Candidate Name	Centre Number	Candidate Number
		2

GCE A level

335/01

CHEMISTRY CH5

A.M. THURSDAY, 19 June 2008

1 hour 40 minutes

FOR EXAMINER'S USE ONLY		
Section	Question	Mark
A	1	
	2	
	3	
B	4	
	5	
TOTAL MARK		

ADDITIONAL MATERIALS

In addition to this examination paper, you will need:
- a calculator;
- an 8 page answer book;
- a copy of the **Periodic Table** supplied by WJEC. Refer to it for any **relative atomic masses** you require.

INSTRUCTIONS TO CANDIDATES

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

Section A Answer **all** the questions in the spaces provided.
Section B Answer **both** questions in **Section B** in a separate answer book which should then be placed inside this question-and-answer book.

Candidates are advised to allocate their time appropriately between **Section A (35 marks)** and **Section B (40 marks)**.

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 75.

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

You are reminded that marking will take into account the Quality of Written Communication used in all your written answers

SECTION A

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

1. *(a)* The thermal decomposition of ozone is shown in the equation below.

$$2O_3(g) \longrightarrow 3O_2(g)$$

Kinetic studies have shown that the reaction is second order with respect to ozone.

(i) Write the rate equation for the reaction and use it to explain the term ***order of reaction***. [2]

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(ii) The value of the rate constant at 298 K is $3{\cdot}4 \times 10^{-5}$ dm^3 $mol^{-1}s^{-1}$. If the concentration of ozone is 0·023 mol dm^{-3}, calculate the rate of reaction at 298 K and state its units. [2]

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(iii) In the stratosphere, chlorine radicals act as catalysts and speed up the decomposition of ozone.

Explain how catalysts increase the rate of reaction. [2]

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(iv) Give **one** problem caused as a consequence of ozone depletion. [1]

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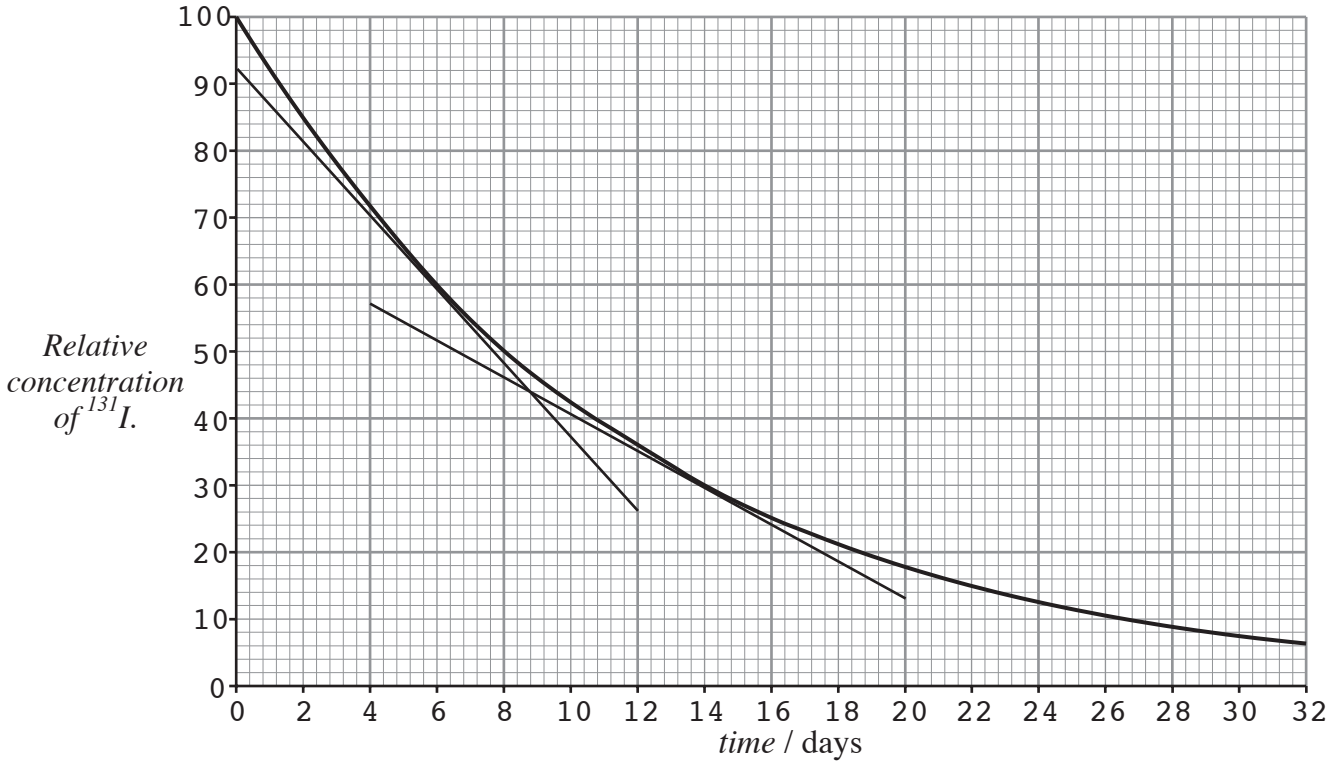
Examiner only

