE OVERALL EQ^{N} ✓

COLLISION INVOLVING 5 SPECIES UNLIKELY ✓

[2]

(b) On the rate–concentration graphs below, sketch lines to show the relationship between initial rate and concentration for $IO_3^-(aq)$ and $H^+(aq)$.



[2]

(c) The table below shows some of the student's results.

(i) Complete the table by adding the missing initial rates in the boxes.

	$[I^-(aq)]$ / mol dm^{-3}	$[IO_3^-(aq)]$ / mol dm^{-3}	$[H^+(aq)]$ / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
Experiment 1	0.015	0.010	0.020	0.60
Experiment 2	0.045	0.010	0.020	5.40
Experiment 3	0.060	0.040	0.080	614.4

$\times 3^2$ $\times 4^2$ $\times 4$ $\times 4^2$ $\times 3^2$

[2]

$= \times 4^2 \times 4 \times 4^2$

$= \times 1024$

(ii) Calculate the rate constant, k, for this reaction. Include units.

Give your answer to **two** significant figures.

$$k = \frac{r}{[I^-]^2 \times [IO_3^-] \times [H^+]^2} \quad \frac{0.6}{(0.0.15)^2 \times 0.01 \times (0.02)^2}$$

k = 6.7×10^8 units mol^{-4} dm^{12} s^{-1} **[3]**

2SF

(iii) The student repeats Experiment 1 using 0.020 mol dm^{-3} methanoic acid, HCOOH(aq) (pK_a = 3.75), instead of 0.020 mol dm^{-3} HCl(aq) as a source of H^+(aq).

Determine the initial rate in this experiment. Show your working.

$K_a = 10^{-3.75} = 1.78 \times 10^{-4}$ ✓

$[H^+] = \sqrt{1.78 \times 10^{-4} \times 0.02}$

$= 1.89 \times 10^{-3}$ ✓

$r = 6.7 \times 10^8 \times \cancel{0.15}\,(0.015)^2 \times 0.01 \times (1.89 \times 10^{-3})^2$

✓

initial rate = 5.33×10^{-3} mol dm^{-3} s^{-1} **[3]**

[Total: 13]

5 Elements in the d-block of the Periodic Table form ions that combine with ligands to form complex ions. Most d-block elements are also classified as transition elements.

(a) Explain why two of the Period 4 d-block elements (Sc–Zn) are **not** also transition elements.

In your answer you should link full electron configurations to your explanations.

MUST WRITE FULL

$Sc: [Ar] 4S^2 3d^1 \Rightarrow Sc^{3+} : [Ar]$

$Zn: [Ar] 4S^2 3d^{10} \Rightarrow Zn^{2+} : [Ar] 3d^{10}$

EMPTY 3d ✓ $\rightarrow Sc^{3+} 1S^2 2S^2 2P^6 3S^2 3P^6$ ✓

FULL 3d ✓ $\rightarrow Zn^{2+} 1S^2 2S^2 2P^6 3S^2 3P^6 3d^{10}$ ✓

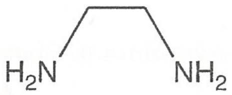
Sc + Zn NOT TM ✓

TM MUST FORM AN ION WITH AN INCOMPLETE d SUB-SHELL ✓

[6]

(b) The cobalt(III) ion, Co^{3+}, forms a complex ion **A** with two chloride ligands and two ethanediamine, $H_2NCH_2CH_2NH_2$, ligands.

The structure of ethanediamine is shown below.



(i) Explain how ethanediamine is able to act as a bidentate ligand. 2 ✓

HAS 2 e PAIRS ON N TO FORM D.C. BOND TO T.M ✓

~~HAS 2~~

[2]

(ii) Write the formula of complex ion **A**.

$[Co(H_2NCH_2CH_2NH_2)_2 Cl_2]^+$

[1]

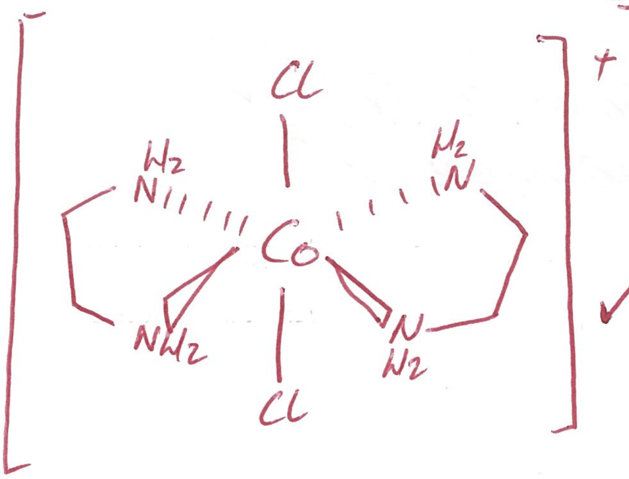
(iii) What is the coordination number of cobalt in complex ion **A**?

