

OXFORD CAMBRIDGE AND RSA EXAMINATIONS
A2 GCE
F325/01

CHEMISTRY A
Equilibria, Energetics and Elements

WEDNESDAY 12 JUNE 2013: Afternoon

DURATION: 2 hours
plus your additional time allowance

MODIFIED ENLARGED

Candidate forename		Candidate surname	
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Centre number						Candidate number				
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Candidates answer on the Question Paper.

OCR SUPPLIED MATERIALS:

Data Sheet for Chemistry A (inserted)

OTHER MATERIALS REQUIRED:


Scientific calculator

READ INSTRUCTIONS OVERLEAF

INSTRUCTIONS TO CANDIDATES

- **The Insert will be found in the centre of this document.**
- **Write your name, centre number and candidate number in the boxes on the first page. 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.**

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.
- Any blank pages are indicated.

Answer ALL the questions.

- 1 Born–Haber cycles can be used to determine lattice enthalpies of ionic compounds.**

(a) Define, in words, the term lattice enthalpy.

[2]

(b) The Born–Haber cycle opposite can be used to determine the lattice enthalpy of calcium oxide. The cycle includes the values for the enthalpy changes of the steps labelled A–G.

(i) Complete the Born–Haber cycle by adding the species present in the two boxes.

Include state symbols.

[2]

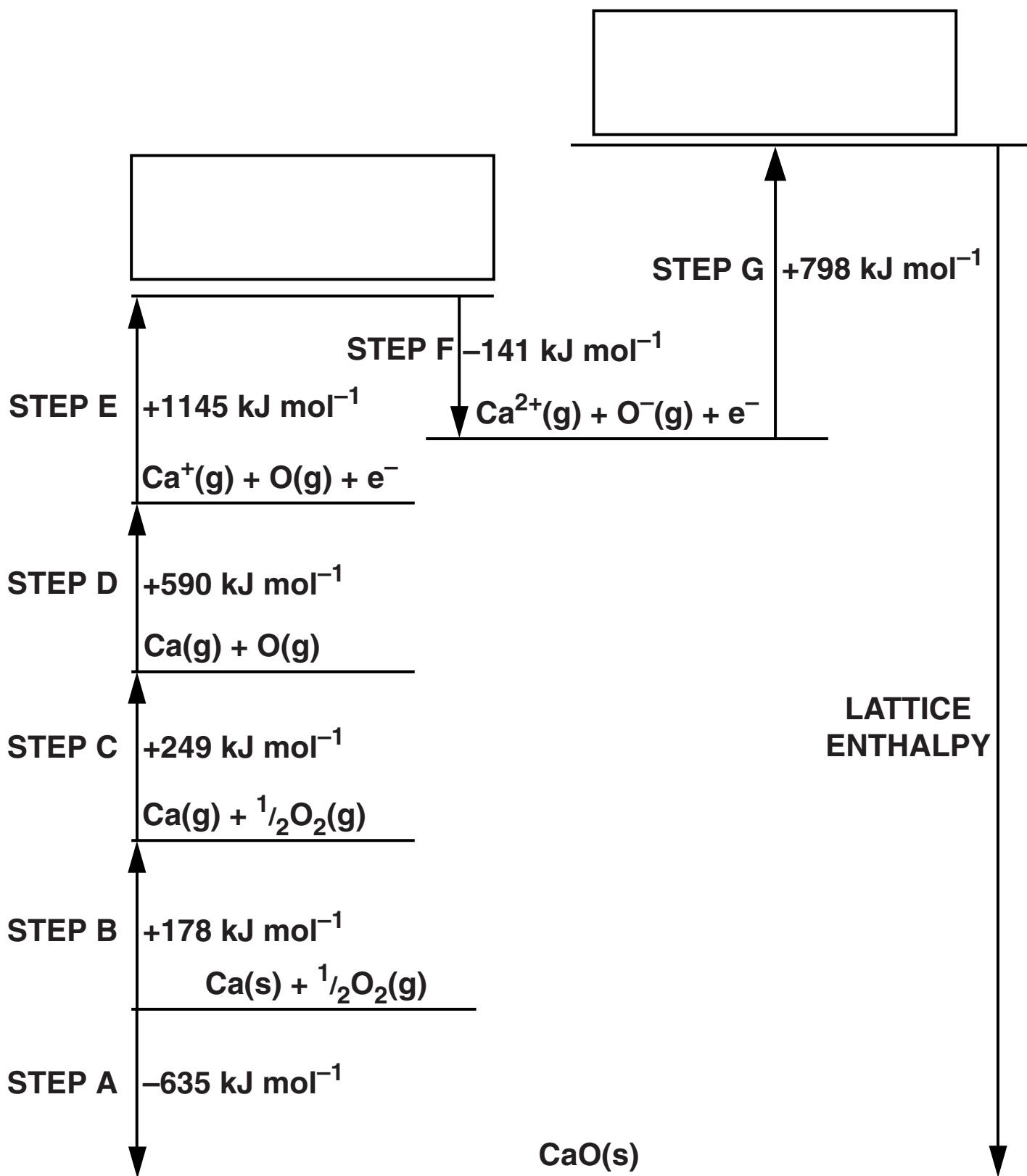
(ii) Name the enthalpy changes for the following steps in the Born–Haber cycle.

