

Do not write
outside the
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						Candidate Number					
Model Answers.											
Candidate Names											
Candidate Signature											



General Certificate of Education
Advanced Level Examination
June 2015

Chemistry

CHEM5

Unit 5 Energetics, Redox and Inorganic Chemistry

Monday 15 June 2015 1.30 pm to 3.15 pm

For this paper you must have:

- the Periodic Table/Data Sheet, provided as an insert (enclosed)
- a ruler with millimetre measurements
- a calculator.

Time allowed

- 1 hour 45 minutes

Instructions

- Use black ink or black ball-point pen.
- Fill in the boxes at the top of this page.
- Answer **all** questions.
- You must answer the questions in the spaces provided. Do not write outside the box around each page or on blank pages.
- All working must be shown.
- Do all rough work in this book. Cross through any work you do not want to be marked.

Information

- The marks for questions are shown in brackets.
- The maximum mark for this paper is 100.
- You are expected to use a calculator, where appropriate.
- The Periodic Table/Data Sheet is provided as an insert.
- Your answers to the questions in **Section B** should be written in continuous prose, where appropriate.
- You will be marked on your ability to:
 - use good English
 - organise information clearly
 - use scientific terminology accurately.

Advice

- You are advised to spend about 70 minutes on **Section A** and about 35 minutes on **Section B**.

For Examiner's Use	
Examiner's Initials	
Question	Mark
1	
2	
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7	
8	
9	
TOTAL	



JUN15CHEM501

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CHEM5

Section A

Answer **all** questions in the spaces provided.

- 1 (a) Define the term **electron affinity** for chlorine.

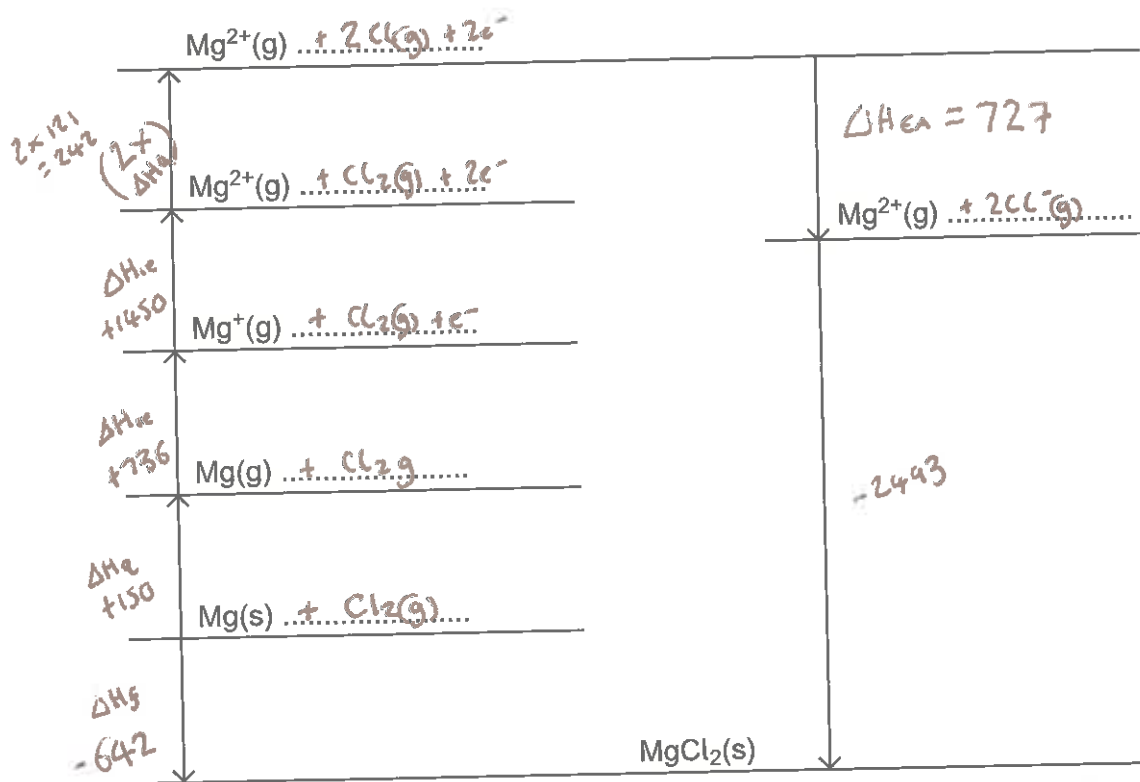
[2 marks]

The enthalpy change for the formation of one mole of chloride ions from chlorine atoms in their gaseous states.

