Q1. Which of these species is the best reducing agent?

A Cl_2 ⬭

B Cl^- ⬭

C I_2 ⬭

D I^- ⬭

(Total 1 mark)

Q2. Which of the following shows chlorine in its correct oxidation states in the compounds shown?

	HCl	$KClO_3$	HClO	
A	−1	+3	+1	⬭
B	+1	−5	−1	⬭
C	−1	+5	+1	⬭
D	+1	+5	−1	⬭

(Total 1 mark)

Q3. Which substance is **not** produced in a redox reaction when solid sodium iodide reacts with concentrated sulfuric acid?

A H_2S ⬭

B HI ⬭

C SO_2 ⬭

D I_2 ⬭

(Total 1 mark)

Q4. In which reaction is hydrogen acting as an oxidising agent?

A $Cl_2 + H_2 \longrightarrow 2HCl$ ⬭

B $(CH_3)_2CO + H_2 \longrightarrow (CH_3)_2CHOH$ ⬭

C $N_2 + 3H_2 \longrightarrow 2NH_3$ ⬭

D $2Na + H_2 \longrightarrow 2NaH$ ⬭

(Total 1 mark)

Q5. In which reaction is the metal oxidised?

A $2Cu^{2+} + 4I^- \longrightarrow 2CuI + I_2$ ⬭

B $[Fe(H_2O)_6]^{3+} + Cl^- \longrightarrow [Fe(H_2O)_5(Cl)]^{2+} + H_2O$ ⬭

C $[CoCl_4]^{2-} + 6H_2O \longrightarrow [Co(H_2O)_6]^{2+} + 4Cl^-$ ⬭

D $Mg + S \longrightarrow MgS$ ⬭

(Total 1 mark)

Q6. Which one of the following is the electron arrangement of the strongest reducing agent?

A $1s^2\ 2s^2\ 2p^5$

B $1s^2\ 2s^2\ 2p^6\ 3s^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

(Total 1 mark)

Q7. Which one of the following is **not** a redox reaction?

A $Br_2 + SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2Br^-$

B $SnCl_2 + HgCl_2 \rightarrow Hg + SnCl_4$

C $Cu_2O + H_2SO_4 \rightarrow CuSO_4 + Cu + H_2O$

D $2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$

(Total 1 mark)

Q8. Refer to the following reaction

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \quad \Delta H^{\ominus} = -11 \text{ kJ mol}^{-1}, \quad \Delta S^{\ominus} = +20 \text{ J K}^{-1} \text{ mol}^{-1}$$

Which one of the following statements is correct?

A This is a redox reaction.

B The reaction is **not** feasible below 298 K

C At equilibrium, the yield of hydrogen iodide is changed by increasing the pressure.

D At equilibrium, the yield of hydrogen iodide increases as the temperature is increased.

(Total 1 mark)

Q9. Photochromic glass contains silver ions and copper ions. A simplified version of a redox equilibrium is shown below. In bright sunlight the high energy u.v. light causes silver atoms to form and the glass darkens. When the intensity of the light is reduced the reaction is reversed and the glass lightens.

$$Cu^{+}(s) + Ag^{+}(s) \rightleftharpoons Cu^{2+}(s) + Ag(s)$$

clear glass dark glass

When the photochromic glass darkens

A the Ag^{+} ion is acting as an electron donor.

B the Cu^{+} ion is acting as a reducing agent.

C the Ag^{+} ion is oxidised.

D the Cu^{+} ion is reduced.

(Total 1 mark)

Q10. Which one of the following statements is **not** correct?

A The first ionisation energy of iron is greater than its second ionisation energy.

B The magnitude of the lattice enthalpy of magnesium oxide is greater than that of barium oxide.

C The oxidation state of iron in $[Fe(CN)_6]^{3-}$ is greater than the oxidation state of copper in $[CuCl_2]^{-}$

D The boiling point of C_3H_8 is lower than that of CH_3CH_2OH

(Total 1 mark)

Q11. The vanadium does **not** have an oxidation state of +3 in

A $[V(H_2O)_6]^{3+}$

B $[V(C_2O_4)_3]^{3-}$

C $[V(OH)_3(H_2O)_3]$

D $[VCl_4]^{3-}$

(Total 1 mark)

Q12. In which one of the following reactions does hydrogen **not** act as a reducing agent?

A $H_2 + Ca \rightarrow CaH_2$

B $2H_2 + O_2 \rightarrow 2H_2O$

C $H_2 + CH_2{=}CH_2 \rightarrow CH_3CH_3$

C $2H_2 + CH_3COCH_3 \rightarrow CH_3CH_2CH_3 + H_2O$

(Total 1 mark)

Q13. In which one of the following reactions is the role of the reagent stated correctly?