• STEP A

• STEP C

• STEP G

[3]



(iii) Calculate the lattice enthalpy of calcium oxide.

answer = _____ kJ mol^{-1} [2]

(c) Describe and explain the factors that affect the values of lattice enthalpies.

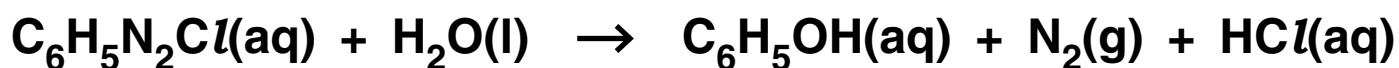
[3]

[TOTAL: 12]

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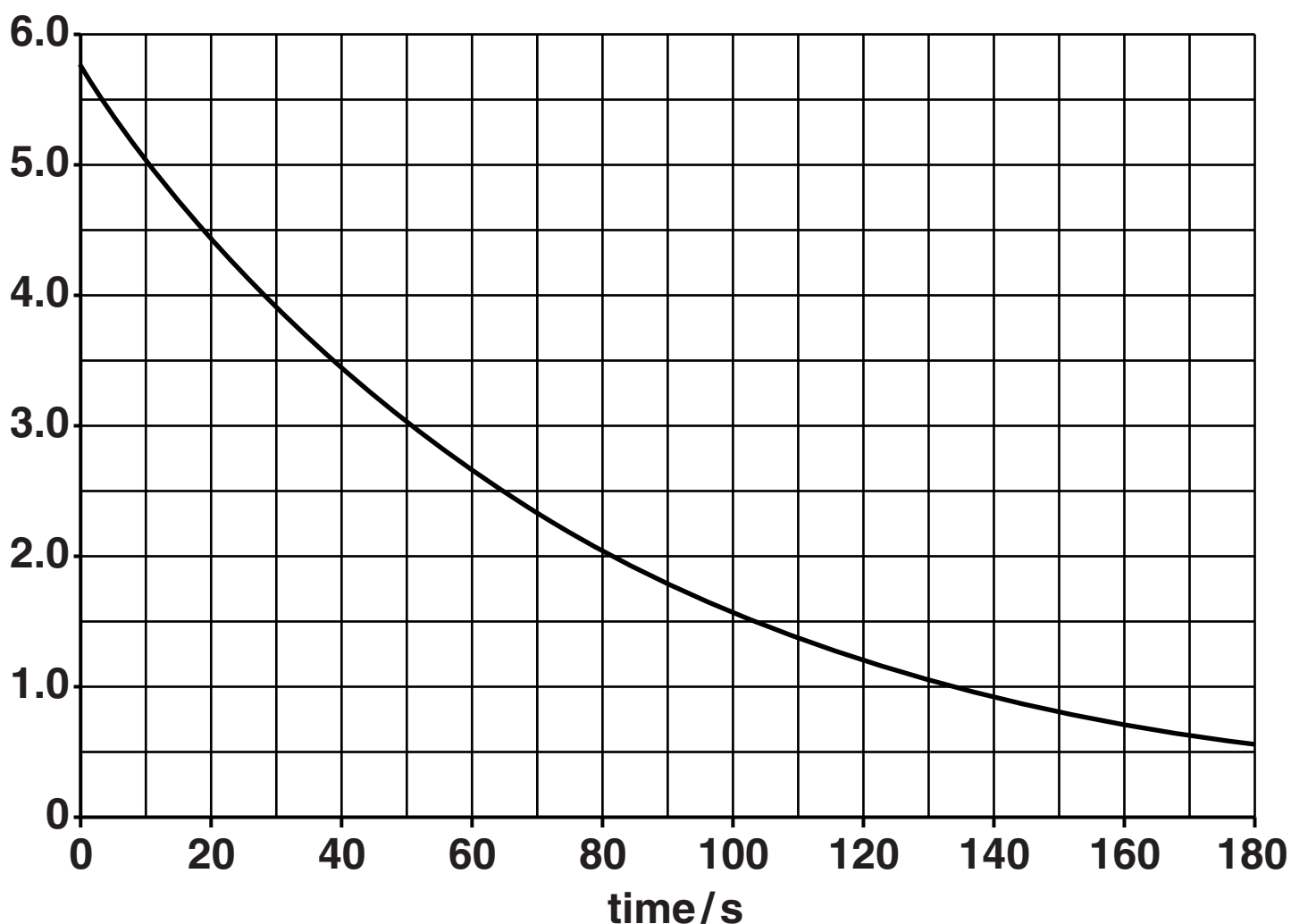
QUESTION 2 BEGINS ON PAGE 8