- 1 (b) Complete this Born-Haber cycle for magnesium chloride by giving the missing species on the dotted lines. Include state symbols where appropriate.

The energy levels are **not** drawn to scale.

[6 marks]



1 (c) Table 1 contains some enthalpy data.

Table 1

	Enthalpy change / kJ mol^{-1}
Enthalpy of atomisation of magnesium	+150
Enthalpy of atomisation of chlorine	+121
First ionisation energy of magnesium	+736
Second ionisation energy of magnesium	+1450
Enthalpy of formation of magnesium chloride	-642
Lattice enthalpy of formation of magnesium chloride	-2493

Use your Born–Haber cycle from Question 1 (b) and data from Table 1 to calculate a value for the electron affinity of chlorine.

[3 marks]

[See cycle on left!]

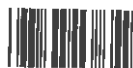
$$(642 + 150 + 736 + 1450 + 242) - 2493 = \Delta H_{EA}$$

$\Delta H_{EA} = -727$ but that is for $2\text{Cl} \rightarrow 2\text{Cl}^-$ so....

$$\frac{-727}{2} = -363.5 \text{ kJ mol}^{-1}$$

Question 1 continues on the next page

Turn over ►



1 (d) Table 2 contains some more enthalpy data.

Table 2

	Enthalpy change / kJ mol^{-1}
Enthalpy of hydration of Mg^{2+} ions	-1920
Enthalpy of hydration of Na^+ ions	-406
Enthalpy of hydration of Cl^- ions	-364

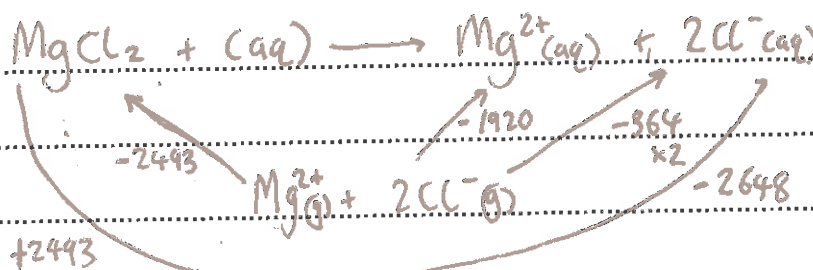
1 (d) (i) Explain why there is a difference between the hydration enthalpies of the magnesium and sodium ions.

[2 marks]

Mg^{2+} has a greater charge density than Na^+ and therefore attracts the water more strongly.

1 (d) (ii) Use data from Table 1 and Table 2 to calculate a value for the enthalpy change when one mole of magnesium chloride dissolves in water.

[2 marks]



$$+2493 + (-2648) = -155$$



Turn over for the next question

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ANSWER IN THE SPACES PROVIDED

Turn over ►



2 Table 3 contains some bond enthalpy data.

Table 3

Bond	H—H	O=O	H—O
Bond enthalpy / kJ mol ⁻¹	436	496	464

2 (a) The value for the H—O bond enthalpy in Table 3 is a mean bond enthalpy.

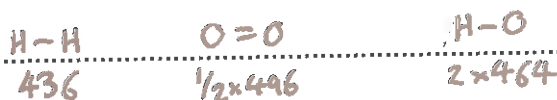
State the meaning of the term **mean bond enthalpy** for the H—O bond.

[2 marks]

The enthalpy change to break 1 mol of H—O bonds averaged over a range of different compounds.

2 (b) Use the bond enthalpies in Table 3 to calculate a value for the enthalpy of formation of water in the gas phase.

[3 marks]



Σ bonds broken - Σ bonds formed

$$436 + (\frac{1}{2} \times 496) \quad 2 \times 464$$

$$= 684 \quad = 928$$

$$684 - 928 = -244 \text{ kJ mol}^{-1}$$



2 (c) The standard enthalpy of combustion of hydrogen, forming water in the gas phase, is almost the same as the correct answer to Question 2 (b).

2 (c) (i) Suggest **one** reason why you would expect the standard enthalpy of combustion of hydrogen to be the same as the answer to Question 2 (b).

[1 mark]

It would be the same equation.

2 (c) (ii) Suggest **one** reason why you would expect the standard enthalpy of combustion of hydrogen to differ slightly from the answer to Question 2 (b).

[1 mark]

Actual bond enthalpies would differ slightly to mean bond enthalpies.