	Reaction	**Role of reagent**
A	$TiO_2 + 2C + 2Cl_2 \rightarrow TiCl_4 + 2CO$	TiO_2 is an oxidising agent
B	$HNO_3 + H_2SO_4 \rightarrow H_2NO_3^+ + HSO_4^-$	HNO_3 is a Brønsted-Lowry acid
C	$CH_3COCl + AlCl_3 \rightarrow CH_3CO^+ + AlCl_4^-$	$AlCl_3$ is a Lewis base
D	$2CO + 2NO \rightarrow 2CO_2 + N_2$	CO is a reducing agent

(Total 1 mark)

Q14. Which one of the following is a redox reaction?

A $2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$

B $3Cl_2 + 6OH^- \rightarrow 5Cl^- + ClO_3^- + 3H_2O$

C $HNO_3 + 2H_2SO_4 \rightarrow NO_2^+ + H_3O^+ + 2HSO_4^-$

D $CaCO_3 + SiO_2 \rightarrow CaSiO_3 + CO_2$

(Total 1 mark)

Q15. Which one of the following is **not** a redox reaction?

A $TiO_2 + 2Cl_2 + C \rightarrow TiCl_4 + CO_2$

B $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

C $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

D $3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$

(Total 1 mark)

Q16. In which one of the following reactions is H_2O_2 behaving as a reducing agent?

A $H_2O_2 + 2I^- + 2H^+ \rightarrow I_2 + 2H_2O$

B $H_2O_2 + 2[Co(NH_3)_6]^{2+} \rightarrow 2[Co(NH_3)_6]^{3+} + 2OH^-$

C $5H_2O_2 + - 2MnO_4^- + 6H+ \rightarrow 2Mn^{2+} + 8H_2O + 5O_2$

D $3H_2O_2 + 2[Cr(OH)_6]^{3-} \rightarrow 2CrO_4^{2-} + 8H_2O + 2OH^-$

(Total 1 mark)

Q17. In which one of the following reactions do two H ions and one electron have to be added to the left-hand side in order to balance the equation?

A $CH_3CHO \rightarrow CH_3CH_2OH$

B $VO^{2+} \rightarrow V^{3+} + H_2O$

C $NO\bar{3} \rightarrow HNO_2 + H_2O$

D $HOCl \rightarrow \frac{1}{2}Cl_2 + H_2O$

(Total 1 mark)

Q18. Which equation does **not** involve the reduction of a transition metal compound?

A $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

B $TiO_2 + 2C + 2Cl_2 \rightarrow TiCl_4 + 2CO$

C $Cr_2O_3 + 2Al \rightarrow 2Cr + Al_2O_3$

D $TiCl_4 + 4Na \rightarrow Ti + 4NaCl$

(Total 1 mark)

Q19. Which one of the following is the electronic configuration of the strongest reducing agent?

A $1s^2\ 2s^2\ 2p^5$

B $1s^2\ 2s^2\ 2p^6\ 3s^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

(Total 1 mark)

Q20. The compound lithium tetrahydridoaluminate(III), $LiAlH_4$, is a useful reducing agent. It behaves in a similar fashion to $NaBH_4$. Carbonyl compounds and carboxylic acids are reduced to alcohols. However, $LiAlH_4$ also reduces water in a violent reaction so that it must be used in an organic solvent.

Which one of the following concerning the violent reaction between $LiAlH_4$ and water is **false**?

A A gas is produced.

B The activation energy for the reaction is relatively high.

C The reaction has a negative free-energy change.

D Aqueous lithium ions are formed.

(Total 1 mark)

Q21. In which one of the following reactions does the metal species undergo reduction?

A $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

B $[Cu(H_2O)_6]^{2+} + 4Cl^- \rightarrow [CuCl_4]^{2-} + 6H_2O$

C $CrO_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$

D $TiO_2 + 2C + 2Cl_2 \rightarrow TiCl_4 + 2CO$

(Total 1 mark)

Q22. Which one of the following contains the metal with the lowest oxidation state?

A CrO_2F_2

B $[Cr_2O_7]^2$

C $[MnCl_6]^{2-}$

D $[Mn(CN)_6]^{3-}$

(Total 1 mark)

Q1. Chlorine is an important industrial chemical.

(a) Chlorine is formed when $KMnO_4$ reacts with hydrochloric acid.
The ionic equation for this redox reaction is

$$16H^+ + 2MnO_4^- + 10Cl^- \longrightarrow 2Mn^{2+} + 8H_2O + 5Cl_2$$

(i) Deduce the half-equation for the oxidation of chloride ions to chlorine.

.....

(1)

(ii) Give the oxidation state of manganese in the MnO_4^- ion.

.....

(1)

(iii) Deduce the half-equation for the reduction of the MnO_4^- ions in acidified solution to manganese(II) ions and water.

.....

(1)

(b) Chlorine behaves as an oxidising agent in the extraction of bromine from seawater.
In this process, chlorine gas is bubbled through a solution containing bromide ions.

(i) Write the **simplest ionic** equation for the reaction of chlorine with bromide ions.

.....

(1)

(ii) Give **one** observation that would be made during this reaction.