2 In aqueous solution, benzenediazonium chloride, $\text{C}_6\text{H}_5\text{N}_2\text{Cl}$, decomposes above 10°C .



A student investigates the rate of this reaction using an excess of water at 50°C . The student takes measurements at intervals during the reaction and then plots his experimental results to give the graph shown below.

$[\text{C}_6\text{H}_5\text{N}_2\text{Cl}(\text{aq})]$
 $/\text{mol dm}^{-3}$



(a) The student uses half-life to suggest the order of reaction with respect to $\text{C}_6\text{H}_5\text{N}_2\text{Cl}$.

(i) What is meant by the half-life of a reaction?

[1]

(ii) Confirm the order of reaction with respect to $\text{C}_6\text{H}_5\text{N}_2\text{Cl}$.

Show your working on the graph.

[2]

(iii) What would be the effect, if any, on the half-life of this reaction of doubling the initial concentration of $\text{C}_6\text{H}_5\text{N}_2\text{Cl}$?

[1]

**(b) The student predicts that the rate equation is:
rate = $k[\text{C}_6\text{H}_5\text{N}_2\text{Cl}]$.**

**(i) Using the graph and this rate equation,
determine the rate of reaction after 40 s.**

Show your working on the graph.

rate = _____ units _____ [3]

**(ii) Calculate the rate constant, k , for this reaction
and give its units.**

k = _____ units _____ [2]

(c) The order of this reaction with respect to H_2O is effectively zero.

Explain why.

[1]

[TOTAL: 10]

- 3 Hydrogen and iodine react together in a reversible reaction:**



A chemist mixes together $2.00 \times 10^{-3} \text{ mol H}_2(\text{g})$ and $4.00 \times 10^{-3} \text{ mol I}_2(\text{g})$ in a 1.00 dm^3 container.

The chemist seals the container.

The mixture is heated and left to reach equilibrium.

At equilibrium, the mixture contains $3.00 \times 10^{-4} \text{ mol}$ of H_2 .

- (a) Calculate the equilibrium constant, K_c , including units, if any, for this equilibrium.**

Give your answer to THREE significant figures.