(b) Radioactive decay shows first-order kinetics.

(i) The graph below shows the decay of a sample of radioactive iodine ^{131}I.



I. Use the graph to find the half-life of ^{131}I. [1]

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II. Using the tangents drawn at relative concentrations of ^{131}I of 60 and 30, or otherwise, explain how the graph shows that the radioactive decay is first-order. [2]

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(ii) Archaeologists can determine the age of dead organic matter by measuring the proportion of radioactive carbon-14 present.

Carbon-14 emits β-particles and has a half-life of 5570 years.

I. Calculate the age of a piece of wood found to contain $\frac{1}{8}$ th as much ^{14}C as living material. [1]

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II. Write an equation to show the radioactive decay of ^{14}C by β emission. [1]

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Total [12]

Turn over.

2. *(a)* Iodine and chlorine both react with aqueous sodium thiosulphate, but in different ways.

(i) Complete equation **A** for the reaction of iodine with thiosulphate ions. [1]

Equation **A**: $I_2 + 2S_2O_3^{2-} \longrightarrow$ ..... + ...

(ii) When chlorine is passed into an aqueous solution containing thiosulphate ions, the products are sulphate ions (SO_4^{2-}), chloride ions (Cl^-) and hydrogen ions (H^+).

Write the half equations for:

I. the reduction of chlorine to chloride ions; [1]

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II. the oxidation of thiosulphate ions in water to sulphate ions and hydrogen ions. [1]

Half equation **B**: .....

III. Hence, write the overall equation for this reaction. [1]

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(iii) I. State the initial and final oxidation numbers of sulphur in both equations **A** and **B** by completing the table below. [1]

	Initial oxidation number of sulphur	Final oxidation number of sulphur
Equation **A**		
Half equation **B**		

II. Explain how these values show that chlorine is a stronger oxidising agent than iodine. [1]

Explanation .....

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(b) Chlorine can react with aqueous sodium hydroxide to form sodium chloride, sodium chlorate(I) and sodium chlorate(V).

Give **one** large scale use for

sodium chloride .....

s*odium chlorate(V)* .....

[1]

(c) Sodium chlorate(I) is used in bleach. The concentration of sodium chlorate(I) in domestic bleach can be found by reacting the bleach with an acidified iodide solution to form iodine and then titrating with a thiosulphate solution.