.....

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(1)

(iii) In terms of electrons, state the meaning of the term **oxidising agent**.

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(1)

(c) In sunlight, chlorine can also oxidise water slowly to form oxygen.

Write an equation for this reaction.

Give the oxidation state of chlorine in the chlorine-containing species that is formed.

Equation

.....

Oxidation state of chlorine in the species formed .....

(2)

(d) Explain why chlorine has a lower boiling point than bromine.

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(2)
(Total 10 marks)

Q2. A student carried out an experiment to find the mass of $FeSO_4.7H_2O$ in an impure sample, **X**.
The student recorded the mass of **X**. This sample was dissolved in water and made up to 250 cm^3 of solution.
The student found that, after an excess of acid had been added, 25.0 cm^3 of this solution reacted with 21.3 cm^3 of a 0.0150 mol dm^{-3} solution of $K_2Cr_2O_7$

(a) Use this information to calculate a value for the mass of $FeSO_4.7H_2O$ in the sample of **X**.

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(5)

(b) The student found that the calculated mass of $FeSO_4.7H_2O$ was greater than the actual mass of the sample that had been weighed out. The student realised that this could be due to the nature of the impurity.

Suggest **one** property of an impurity that would cause the calculated mass of $FeSO_4.7H_2O$ in **X** to be greater than the actual mass of **X**.
Explain your answer.

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(2)
(Total 7 marks)

Q3. Vanadium is an important metal. Ferrovanadium, an alloy of iron and vanadium, is used to make a strong type of vanadium-steel. Pure vanadium is used in nuclear reactors.

(a) The table shows some standard enthalpy of formation data.

	V_2O_5(s)	**CaO(s)**
ΔH_f^θ / kJ mol^{-1}	−1560	−635

In the oldest method of extraction of vanadium, V_2O_5 is reacted with calcium at a high temperature.

$$5Ca(s) + V_2O_5(s) \longrightarrow 2V(s) + 5CaO(s)$$

Use data from the table and the equation to calculate the standard enthalpy change for this reaction.

State the type of reaction that V_2O_5 has undergone.

Suggest **one** major reason why this method of extracting vanadium is expensive, other than the cost of heating the reaction mixture.

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(5)

(b) Ferrovanadium is produced by the reaction of aluminium with a mixture of V_2O_5 and iron(III) oxide.

Write an equation for the reaction of aluminium with iron(III) oxide.

State the change in oxidation state of aluminium in this reaction.

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(2)

(c) Pure vanadium, for nuclear reactors, is formed by the reaction of hydrogen with purified VCl_2

Write an equation for this reaction in which the only other product is HCl gas.

Identify **two** hazards in this process, other than the fact that it operates at a high temperature.

Deduce why this process produces **pure** vanadium, other than the fact that purified

VCl_2 is used.

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(4)
(Total 11 marks)

Q4. Antimony is a solid element that is used in industry. The method used for the extraction of antimony depends on the grade of the ore.

(a) Antimony can be extracted by reacting scrap iron with low-grade ores that contain antimony sulfide (Sb_2S_3).

(i) Write an equation for the reaction of iron with antimony sulfide to form antimony and iron(II) sulfide.

.....

(1)

(ii) Write a half-equation to show what happens to the iron atoms in this reaction.

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(1)

(b) In the first stage of the extraction of antimony from a high-grade ore, antimony sulfide is roasted in air to convert it into antimony(III) oxide (Sb_2O_3) and sulfur dioxide.

(i) Write an equation for this reaction.

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(1)

(ii) Identify **one** substance that is manufactured directly from the sulfur dioxide formed in this reaction.

.....

(1)

(c) In the second stage of the extraction of antimony from a high-grade ore, antimony(III) oxide is reacted with carbon monoxide at high temperature.

(i) Use the standard enthalpies of formation in the table and the equation given below the table to calculate a value for the standard enthalpy change for this reaction.

	$Sb_2O_3(s)$	**CO(g)**	**Sb(l)**	**$CO_2(g)$**
$\Delta H_f^{\ominus}$ / kJ mol^{-1}	-705	-111	+20	-394

$$Sb_2O_3(s) + 3CO(g) \longrightarrow 2Sb(l) + 3CO_2(g)$$

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(3)

(ii) Suggest why the value for the standard enthalpy of formation of liquid antimony, given in the table above, is **not** zero.

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(1)

(iii) State the type of reaction that antimony(III) oxide has undergone in this reaction.

.....

(1)

(d) Deduce **one** reason why the method of extraction of antimony from a low-grade ore, described in part (a), is a low-cost process. Do **not** include the cost of the ore.

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(1)

(Total 10 marks)

Q5. In which reaction is the metal oxidised?