$K_c =$ _____ units _____ [5]

(b) The chemist repeats the experiment several times. In each experiment, the chemist makes one change.

(i) The chemist uses $3.00 \times 10^{-3} \text{ mol H}_2(\text{g})$ instead of $2.00 \times 10^{-3} \text{ mol H}_2(\text{g})$.

Predict whether the amounts of $\text{H}_2(\text{g})$, $\text{I}_2(\text{g})$ and $\text{HI}(\text{g})$ in the equilibrium mixture would be greater, smaller or the same as in the original experiment.

Answer by placing ticks in the appropriate boxes of the table below.

	$\text{H}_2(\text{g})$	$\text{I}_2(\text{g})$	$\text{HI}(\text{g})$
Greater			
Smaller			
The same			

[2]

(ii) The chemist heats the mixture to a higher temperature at constant pressure.

Explain whether the value of K_c would be greater, smaller or the same.

[1]

- (iii) The chemist increases the pressure of the mixture at constant temperature.

Explain whether the value of K_c would be greater, smaller or the same.

[1]

[TOTAL: 9]

4 This question looks at pH values and reactions of acids, bases and buffers.

(a) **0.14 mol dm⁻³ solutions of hydrochloric acid, HCl, and chloric(I) acid, HClO ($pK_a = 7.43$), have different pH values.**

Explain why the pH values are different and calculate the pH of 0.14 mol dm⁻³ solutions of HCl and HClO to TWO DECIMAL PLACES.

Show any working in calculations.

[illegible]

- (b) Aluminium powder is added to aqueous ethanoic acid, CH_3COOH .**

Write full and ionic equations for the reaction that takes place.

full equation _____

ionic equation _____ **[2]**

- (c) Calculate the pH of a 0.40 mol dm^{-3} solution of NaOH.**

[2]

(d) In biochemistry, buffer solutions based on methanoic acid can be used in the analysis of urine samples.

(i) Explain what is meant by the term buffer solution.

Describe how a buffer solution based on methanoic acid can act as a buffer.



In your answer you should explain how the equilibrium system allows the buffer solution to control the pH. [7]

[illegible]

(ii) A chemist prepares a buffer solution by mixing together the following:

**200 cm³ of 3.20 mol dm⁻³ HCOOH ($K_a = 1.70 \times 10^{-4}$ mol dm⁻³)
and
800 cm³ of 0.500 mol dm⁻³ NaOH.**

The volume of the buffer solution is 1.00 dm³.

- Explain why a buffer solution is formed when these two solutions are mixed together.**
- Calculate the pH of this buffer solution. [6]**

Give your answer to TWO decimal places.

[6]

[TOTAL: 22]

- 5 Iron is heated with chlorine to form an orange–brown solid, A.**

Solid A is dissolved in water to form an orange–brown solution, X, containing the complex ion $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$.

Separate portions of solution X are reacted as shown in EXPERIMENTS 1–4 below.

EXPERIMENT 1

Aqueous sodium hydroxide is added to solution X. An orange–brown precipitate B forms.

EXPERIMENT 2

Excess zinc powder is added to solution X and the mixture is heated. The excess zinc is removed leaving a pale-green solution containing the complex ion C and aqueous Zn^{2+} ions.

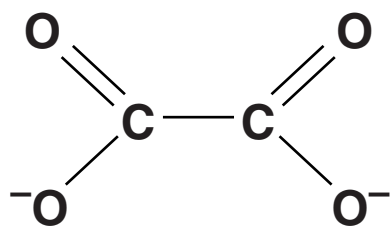
EXPERIMENT 3

An excess of aqueous potassium cyanide, $\text{KCN}(\text{aq})$, is added to solution X. The solution turns a yellow colour and contains the complex ion E. E has a molar mass of 211.8 g mol^{-1} .

EXPERIMENT 4

An aqueous solution containing ethanedioate ions, $(\text{COO}^-)_2$, is added to solution X. A coloured solution forms containing a mixture of optical isomers F and G.