(i) Name a suitable indicator for this reaction. [1]

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(ii) The equation for this reaction is given below.

$$ClO^-(aq) + 2H^+(aq) + 2I^-(aq) \longrightarrow I_2(aq) + Cl^-(aq) + H_2O(l)$$

A 25·0 cm^3 sample of domestic bleach was diluted to 250 cm^3 in a volumetric flask. 25·0 cm^3 of this solution was added to an excess of acidified potassium iodide and the iodine produced reacted with 20·4 cm^3 of aqueous sodium thiosulphate of concentration 0·092 mol dm^{-3}.

Use the equation above, together with equation **A** in *(a)*(i), to calculate the concentration of sodium chlorate(I) in the original bleach sample to three significant figures. [4]

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..... mol dm^{-3}

(d) Describe how you could distinguish between separate aqueous solutions containing chloride ions and iodide ions. [2]

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Total [14]

3. In living systems, enzymes that enable biochemical reactions to take place can function only within a narrow pH range. An important buffer system in human beings is the carbonate buffer, consisting of carbonic acid and its conjugate base, the hydrogencarbonate ion.

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

(K_a for carbonic acid = $4{\cdot}5 \times 10^{-7}$ mol dm^{-3} at body temperature)

(a) Write an expression for the dissociation constant, K_a, for the acid. [1]

(b) Define the term pH. [1]

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(c) Calculate the pH of blood in a person, given that at body temperature: [3]

$[H_2CO_3] = 1{\cdot}85 \times 10^{-3}$ mol dm^{-3} $[HCO_3^-] = 2{\cdot}09 \times 10^{-2}$ mol dm^{-3}

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(d) State the purpose of a buffer solution. [1]

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(e) Explain how an aqueous solution of ethanoic acid and sodium ethanoate can act as a buffer solution when a small amount of acid or alkali is **separately** added to it. [3]

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Total [9]

Total Section A [35]

SECTION B

Answer **both** *questions in this section in the separate answer book provided.*

4. Copper and iron are typical transition metals.

(a) A characteristic of these metals is the ability to form coloured ions.

(i) Give the electronic configuration of copper(I) ions, Cu^+, and state why copper(I) compounds are not usually coloured. [2]

(ii) Copper(II) ions can form the following coloured complexes:

$$[Cu(H_2O)_6]^{2+} \text{ and } [CuCl_4]^{2-}$$

I. State the shape and colour for **each** complex. [4]

II. Describe the bonding in copper(II) complexes. [2]

(b) (i) Define the term ***standard molar enthalpy change of formation***, $\Delta H_f^{\ominus}$ [1]

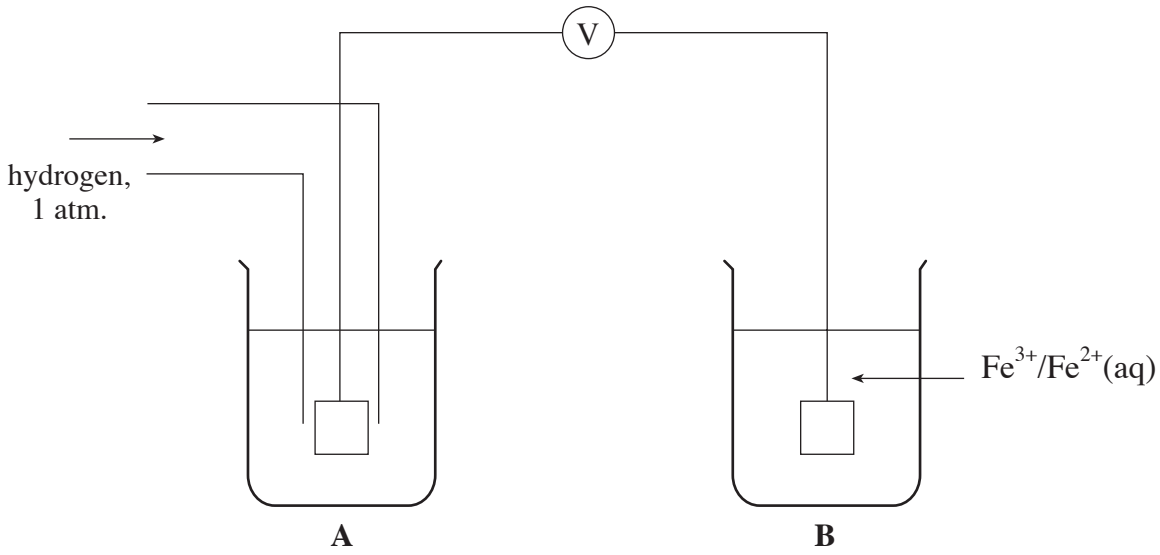
(ii) The enthalpy of formation of copper(II) fluoride, CuF_2, can be determined indirectly using a Born-Haber cycle.
Use the data given below to calculate the enthalpy of formation of copper(II) fluoride in kJ mol^{-1}. [4]