A $2Cu^{2+} + 4I^- \longrightarrow 2CuI + I_2$ ⬭

B $[Fe(H_2O)_6]^{3+} + Cl^- \longrightarrow [Fe(H_2O)_5(Cl)]^{2+} + H_2O$ ⬭

C $[CoCl_4]^{2-} + 6H_2O \longrightarrow [Co(H_2O)_6]^{2+} + 4Cl^-$ ⬭

D $Mg + S \longrightarrow MgS$ ⬭

(Total 1 mark)

Q6. In which reaction is hydrogen acting as an oxidising agent?

A $Cl_2 + H_2 \longrightarrow 2HCl$ ⬭

B $(CH_3)_2CO + H_2 \longrightarrow (CH_3)_2CHOH$ ⬭

C $N_2 + 3H_2 \longrightarrow 2NH_3$ ⬭

D $2Na + H_2 \longrightarrow 2NaH$ ⬭

(Total 1 mark)

Q7. Which substance is **not** produced in a redox reaction when solid sodium iodide reacts with concentrated sulfuric acid?

A H_2S ⬭

B HI ⬭

C SO_2 ⬭

D I_2 ⬭

(Total 1 mark)

Q8. For many years, swimming pool water has been treated with chlorine gas. The chlorine is added to kill any harmful bacteria unintentionally introduced by swimmers. Pool managers are required to check that the chlorine concentration is high enough to kill the bacteria without being a health hazard to the swimmers.

When chlorine reacts with water in the absence of sunlight, the chlorine is both oxidised and reduced and an equilibrium is established.

(a) Write an equation for this equilibrium.

For each chlorine-containing species in the equation, write the oxidation state of chlorine below the species.

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(2)

(b) The pool manager maintains the water at a pH slightly greater than 7.0

Explain how this affects the equilibrium established when chlorine is added to water.

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(2)

(c) Explain why chlorine is used to kill bacteria in swimming pools, even though chlorine is toxic.

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(2)

(Total 6 marks)

Q9. Which of the following shows chlorine in its correct oxidation states in the compounds shown?

	HCl	$KClO_3$	HClO	
A	−1	+3	+1	⬭
B	+1	−5	−1	⬭
C	−1	+5	+1	⬭
D	+1	+5	−1	⬭

(Total 1 mark)

Q10. Which of these species is the best reducing agent?

A Cl_2 ⬭

B Cl^- ⬭

C I_2 ⬭

D I^- ⬭

(Total 1 mark)

Q1. This question is about Group 7 chemistry.

(a) Sea water is a major source of iodine.
The iodine extracted from sea water is impure. It is purified in a two-stage process.

Stage **1** $I_2 + 2H_2O + SO_2 \longrightarrow 2HI + H_2SO_4$

Stage **2** $2HI + Cl_2 \longrightarrow I_2 + 2HCl$

(i) State the initial oxidation state and the final oxidation state of sulfur in Stage **1**.

Oxidation state of S in SO_2 .....

Oxidation state of S in H_2SO_4 .....

(2)

(ii) State, in terms of electrons, what has happened to chlorine in Stage **2**.

.....

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(1)

(b) When concentrated sulfuric acid is added to potassium iodide, iodine is formed in the following redox equations.

$$\text{.....}KI + \text{.....}H_2SO_4 \longrightarrow \text{.....}KHSO_4 + \text{.....}I_2 + S + \text{.....}H_2O$$

$$8KI + 9H_2SO_4 \longrightarrow 8KHSO_4 + 4I_2 + H_2S + 4H_2O$$

(i) Balance the equation for the reaction that forms sulfur.

(1)

(ii) Deduce the half-equation for the formation of iodine from iodide ions.

.....

(1)

(iii) Deduce the half-equation for the formation of hydrogen sulfide from concentrated sulfuric acid.

.....

(1)

(c) A yellow precipitate is formed when silver nitrate solution, acidified with dilute nitric acid, is added to an aqueous solution containing iodide ions.

(i) Write the **simplest ionic** equation for the formation of the yellow precipitate.

.....

(1)

(ii) State what is observed when concentrated ammonia solution is added to this yellow precipitate.

.....

.....

(1)

(iii) State why the silver nitrate solution is acidified when testing for iodide ions.

.....

.....

.....

(1)

(iv) Explain why dilute hydrochloric acid is **not** used to acidify the silver nitrate solution in this test for iodide ions.

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(1)

(d) Chlorine is toxic to humans. This toxicity does not prevent the large-scale use of chlorine in water treatment.

(i) Give **one** reason why water is treated with chlorine.

.....

.....

(1)

(ii) Explain why the toxicity of chlorine does **not** prevent this use.

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(1)

(iii) Write an equation for the reaction of chlorine with cold water.

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(1)

(e) Give the formulas of the **two** different chlorine-containing compounds that are formed when chlorine reacts with cold, dilute, aqueous sodium hydroxide.

Formula 1 .....

Formula 2 .....

(1)

(Total 14 marks)

Q2. The table below shows some standard electrode potential data.

	$E^{\ominus}$ / V
$ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	–1.25
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	–0.44
$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$	+1.64

(a) Give the conventional representation of the cell that is used to measure the standard electrode potential of iron as shown in the table.