6

[1]

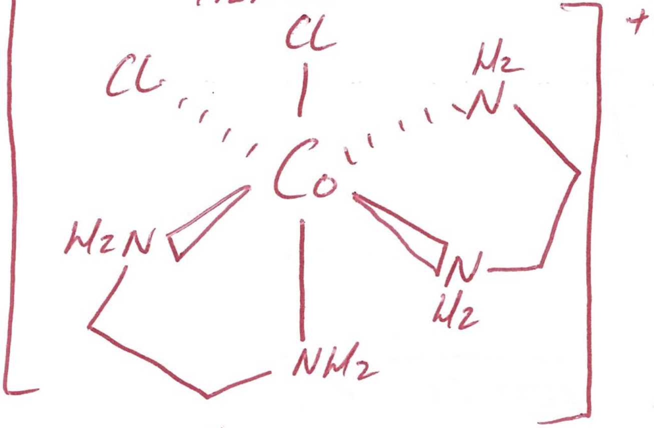
(iv) Complex ion **A** has *cis* and *trans* stereoisomers. One of these stereoisomers also has an optical isomer.

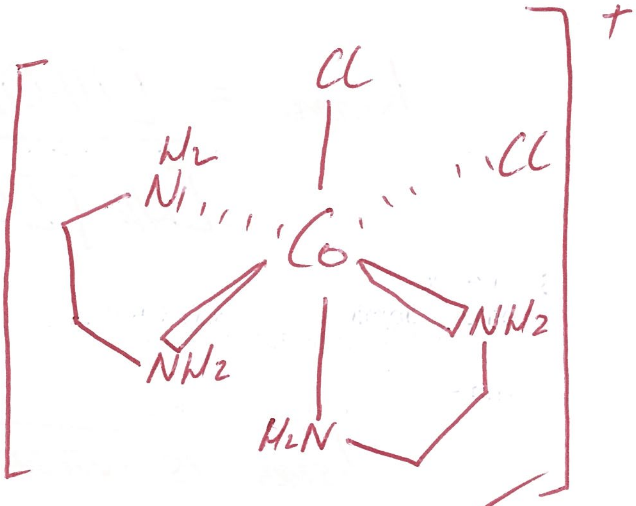
Draw 3-D diagrams to show the three stereoisomers.



✓

TRANS





✓ ✓

CIS CIS

[3]

OPTICAL ISOMERS.

Question 5 continues on page 12

(c) The equilibrium reaction for the transport of oxygen by haemoglobin (Hb) in blood can be represented as **equation 5.1**.

$$Hb(aq) + O_2(aq) \rightleftharpoons HbO_2(aq)$$ **equation 5.1**

(i) Explain how ligand substitution reactions allow haemoglobin to transport oxygen in blood.

O_2 FORMS D.C. BOND TO Fe^{2+} ✓

O_2 RELEASED WHEN REQUIRED ✓

.... **[2]**

(ii) Write an expression for the stability constant, K_{stab}, for the equilibrium involved in the transport of oxygen by haemoglobin.

Use the simplified species in **equation 5.1**.

$$K_{STAB} = \frac{[HbO_2]}{[Hb][O_2]}$$

[1]

(iii) In the presence of carbon monoxide, less oxygen is transported in the blood.

Suggest why, in terms of bond strength and stability constants.

K_{STAB} IS GREATER FOR CO ✓

AS IT FORMS STRONGER D.C. BOND ✓

.... **[2]**

[Total: 18]

6 Ethanoic acid, CH_3COOH, is a weak Brønsted–Lowry acid.

(a) An acid–base equilibrium is set up when ethanoic acid is added to water.

Write the equation for the equilibrium that would be set up and label the two conjugate acid–base pairs.