7

Turn over for the next question

Turn over ►



3 Hydrogen can be manufactured from the reaction of steam with methane.



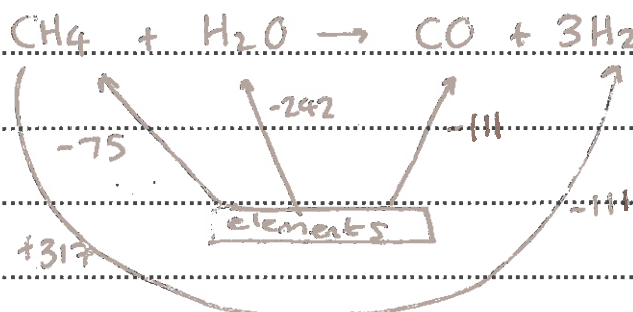
3 (a) Table 4 contains some enthalpy of formation and entropy data.

Table 4

Substance	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$	$S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$
$\text{CH}_4(\text{g})$	-75	186
$\text{H}_2\text{O}(\text{g})$	-242	189
$\text{CO}(\text{g})$	-111	198
$\text{H}_2(\text{g})$	0	131
$\text{CO}_2(\text{g})$	-394	214

3 (a) (i) Use data from Table 4 to calculate the enthalpy change, ΔH , for the reaction of steam with methane.

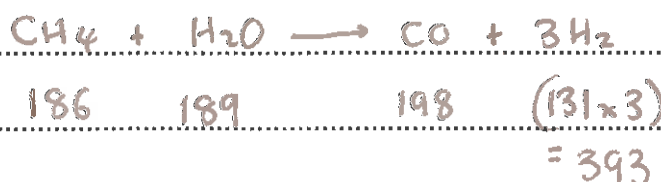
[3 marks]



$$317 + (-111) = \underline{\underline{+206 \text{ kJ mol}^{-1}}}$$

3 (a) (ii) Use data from Table 4 to calculate the entropy change, ΔS , for the reaction of steam with methane.

[2 marks]



$$\sum \text{products} - \sum \text{reactants} \quad \therefore \quad 591 - 375 = 216 \text{ J K}^{-1} \text{ mol}^{-1}$$



- 3 (b) Use your values of ΔH and ΔS from Questions 3 (a) (i) and 3 (a) (ii) to calculate the temperature above which this reaction is feasible.

[4 marks]

$$T = \frac{\Delta H}{\Delta S} = \frac{206}{(216/1000)} = \underline{\underline{953.7 \text{ K}}}$$

↑
this is for the point of
feasibility or $\Delta G = 0$

- 3 (c) The temperature used for this manufacture of hydrogen is usually about 1300 K.

Suggest **one** reason, other than changing the position of equilibrium, why this temperature is used rather than the value that you calculated in Question 3 (b).

[1 mark]

To speed up the reaction rate

Question 3 continues on the next page

Turn over ►



- 3 (d) Hydrogen can also be obtained by reaction of carbon monoxide with steam.



- 3 (d) (i) Explain, using a calculation, why this reaction should **not** occur at 1300 K.

[3 marks]

$$\Delta G = \Delta H - T\Delta S \quad \Delta G = -41 - (1300 \times (-42/1000))$$

$$\Delta G = -41 - (-54.6) \quad \Delta G = +13.6$$

ΔG must be negative (or zero) for reaction to be feasible.

- 3 (d) (ii) Explain how the conditions for the reaction could be changed to allow this reaction to take place.

[2 marks]

A lower temperature could be used which could make this feasible

$$T = \frac{\Delta H}{\Delta S} \quad T = \frac{-41}{(-42/1000)} \quad T = 976.2 \text{ K}$$

a temperature below 976.2 K needs to be used.
This would cause ΔG to be negative.



- 4 (a) Write an equation for the reaction that occurs when magnesium is heated in steam. Describe what you would observe when this reaction occurs.

[3 marks]

Equation $Mg + H_2O \rightarrow MgO + H_2$

Observations white solid and bright white light.

- 4 (b) Write an equation for the reaction that occurs when sodium is heated in oxygen. Describe what you would observe when this reaction occurs.

[3 marks]

Equation $4Na + O_2 \rightarrow 2Na_2O$

Observations white solid and orange/yellow flame.

Turn over for the next question

Turn over ►



- 5 (a) Table 5 contains data that show a trend in the melting points of some oxides of the Period 3 elements.

Table 5

Oxide	Sodium oxide	Magnesium oxide	Aluminium oxide	Silicon(IV) oxide	Phosphorus(V) oxide	Sulfur(IV) oxide
Melting point / K		3125	2345	1883	573	

- 5 (a) (i) Use data from Table 5 to predict an approximate melting point for sodium oxide.

Tick (✓) one box.