.....

(2)

(b) With reference to electrons, give the meaning of the term **reducing agent**.

.....

.....

(1)

(c) Identify the weakest reducing agent from the species in the table.

Explain how you deduced your answer.

Species.....

Explanation.....

.....

(2)

(d) When HOCl acts as an oxidising agent, one of the atoms in the molecule is reduced.

(i) Place a tick (✓) next to the atom that is reduced.

Atom that is reduced	**Tick (✓)**
H	
O	
Cl	

(1)

(ii) Explain your answer to part (i) in terms of the change in the oxidation state of this atom.

.....

.....

(1)

(e) Using the information given in the table, deduce an equation for the redox reaction that would occur when hydroxide ions are added to HOCl

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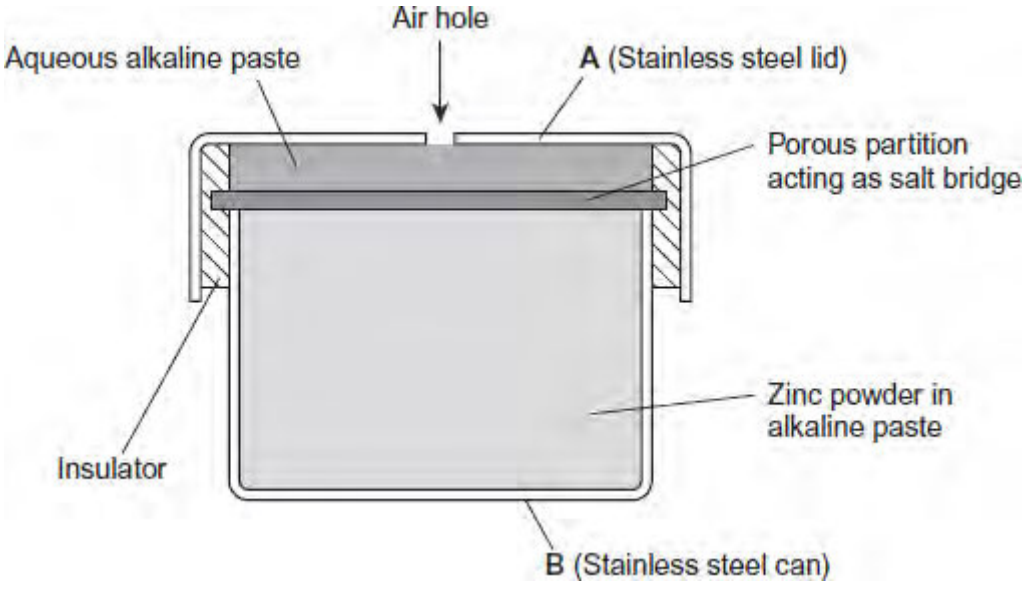
(2)

(f) The table is repeated to help you answer this question.

	$E^{\ominus}$ / V
$ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	–1.25
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	–0.44
$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$	+1.64

The half-equations from the table that involve zinc and oxygen are simplified versions of those that occur in hearing aid cells.

A simplified diagram of a hearing aid cell is shown in the following figure.



(i) Use data from the table to calculate the e.m.f. of this cell.

.....

.....

Answer = ...

(1)

(ii) Use half-equations from the table to construct an overall equation for the cell reaction.

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.....

(1)

(iii) Identify which of **A** or **B**, in the figure, is the positive electrode. Give a reason for your answer.

Positive electrode .....

Reason .....

.....

.....

(2)

(iv) Suggest **one** reason, other than cost, why this type of cell is **not** recharged.

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.....

(1)

(Total 14 marks)

Q3. The silicon chip industry requires the production of pure silicon. Silicon is extracted from its ore, silicon dioxide (SiO_2), by a process similar to that used in the extraction of titanium.

(a) (i) Write an equation for the formation of $SiCl_4$ from SiO_2 using chlorine and

carbon.

.....

(1)

(ii) Suggest how the liquid $SiCl_4$ is purified.

.....

.....

(1)

(b) The final stage in the extraction of silicon involves the use of hydrogen gas to convert the $SiCl_4$ into silicon and hydrogen chloride.

(i) Write an equation for this reaction.

.....

(1)

(ii) State the role of hydrogen in this reaction.

.....

(1)

(iii) Give **one** risk associated with the use of hydrogen gas.

.....

(1)

(c) The magnesium used to make magnesium ferrosilicon alloys is extracted from magnesium oxide using silicon.
Write an equation for this reaction to produce magnesium and silicon dioxide.

.....

(1)

(Total 6 marks)

Q4. The price of copper is increasing as supplies of high-grade ores start to run out.
The mineral covellite (CuS), found in low-grade ores, is a possible future source of copper.

(a) When copper is extracted from covellite, a reaction occurs between copper(II) sulfide and nitric acid to form a dilute solution of copper(II) sulfate.

(i) Balance the equation for this reaction.