The structure of the ethanedioate ion is shown below.



(a) Write an equation for the formation of solid A.

_____ [1]

(b) In EXPERIMENT 1, write an ionic equation for the formation of precipitate B.

_____ [1]

(c) In EXPERIMENT 2,

(i) write an equation for the formation of complex ion C

_____ [2]

(ii) state the type of reaction taking place.

_____ [1]

(d) In EXPERIMENT 3,

(i) write an equation for the formation of complex ion E

_____ [2]

(ii) state the type of reaction taking place.

_____ [1]

(e) In EXPERIMENT 4, optical isomers F and G are formed.

Show the 3-D shapes of F and G.

In your diagrams, show the ligand atoms that are bonded to the metal ions and any overall charges.

[3]

- (f) In a separate experiment, iron metal is heated with potassium nitrate, KNO_3 , a strong oxidising agent. A reaction takes place and the resulting mixture is poured into water. A dark red solution forms containing ferrate(VI) ions. The ferrate(VI) ion has a 2– charge.

Suggest a possible formula for the ferrate(VI) ion.

_____ [1]

[TOTAL: 12]

6 The equation for the reaction of CO₂ and H₂O to produce glucose, C₆H₁₂O₆, and O₂ is shown below.



Standard entropies are given in the table below.

Substance	CO ₂ (g)	H ₂ O(l)	O ₂ (g)
S [⦿] / J K ⁻¹ mol ⁻¹	214	70	205

24

(a) (i) Calculate the standard entropy of glucose.

$$S^{\ominus} = \text{_____ J K}^{-1} \text{ mol}^{-1} \text{ [2]}$$

(ii) Calculate ΔG , in kJ mol^{-1} , at 25°C .

Show all your working.

25

$\Delta G =$ _____ kJ mol^{-1} [2]

(iii) Explain why this reaction is NOT feasible at ANY temperature.

[1]

- (b) Although the reaction between CO_2 and H_2O to form $\text{C}_6\text{H}_{12}\text{O}_6$ and O_2 appears not to be feasible, plants are able to make the reaction take place spontaneously by photosynthesis.

Each year, $3.4 \times 10^{18} \text{ kJ}$ of solar energy is taken in by all the plants on the Earth to make photosynthesis take place.

Calculate the mass of carbon dioxide that is removed each year from the atmosphere by photosynthesis on Earth.

mass of CO_2 = _____ [2]

[TOTAL: 7]

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QUESTION 7 BEGINS ON PAGE 28

- 7** Standard electrode potentials for seven redox systems are shown in TABLE 7.1 opposite. You may need to use this information in parts (a)–(d) of this question.

- (a)** Define the term standard electrode potential. Include all standard conditions in your answer.

[2]

- (b)** An electrochemical cell can be made based on redox systems 1 and 2.

Write down the standard cell potential of this cell.

standard cell potential = _____ V [1]

Redox system		E^\ominus / V
1	$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Mg}(\text{s})$	-2.37
2	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34
3	$\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Al}(\text{s})$	-1.66
4	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
5	$\text{I}_2(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{I}^-(\text{aq})$	+0.54
6	$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$	+1.36
7	$\text{ClO}^-(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightleftharpoons \frac{1}{2}\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$	+1.63

TABLE 7.1

(c) Using redox systems 3, 4 and 5 ONLY in TABLE 7.1, predict THREE reactions that might be feasible.

(i) Write the overall equation for each predicted reaction.

[3]

(ii) Give TWO reasons why it is uncertain whether reactions predicted from E^\ominus values may actually take place.

[2]

- (d) In aqueous acid, $\text{Cl}^{-}(\text{aq})$ ions react with $\text{ClO}^{-}(\text{aq})$ ions to form chlorine gas, $\text{Cl}_2(\text{g})$.
In aqueous alkali, chlorine gas, $\text{Cl}_2(\text{g})$, reacts to form $\text{Cl}^{-}(\text{aq})$ and $\text{ClO}^{-}(\text{aq})$ ions.

Explain this difference.

Use TABLE 7.1 to help you with your answer.