[1 mark]

250 K

500 K

1500 K

3500 K

- 5 (a) (ii) Explain, in terms of structure and bonding, why sodium oxide has a high melting point. [2 marks]

Giant ionic lattice. Strong attraction between positively charged Na^+ and negatively charged O^{2-} . This means lots of energy needed to overcome this attraction.



5 (a) (iii) Use data from Table 5 to predict a value for the melting point of sulfur(IV) oxide.

Suggest, in terms of structure and bonding, why the melting point of sulfur(IV) oxide is different from that of phosphorus(V) oxide.

[3 marks]

Predicted melting point of sulfur(IV) oxide 200K

Why the melting point is different from phosphorus(V) oxide

SO₂ is smaller than P₄O₁₀ so van der Waals between molecules are weaker so less energy needed to overcome this.

5 (b) Write an equation for the reaction of sulfur(IV) oxide with water.

Suggest the pH value of the resulting solution.

[2 marks]

Equation



pH value 1

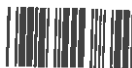
5 (c) Silicon(IV) oxide is insoluble in water.

Explain, using an equation, why silicon(IV) oxide is classified as an acidic oxide.

[2 marks]



It reacts with bases!



6 Table 6 contains some standard electrode potential data.

Table 6

Electrode half-equation	E^\ominus / V
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Au^+ + e^- \longrightarrow Au$	+1.68
$2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$	+1.64
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Ag^+ + e^- \longrightarrow Ag$	+0.80
$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$Fe^{2+} + 2e^- \longrightarrow Fe$	-0.44

6 (a) In terms of electrons, explain the meaning of the term oxidising agent.

[1 mark]

An electron acceptor.

6 (b) Identify the weakest oxidising agent in Table 6.
Explain your choice.

[2 marks]

Weakest oxidising agent Fe^{2+}

Explanation Has the most negative E^\ominus value.



- 6 (c) Write the conventional representation of the cell used to measure the standard electrode potential for the Ag^+/Ag electrode.

State the conditions necessary when measuring this value.

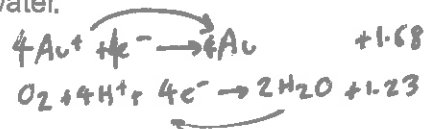
[4 marks]

Conventional representation $\text{Pt} | \text{H}_2 | \text{H}^+ || \text{Ag}^+ | \text{Ag}$

Conditions $298\text{K}, 1\text{mol dm}^{-3}, 100\text{kPa}$.

- 6 (d) Use data from Table 6 to explain, in terms of redox, what happens when a soluble gold(I) compound containing Au^+ ions is added to water.

State what you would observe.



Write an equation for the reaction that occurs.

[4 marks]

Explanation $E_{\text{cell}} = R - L$ $E_{\text{cell}} = +1.68 - (+1.23)$

$$E_{\text{cell}} = +0.45\text{V}$$

So Au^+ would oxidise water.

Observation Gold metal solid and bubbles of O_2 .



Question 6 continues on the next page

Turn over ►



Table 6 is repeated below to help you answer these questions.

Table 6

Electrode half-equation	E^\ominus / V
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Au^+ + e^- \longrightarrow Au$	+1.68
$2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$	+1.64
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Ag^+ + e^- \longrightarrow Ag$	+0.80
$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$Fe^{2+} + 2e^- \longrightarrow Fe$	-0.44

6 (e) A cell is made by connecting Fe^{2+}/Fe and Ag^+/Ag electrodes with a salt bridge.

6 (e) (i) Calculate the e.m.f. of this cell.

[1 mark]

$$E_{cell} = R - L \quad E_{cell} = 0.80 - (-0.44)$$

$$\text{Answer} = +1.24V$$

6 (e) (ii) Suggest why potassium chloride would **not** be suitable for use in the salt bridge of this cell.

[1 mark]

Chloride ions would react with Ag^+ to form $AgCl$.



- 6 (f) Use data from Table 6 to explain what happens when a solution of iron(II) chloride is exposed to the air.

[2 marks]

$$E_{\text{cell}} = R - L$$

$$E_{\text{cell}} = +0.23 - (+0.77)$$

$$E_{\text{cell}} = 0.46\text{V}$$

Fe^{2+} would be oxidised to Fe^{3+} , or Fe^{2+} reduces

O_2 to water.

15

Turn over ►



Section B

Answer **all** questions in the spaces provided.

- 7 (a) A sample of solid chromium(III) hydroxide displays amphoteric character when treated separately with dilute hydrochloric acid and with dilute aqueous sodium hydroxide.

Write an ionic equation for each of these reactions. Include the formula of each complex ion formed.

Describe the changes that you would observe in each reaction.

[5 marks]



Green solid \rightarrow Green solution.