$3CuS(s) + ...HNO_3(aq) \longrightarrow .....CuSO_4(aq) + ...NO(g) + ...H_2O(l)$

(1)

(ii) Give the oxidation state of nitrogen in each of the following.

HNO_3.....

NO .....

(2)

(iii) Deduce the redox half-equation for the reduction of the nitrate ion in acidified solution to form nitrogen monoxide and water.

.....

(1)

(iv) Deduce the redox half-equation for the oxidation of the sulfide ion in aqueous solution to form the sulfate ion and $H^+(aq)$ ions.

.....

(1)

(b) Use your knowledge of metal reactivity to state and explain a low-cost method for the extraction of copper from a dilute aqueous solution of copper(II) sulfate.
Write the **simplest ionic** equation for the reaction that occurs during this extraction process.

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Simplest ionic equation

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(4)
(Total 9 marks)

Q5. Metals can be extracted by different methods.

(a) Give **one** reason why titanium cannot be extracted directly from titanium(IV) oxide using carbon.

.....

.....

(1)

(b) Titanium steel is an alloy of titanium and iron. Titanium steel is extracted from the mineral ilmenite ($FeTiO_3$) in a two-stage process.
Purified $FeTiO_3$ is first converted into a mixture of two metal chlorides. These two metal chlorides are then reduced simultaneously using sodium.

(i) Write an equation for the reaction of $FeTiO_3$ with chlorine and carbon to produce iron(III) chloride ($FeCl_3$), titanium(IV) chloride and carbon monoxide.

.....

(1)

(ii) Write an equation for the simultaneous reduction of the mixture of iron(III) chloride and titanium(IV) chloride to iron and titanium using sodium.

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(1)

(c) Scrap iron is used to extract copper from dilute aqueous solutions containing copper(II) ions.
Explain, in terms of redox, what happens to the copper(II) ions in this extractio.

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(2)

(d) Aluminium is an expensive metal because it is extracted from molten aluminium oxide using electrolysis.
Write the half-equation for the reaction that occurs at the positive electrode during this extraction.

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(1)
(Total 6 marks)

Q6. The manufacture of food grade phosphoric acid for use in cola drinks begins with the production of pure white phosphorus from the mineral fluoroapatite, $Ca_5F(PO_4)_3$

(a) Complete the following equation for the manufacture of phosphorus.

$$....Ca_5F(PO_4)_3 + 9SiO_2 + ...C \longrightarrow 9CaSiO_3 + CaF_2 + ...CO + ...P$$

(1)

(b) As the phosphorus cools, it forms white phosphorus, P_4

Give the oxidation state of phosphorus in each of the following.

P_4 .....

H_3PO_4 .....

(2)

(c) Fertiliser grade phosphoric acid is manufactured from sulfuric acid and calcium phosphate.
Use the following precise relative atomic mass data to show how mass spectrometry can be used to distinguish between pure sulfuric acid (H_2SO_4) and pure phosphoric acid (H_3PO_4) which both have M_r = 98 to two significant figures.

Atom	Precise relative atomic mass
^{1}H	1.00794
^{16}O	15.99491
^{31}P	30.97376
^{32}S	32.06550

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(1)

(d) Concentrated phosphoric acid is used as a catalyst in the hydration of propene to form the alcohol $CH_3CH(OH)CH_3$ as the main organic product.
The industrial name for this alcohol is isopropyl alcohol.

(i) State the meaning of the term *catalyst*.

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(Extra space) .....

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(1)

(ii) State the meaning of the term *hydration*.

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(Extra space) .....

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(1)

(iii) Write an equation for the hydration of propene to form isopropyl alcohol. Give the IUPAC name for isopropyl alcohol.

Equation .....

IUPAC name .....

(2)
(Total 8 marks)

Q1.A student investigated the chemistry of the halogens and the halide ions.

(a) In the first two tests, the student made the following observations.

Test	Observation
1. Add chlorine water to aqueous potassium iodide solution.	The colourless solution turned a brown colour.
2. Add silver nitrate solution to aqueous potassium chloride solution.	The colourless solution produced a white precipitate.

(i) Identify the species responsible for the brown colour in Test **1**.

Write the **simplest ionic** equation for the reaction that has taken place in Test **1**.

State the type of reaction that has taken place in Test **1**.

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(Extra space) .....

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(3)

(ii) Name the species responsible for the white precipitate in Test **2**.

Write the **simplest ionic** equation for the reaction that has taken place in Test **2**.

State what would be observed when an excess of dilute ammonia solution is added to the white precipitate obtained in Test **2**.

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(Extra space) .....

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(3)

(b) In two further tests, the student made the following observations.

Test	Observation
3. Add concentrated sulfuric acid to solid potassium chloride.	The white solid produced misty white fumes which turned blue litmus paper to red.
4. Add concentrated sulfuric acid to solid potassium iodide.	The white solid turned black. A gas was released that smelled of rotten eggs. A yellow solid was formed.