CH_3COOH + H_2O $\rightleftharpoons$ CH_3COO^- + H_3O^+

ACID 1 BASE 2 BASE 1 ACID 2

[2]

(b) An aqueous solution of CH_3COOH has a pH of 3.060.
This solution contains both hydrogen ions and hydroxide ions.

(i) How can an aqueous solution of an acid contain hydroxide ions?

WATER DISSOCIATES

[1]

(ii) Calculate the concentration of hydroxide ions in this solution of ethanoic acid.

$[H^+] = 10^{-3.06}$

$= 8.71 \times 10^{-4}$

$[OH^-] = \frac{1\times10^{-14}}{8.71\times10^{-4}}$

concentration of hydroxide ions = 1.15×10^{-11} mol dm^{-3} [2]

(c) A student adds an excess of aqueous ethanoic acid to solid calcium carbonate.
The resulting solution is able to act as a buffer solution.

(i) Write a full equation for the reaction between ethanoic acid and solid calcium carbonate.

$2CH_3COOH + CaCO_3 \rightarrow (CH_3COO)_2Ca + CO_2 + H_2O$ **[1]**

(ii) Explain why this buffer solution has formed.

CONTAINS CH_3COOH AND CH_3COO^-

[1]

(iii) Explain how this buffer solution controls pH when either an acid or an alkali is added.

In your answer you should explain how the equilibrium system allows the buffer solution to control the pH.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ ✓

ADD H^+ EQ^m SHIFTS TO LHS ✓
BY H^+ REACTING WITH CH_3COO^- ✓

ADD OH^- EQ^m SHIFTS TO RHS ✓
AS $OH^- + H^+ \rightarrow H_2O$ ✓

[5]

DO NOT WRITE IN THIS MARGIN

(d) A biochemist plans to make up a buffer solution with a pH of 5.000.
The biochemist adds solid sodium ethanoate, CH_3COONa, to 400 cm³ of 0.200 mol dm^{-3} ethanoic acid.
K_a for ethanoic acid = 1.75×10^{-5} mol dm^{-3}

Calculate the mass of sodium ethanoate that the biochemist needs to dissolve in the ethanoic acid to prepare this buffer solution.

Assume that the volume of the solution remains constant at 400 cm³ on dissolving the sodium ethanoate.

$[H^+] = 10^{-5} = 1\times10^{-5}$ ✓

$$Ka = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$[CH_3COO^-] = \frac{Ka \times [CH_3COOH]}{[H^+]}$$

$$= \frac{1.75\times10^{-5} \times 0.2}{1\times10^{-5}}$$ ✓

$= 0.35$ Moldm^{-3} ✓

Moles = C × V ✓

$= 0.35 \times 0.4$

$= 0.14$ ✓

MASS = MOLES × Mr

$= 0.14 \times 82.0$

$= 11.5$g ✓

[5]

[Total: 17]

7 Electrochemical cells contain two redox systems, one providing electrons and the other accepting electrons. The tendency to lose or gain electrons is measured using values called standard electrode potentials.

(a) Define the term *standard electrode potential*.

Include all standard conditions in your answer.

EMF OF A ½ CELL WHEN CONNECTED TO

STANDARD H_2 ½ CELL ✓

~~HB~~ 1 ATM 298K 1mol dm^{-3} ✓

[2]

(b) The table below shows two redox systems and their standard electrode potentials, $E^\ominus$.

←

Redox system	$E^\ominus$/V
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

→

MOST (-)VE

MOST (+)VE

A standard $Cu^{2+}(aq)/Cu(s)$ half-cell is connected to a standard $Ag^+(aq)/Ag(s)$ half-cell. The potential of the cell is measured.

Water is then added to the $Cu^{2+}(aq)/Cu(s)$ half-cell. This changes the position of equilibrium in the half-cell. The cell potential increases.

(i) Write down the equation for the overall cell reaction.

$2Ag^+ + Cu \rightarrow 2Ag + Cu^{2+}$ **[1]**

(ii) Explain, in terms of equilibrium, why the cell potential increases.

$Cu^{2+}_{(aq)} + 2e^- \rightleftharpoons Cu_{(s)}$

ADDING H_2O REDUCES THE $[Cu^{2+}]$ ∴

EQ^{M} SHIFTS TO LHS ✓

THIS RELEASES MORE e's ✓

MAKING THIS ½ CELL MORE (-)VE

∴ BIGGER DIFFERENCE ✓

[3]

(c) Hydrogen fuel cells are being developed for powering vehicles.