Process	$\Delta H^{\ominus}$ / kJ mol^{-1}
$Cu(s) \longrightarrow Cu(g)$	339
$\frac{1}{2}F_2(g) \longrightarrow F(g)$	79
$Cu(g) \longrightarrow Cu^+(g) + e^-$	745
$Cu^+(g) \longrightarrow Cu^{2+}(g) + e^-$	1960
$F(g) + e^- \longrightarrow F^-(g)$	–348
$Cu^{2+}(g) + 2F^-(g) \longrightarrow CuF_2(s)$	–3037

(c) Another transition metal characteristic is the ability to show variable oxidation states.

The diagram below shows part of the apparatus that was used to measure the standard electrode potential of the $Fe^{3+}(aq)/Fe^{2+}$ (aq) half cell.

The standard electrode potential of the half cell was +0·77V.



(i) Name a solution that must be placed in beaker **A** and state its concentration. [1]

(ii) Name a metal that could be used as an electrode in beaker **B**. [1]

(iii) Name the part of the cell that is missing and state its purpose. [1]

(iv) State, giving a reason, the direction in which the electrons flow along the wire through the voltmeter. [1]

(v) The standard electrode potential for the $Cu^{2+}(aq)/Cu(s)$ electrode is +0·34 V. If the hydrogen electrode is replaced by the $Cu^{2+}(aq)/Cu(s)$ electrode, calculate the new reading on the voltmeter, V. [1]

(d) Some standard electrode potentials, $E^{\ominus}$, are given below.

System	$E^{\ominus}$ / volts
$\frac{1}{2}I_2(s) + e^- \rightleftharpoons I^-(aq)$	0·54
$\frac{1}{2}Br_2(l) + e^- \rightleftharpoons Br^-(aq)$	1·09
$\frac{1}{2}Cl_2(g) + e^- \rightleftharpoons Cl^-(aq)$	1·36

Using the information from the table, state which of the halides will reduce Fe^{3+} to Fe^{2+}. Give a reason for your answer. [2]

Total [20]

Turn over.

5. Two important trends in inorganic chemistry are that:

- metallic character increases down a group;
- metallic character decreases across a period.

(a) Show how the first trend is true by describing the structure and bonding of carbon (diamond) and of lead. [5]

(b) Group IV elements can show oxidation states of II and IV in their compounds. State, giving examples, how the relative stability of these oxidation states changes as the group is descended and give a reason for this trend. [3]

(c) Explain why the reactivity of Group I metals increases as the group is descended. [2]

(d) State what would be **observed** if sodium hydride were added to water and predict the pH of the resulting solution. [2]

(e) State the trend in the bonding of the chlorides across Period 3. [1]

(f) Write equations to show what happens when sodium chloride and phosphorus(V) chloride are separately added to water. Predict a value for the pH of the resulting solutions and explain how your answer confirms the trend given in part *(e)*. [5]

(g) Name an oxide of one of the elements of Period 2 or 3 which would dissolve in water to give an acidic solution and give the formula of the salt which would form if the acidic solution were added to aqueous sodium hydroxide. [2]

Total [20]

Total Section B [40]