(i) Write the **simplest ionic** equation for the reaction that has taken place in Test **3**.

Identify the species responsible for the misty white fumes produced in Test **3**.

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(Extra space) .....

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(2)

(ii) The student had read in a textbook that the equation for one of the reactions in Test **4** is as follows.

$$8H^+ + 8I^- + H_2SO_4 \longrightarrow 4I_2 + H_2S + 4H_2O$$

Write the **two** half-equations for this reaction.

State the role of the sulfuric acid and identify the yellow solid that is also observed in Test **4**.

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(Extra space) .....

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(4)

(iii) The student knew that bromine can be used for killing microorganisms in swimming pool water.
The following equilibrium is established when bromine is added to cold water.

$Br_2(l) + H_2O(l) \rightleftharpoons HBrO(aq) + H^+(aq) + Br^-(aq)$

Use Le Chatelier's principle to explain why this equilibrium moves to the right when sodium hydroxide solution is added to a solution containing dissolved bromine.

Deduce why bromine can be used for killing microorganisms in swimming pool water, even though bromine is toxic.

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(Extra space) .....

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(3)
(Total 15 marks)

Q2. A sample of nitrogen dioxide gas (NO_2) was prepared by the reaction of copper with concentrated nitric acid.

(a) (i) Balance the equation for the reaction of copper with concentrated nitric acid.

$$Cu + ... HNO_3 \rightarrow Cu(NO_3)_2 + ... NO_2 + ... H_2O$$

(1)

(ii) Give the oxidation state of nitrogen in each of the following compounds.

HNO_3 .....

NO_2 .....

(2)

(iii) Deduce the half-equation for the conversion of HNO_3 into NO_2 in this reaction.

.....

(1)

(b) The following equilibrium is established between colourless dinitrogen tetraoxide gas (N_2O_4) and dark brown nitrogen dioxide gas.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = 58 \text{ kJ mol}^{-1}$$

(i) Give two features of a reaction at equilibrium.

Feature 1 .....

.....

.....

.....

Feature 2 .....

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(2)

(ii) Use Le Chatelier's principle to explain why the mixture of gases becomes darker in colour when the mixture is heated at constant pressure.

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(2)

(iii) Use Le Chatelier's principle to explain why the amount of NO_2 decreases when the pressure is increased at constant temperature.

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(2)
(Total 10 marks)

Q3. Iodine reacts with concentrated nitric acid to produce nitrogen dioxide (NO_2).

(a) (i) Give the oxidation state of iodine in each of the following.

I_2 .....

HIO_3.....

(2)

(ii) Complete the balancing of the following equation.

$$I_2 + 10HNO_3 \longrightarrow .....HIO_3 + ...NO_2 + ...H_2O$$

(1)

(b) In industry, iodine is produced from the $NaIO_3$ that remains after sodium nitrate has been crystallised from the mineral Chile saltpetre.
The final stage involves the reaction between $NaIO_3$ and NaI in acidic solution.
Half-equations for the redox processes are given below.

$$IO_3^- + 5e^- + 6H^+ \longrightarrow 3H_2O + \frac{1}{2}I_2$$

$$I^- \longrightarrow \frac{1}{2}I_2 + e^-$$

Use these half-equations to deduce an overall ionic equation for the production of iodine by this process. Identify the oxidising agent.

Overall ionic equation

The oxidising agent .....

(2)

(c) When concentrated sulfuric acid is added to potassium iodide, solid sulfur and a black solid are formed.

(i) Identify the black solid.

.....

(1)

(ii) Deduce the half-equation for the formation of sulfur from concentrated sulfuric acid.

.....

(1)

(d) When iodide ions react with concentrated sulfuric acid in a different redox reaction, the oxidation state of sulfur changes from +6 to -2. The reduction product of this reaction is a poisonous gas that has an unpleasant smell.
Identify this gas.

.....

(1)

(e) A yellow precipitate is formed when silver nitrate solution, acidified with dilute nitric acid, is added to an aqueous solution containing iodide ions.

(ii) Write the **simplest ionic** equation for the formation of the yellow precipitate.

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(1)

(ii) State what is observed when concentrated ammonia solution is added to this precipitate.

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(1)

(iii) State why the silver nitrate is acidified when testing for iodide ions.

.....

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(1)

(f) Consider the following reaction in which iodide ions behave as reducing agents.

$$Cl_2(aq) + 2I^-(aq) \longrightarrow I_2(aq) + 2Cl^-(aq)$$

(i) In terms of electrons, state the meaning of the term *reducing agent*.

.....

.....

(1)

(ii) Write a half-equation for the conversion of chlorine into chloride ions.

.....

(1)

(iii) Suggest why iodide ions are stronger reducing agents than chloride ions.

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(Extra space) .....

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(2)

(Total 15 marks)

Q4. Metals are usually extracted from oxides.

Some of these oxides occur naturally. Other oxides are made by roasting sulfide ores in air, producing sulfur dioxide as a by-product.