(i) State **one** advantage of using hydrogen as a fuel compared with conventional fuels.

LESS CO_2 MADE / ONLY H_2O MADE

.... [1]

(ii) In vehicles, hydrogen can be stored on the surface of a solid material or within a solid material.

State **one** other way that hydrogen can be stored as a fuel for vehicles.

AS A LIQUID UNDER PRESSURE

.... [1]

(d) Aluminium–oxygen cells are being investigated for powering vehicles.

The **reactions** at each electrode are shown below.

×4 $Al(s) + 4OH^-(aq) \rightarrow Al(OH)_4^-(aq) + 3e^-$ — MOST (-)VE

×3 $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$ — MOST (+)VE +0.40

(i) The standard electrode potential for the O_2/OH^- redox system is +0.40 V.
The standard cell potential of an aluminium–oxygen cell is 2.71 V.

What is the standard electrode potential of the aluminium redox system in this cell?

$E_{cell} = E_P - E_N \quad 2.71 = 0.4 - E$

$E = 0.4 - 2.71$

standard electrode potential = −2.31 V [1]

(ii) Construct the overall cell equation for an aluminium–oxygen cell.

$4Al + 16OH^- + 3O_2 + 6H_2O \rightarrow 4Al(OH)_4^- + \cancel{12OH^-}$

$\frac{-12}{4}$

$4Al + 4OH^- + 3O_2 + 6H_2O \rightarrow 4Al(OH)_4^-$

✓✓

[2]

[Total: 11]

8 A student carries out an investigation to prepare and analyse a sample of barium ferrate(VI), $BaFeO_4$. The steps in the investigation are shown below.

PREPARE

Step 1

The student adds solid iron(III) oxide to a hot aqueous solution containing an excess of hydroxide ions. The student bubbles chlorine gas through the mixture.
A solution forms containing aqueous ferrate(VI) ions, $FeO_4^{2-}(aq)$, and aqueous chloride ions.

Step 2

The student adds aqueous barium chloride to the resulting solution.
A precipitate of impure barium ferrate(VI) forms.
The precipitate is filtered, washed with distilled water and dried.
The student obtains 0.437 g of impure solid barium ferrate(VI).

ANALYSE

Step 3

An excess of acidified aqueous potassium iodide is added to the solid from **step 2**.
The $BaFeO_4$ reacts as shown below, and the impurity does not react. A solution forms containing aqueous iodine, $I_2(aq)$.

OX ∴ RED. AGENT

$$\underset{+6}{BaFeO_4(s)} + 8H^+(aq) + \underset{-1}{3I^-(aq)} \longrightarrow \underset{+3}{Fe^{3+}(aq)} + Ba^{2+}(aq) + \underset{0}{1\tfrac{1}{2}I_2(aq)} + 4H_2O(l)$$

RED

Step 4

The student determines the amount of I_2 formed by carrying out a titration with aqueous sodium thiosulfate, $Na_2S_2O_3(aq)$.

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

26.4 cm³ of 0.100 mol dm⁻³ $Na_2S_2O_3(aq)$ are required to reach the end point.

(a) Construct an equation for the oxidation of iron(III) oxide (**step 1**).

*

$$\tfrac{1}{2}\underset{+3}{Fe_2O_3} + OH^- + \underset{0}{1\tfrac{1}{2}Cl_2} \longrightarrow \underset{+6}{FeO_4^{2-}} + \underset{-1}{Cl^-}$$

③ ① $\boxed{\times 3}$

LEAVE OUT OH⁻

$$\underset{1\frac{1}{2}}{\tfrac{1}{2}Fe_2O_3} + 5OH^- \quad 1\tfrac{1}{2}Cl_2 \longrightarrow \underset{4}{FeO_4^{2-}} + 3Cl^- + 2\tfrac{1}{2}H_2O$$

0

[2]

(b) Write an **ionic** equation for the formation of barium ferrate(VI) (**step 2**).

Include state symbols.

$$Ba^{2+}_{(aq)} + FeO^{2-}_{4(aq)} \longrightarrow BaFeO_{4(s)}$$ [1]

SEE BACK PAGE *

(c) In **step 3**, what is the reducing agent?

Explain your answer in terms of electrons.

reducing agent I^-

explanation AS IT LOSES e's

REDUCING $FeO_4^{2-} \rightarrow Fe^{3+}$

[2]

(d) The solid sample of barium ferrate(VI) obtained in step 2 is impure.