[4]

(e) In acidic conditions, Sn^{2+} ions react with IO_3^- ions to produce iodine and Sn^{4+} ions.

**(i) What is the oxidising agent in this reaction?
Explain your answer.**

_____ **[1]**

(ii) Construct an equation for this reaction.

_____ **[2]**

[TOTAL: 15]

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QUESTION 8 BEGINS ON PAGE 34

8 Dimethylglyoxime, DMGH, can be used to analyse nickel(II) compounds.

An excess of a solution of DMGH is added to an acidic solution of a nickel(II) compound. Aqueous ammonia is added which precipitates out a nickel(II) complex, $\text{Ni}(\text{DMG})_2$, as a red solid.

(a) Complete the electron configurations of nickel as the element and in the +2 oxidation state.

nickel as the element:

$1s^2 2s^2 2p^6$ _____

nickel in the +2 oxidation state:

$1s^2 2s^2 2p^6$ _____

[2]

A sample of a hydrated nickel(II) salt is analysed using the procedure below.

STEP 1

2.50 g of the hydrated nickel(II) salt is dissolved in dilute acid.

An excess of an aqueous solution of DMGH is added.

STEP 2

An excess of aqueous ammonia is added and the mixture is heated.

A red precipitate of $\text{Ni}(\text{DMG})_2$ forms.

An equation for the reaction is shown below.



RED PRECIPITATE

STEP 3

The red precipitate is filtered, washed with water, dried and then weighed.

The precipitate of $\text{Ni}(\text{DMG})_2$ has a mass of 2.57 g.

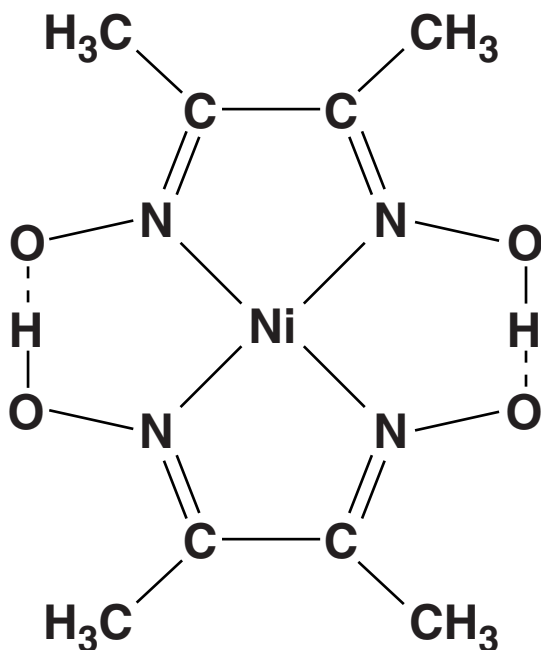
Assume that all $\text{Ni}^{2+}(\text{aq})$ ions have been converted into $\text{Ni}(\text{DMG})_2(\text{s})$.
 $M[\text{Ni}(\text{DMG})_2] = 288.7 \text{ g mol}^{-1}$.

STEP 4

A second 2.50 g sample of the hydrated nickel(II) salt is heated in a crucible to remove the water of crystallisation.

1.38 g of the anhydrous salt remains.

(b) The structure of $\text{Ni}(\text{DMG})_2$ is shown below.



(i) State and explain the role of ammonia in STEP 2 of this experiment.

_____ [1]

(ii) State the coordination number of Ni in $\text{Ni}(\text{DMG})_2$.

_____ [1]

(iii) Why does the $\text{Ni}(\text{DMG})_2$ complex have no overall charge?

_____ [1]

(iv) Draw the structure of dimethylglyoxime, DMGH.

[1]

(c) Determine a possible formula of the hydrated nickel(II) salt.

Your answer MUST show relevant working.

[7]

[TOTAL: 13]

END OF QUESTION PAPER

ADDITIONAL ANSWER SPACE

If additional answer space is required, you should use the following lined page(s). The question number(s) must be clearly shown in the margins.

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