Green solid \rightarrow Green solution.



- 7 (b) Aqueous solutions of copper(II) sulfate and cobalt(II) sulfate undergo ligand substitution reactions when treated separately with an excess of dilute aqueous ammonia.

Write equations for these reactions. Include the formulae for any complex ions. Describe the changes that you would observe in each reaction.

[6 marks]



Blue solution \rightarrow Deep Blue solution



Pink solution \rightarrow Straw coloured solution.

Turn over for the next question

Turn over ►



8 A green solution, X, is thought to contain $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ ions.

8 (a) The presence of these ions can be confirmed by reacting separate samples of solution X with aqueous ammonia and with aqueous sodium carbonate.

Write equations for each of these reactions and describe what you would observe.

[4 marks]



Green solution \rightarrow Green ppt



Green solution \rightarrow Green ppt



- 8 (b) A 50.0 cm³ sample of solution X was added to 50 cm³ of dilute sulfuric acid and made up to 250 cm³ of solution in a volumetric flask.

A 25.0 cm³ sample of this solution from the volumetric flask was titrated with a 0.0205 mol dm⁻³ solution of KMnO₄

At the end point of the reaction, the volume of KMnO₄ solution added was 18.70 cm³.

- 8 (b) (i) State the colour change that occurs at the end point of this titration and give a reason for the colour change.

[2 marks]

green solution to purple. After end point MnO₄⁻ ions become present which are purple.

- 8 (b) (ii) Write an equation for the reaction between iron(II) ions and manganate(VII) ions.

Use this equation and the information given to calculate the concentration of iron(II) ions in the original solution X.

[5 marks]



$$C = \frac{n}{V}$$

$$\left(\frac{18.7}{1000}\right) \times 0.0205 = 3.8 \times 10^{-4} \text{ moles of KMnO}_4$$

$$\therefore \times 5 = 1.92 \times 10^{-3} \text{ moles of Fe}^{2+}$$

$$\times 10 \text{ (for 250 cm}^3\text{)} = 1.92 \times 10^{-2} \text{ moles}$$

$$\frac{1.92 \times 10^{-2}}{(50/1000)} = 0.384 \text{ mol dm}^{-3}$$

$$\frac{1.92 \times 10^{-2}}{(50/1000)}$$

Turn over for the next question

Turn over ▶



9 The redox reaction, in aqueous solution, between acidified potassium manganate(VII) and sodium ethanedioate is autocatalysed.

9 (a) Write an equation for this redox reaction.

Identify the species that acts as the catalyst.

Explain how the properties of the species enable it to act as a catalyst in this reaction.

[6 marks]



Mn^{2+} is the catalyst.

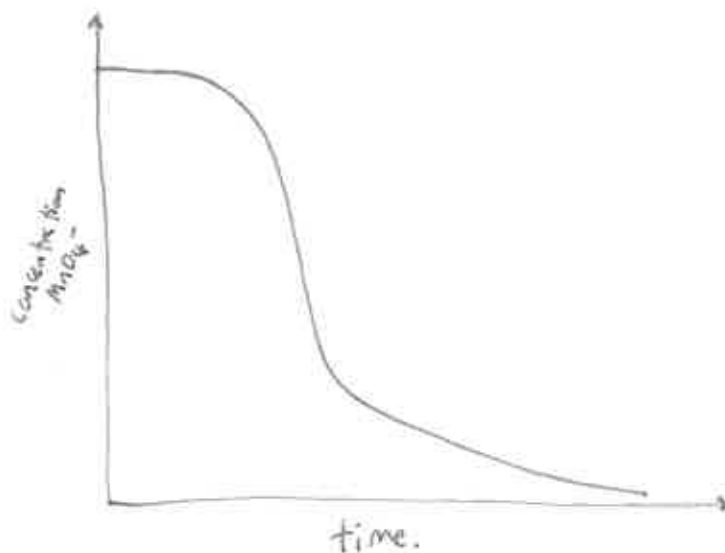
It is a good catalyst due to its variable oxidation states.

Initially slow due to two negative ions repelling each other but formation of Mn^{2+} initially then forms Mn^{3+} intermediate which can react with ethanedioate ions and due to opposite charges lowers E_a . the Mn^{3+} is then converted back to Mn^{2+} species.



- 9 (b) Sketch a graph to show how the concentration of MnO_4^- ions varies with time in this reaction.
Explain the shape of the graph.

[4 marks]



slow at first (uncatalysed) Mn^{2+} then formed and rate
increases as this catalyses the reaction. This starts to
decrease as reactants are used up.

10

END OF QUESTIONS



There are no questions printed on this page

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