Determine the percentage, by mass, of barium ferrate(VI) in the 0.437 g of solid formed in **step 2**.

Give your answer to **one** decimal place.

MOLES ~~2s~~ $S_2O_3^{2-} = 0.1 \times 0.0264$

$= 2.64 \times 10^{-3}$ ✓

MOLES $I_2 = 1.32 \times 10^{-3}$ (2:1)

MOLES $BaFeO_4 = 8.8 \times 10^{-4}$ (1½ : 1) ✓

MASS $= 8.8 \times 10^{-4} \times 257.1$

$= 0.226g$ ✓

$\% = \frac{0.226}{0.437} \times 100$

$= 51.77$

percentage of barium ferrate(VI) = 51.8 ✓ 1dp * % **[4]**

Question 8 continues on page 20

(e) When the solution is not alkaline, ferrate(VI) ions react with water.
The reaction forms a gas with a density of $1.333 \times 10^{-3}\,g\,cm^{-3}$, measured at room temperature and pressure, and an orange–brown precipitate.

Fe^{3+} WITH OH^-?   MASS   1cm3 VOL OF GAS

a)• Determine the formulae of the gas and the precipitate.
b)• Write an equation for the reaction that takes place.

a) MOLES $= \frac{1}{24000} = 4.167 \times 10^{-4}$ $Mr = \frac{1.333 \times 10^{-3}}{4.167 \times 10^{-5}}$

$= 32 \therefore O_2$ ✓

$Fe(OH)_3$ PRECIPITATE ✓

~~B~~)

gas O_2

precipitate $Fe(OH)_3$

b) equation $\underset{+6}{FeO_4} + \underset{-2}{H_2O} \rightarrow \underset{+3}{Fe(OH)_3} + \frac{1}{2}\underset{0}{O_2}$

③ ②

$2FeO_4^{2-} + 3H_2O \rightarrow 2Fe(OH)_3 + 1\frac{1}{2}O_2$

[3]

OXYGEN ⑪ ⑨

[Total: 12]

$\therefore + 2OH^-$

END OF QUESTION PAPER

HYDROGEN ⑥ ⑧

$\therefore + 2H_2O$

$2FeO_4^{2-} + 5H_2O \rightarrow 2Fe(OH)_3 + 1\frac{1}{2}O_2 + 2OH^-$

OXYGEN ⑬ ⑪ $\therefore + 2OH^-$

HYDROGEN ⑩ ⑩ ✓

$2FeO_4^{2-} + 5H_2O \rightarrow 2Fe(OH)_3 + 1\frac{1}{2}O_2 + 4OH^-$

Qu <u>8a</u> * LEAVE $OH^-/H_2O + H^+$ OUT

½ FOR OX NR'S

$$\frac{1}{2}Fe_2O_3 + \frac{1}{2}Cl_2 \longrightarrow FeO_4^{2-} + Cl^-$$

+3 0 +6 −1

+3 → +6 ③

0 → −1 ① ∴ ×3

$$\Rightarrow \frac{1}{2}Fe_2O_3 + 1\frac{1}{2}Cl_2 \longrightarrow FeO_4^{2-} + 3Cl^-$$

DOUBLE EQ^N AS ½ OXYGENS WOULD BE IN BALANCING.

$$Fe_2O_3 + 3Cl_2 \longrightarrow 2FeO_4^{2-} + 6Cl^-$$

OXYGEN ③ ⑧

ADD 5× OH^- ON LHS

$$Fe_2O_3 + 5OH^- + 3Cl_2 \rightarrow 2FeO_4^{2-} + 6Cl^-$$

OXYGEN ⑧ ⑧

HYDROGEN ⑤ ⓪ ∴ ADD 5× H_2O

$$Fe_2O_3 + 5OH^- + 3Cl_2 \longrightarrow 2FeO_4^{2-} + 6Cl^- + 5H_2O$$

OXYGEN ⑧ ⑬

∴ ADD 5× OH^-

$$Fe_2O_3 \ ~~KOH~~ + 10OH^- + 3Cl_2 \longrightarrow 2FeO_4^{2-} + 6Cl^- + 5H_2C$$

OXYGEN ⑬ ⑬

HYDROGEN ⑩ ⑩ ✓

BALANCED.