For the extraction of some metals, the oxide needs to be converted into a chloride.

(a) The ore molybdenite contains molybdenum disulfide (MoS_2).
The first stage in the extraction of molybdenum is to roast the ore in air to form molybdenum oxide (MoO_3) and sulfur dioxide.

(i) Write an equation for the first stage in this extraction.

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(1)

(ii) The release of sulfur dioxide into the atmosphere causes environmental problems and wastes a valuable resource. Identify **one** environmental problem and identify **one** use for the sulfur dioxide.

Environmental problem .....

.....

.....

Use for sulfur dioxide .....

.....

.....

(2)

(iii) Pure molybdenum is formed in the second stage by the reduction of MoO_3 using hydrogen.

Write an equation for this reaction.

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(1)

(iv) State **one** risk in using hydrogen gas in metal extractions.

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(1)

(b) Calcium is an expensive metal. It is extracted by the electrolysis of molten calcium chloride.

(i) State why calcium chloride must be molten for electrolysis to occur.

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.....

(1)

(ii) Write an equation for the reaction that takes place at the negative electrode during this electrolysis.

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(1)

(iii) Identify the major cost in this extraction of calcium.

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(1)

(Total 8 marks)

Q5.(a) Iron is extracted from iron(III) oxide using carbon at a high temperature.

(i) State the type of reaction that iron(III) oxide undergoes in this extraction.

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(1)

(ii) Write a half-equation for the reaction of the iron(III) ions in this extraction.

.....

(1)

(b) At a high temperature, carbon undergoes combustion when it reacts with oxygen.

(i) Suggest why it is **not** possible to measure the enthalpy change directly for the following combustion reaction.

$$C(s,graphite) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$$

.....

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(1)

(ii) State Hess's Law.

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(1)

(iii) State the meaning of the term *standard enthalpy of combustion.*

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(Extra space) .....

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(3)

(c) Use the standard enthalpies of formation in the table below and the equation to calculate a value for the standard enthalpy change for the extraction of iron using carbon monoxide.

	$Fe_2O_3(s)$	CO(g)	Fe(l)	$CO_2(g)$
ΔH_f/ kJ mol^{-1}	- 822	- 111	+14	- 394

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

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(Extra space) .....

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(3)

(d) (i) Write an equation for the reaction that represents the standard enthalpy of formation of carbon dioxide.

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(1)

(ii) State why the value quoted in part (c) for the standard enthalpy of formation of $CO_2(g)$ is the same as the value for the standard enthalpy of combustion of carbon.

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(1)

(Total 12 marks)

Q1. The method of extraction of zinc has changed as different ores containing the element have been discovered and as technology has improved.

Extraction process 1
In the earliest process, calamine (impure zinc carbonate) was heated with charcoal in earthenware pots. This two-stage process gave a low yield of zinc.

$$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$$

$$ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$$

Extraction process 2
Deposits of calamine were being used up and a new two-stage process was developed using zinc sulfide ores. All of the waste gases from this process were released into the atmosphere.

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$$

$$ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$$

Extraction process 3
The modern process uses the electrolysis of aqueous solutions of very pure zinc sulfate. The first step in this process is the same as the first step in Extraction process **2**. The second step uses sulfuric acid made from the SO_2 collected in the first step. The third step involves the electrolysis of zinc sulfate solution to form pure zinc.

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$$

$$ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$$

$$ZnSO_4(aq) \xrightarrow{\text{electrolysis}} Zn(s)$$

(a) In the first stage of Extraction process **1** the following equilibrium is established when zinc carbonate is heated in a closed container.

$$ZnCO_3(s) \rightleftharpoons ZnO(s) + CO_2(g)$$

Use Le Chatelier's principle to suggest and explain the effect on the yield of zinc oxide of allowing the carbon dioxide to escape from the container.

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(3)

(b) State and explain **one** environmental reason why Extraction process **3** is an improvement over Extraction process **2**.

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(3)

(c) Give **one** reason why Extraction process **3** is an expensive method of making zinc but one which is justified in terms of the product formed.

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(2)

(d) Deduce the half-equation for the formation of zinc from zinc ions during the electrolysis of zinc sulfate solution and identify the electrode at which this reaction occurs.

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(2)

(e) Identify **one** reaction from the three extraction processes that is **not** a redox reaction and state the type of reaction that it is. In terms of redox, state what happens to the carbon in Extraction process **2**.

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(3)

(f) Zinc and magnesium both react with steam in a similar way. Write an equation for the reaction of zinc with steam and name the products of this reaction.

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(2)
(Total 15 marks)

Q2. The electrons transferred in redox reactions can be used by electrochemical cells to provide energy.

Some electrode half-equations and their standard electrode potentials are shown in the table below.

Half-equation	$E^{\ominus}$**/V**
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77

$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	–0.44
$Li^+(aq) + e^- \rightarrow Li(s)$	–3.04

(a) Describe a standard hydrogen electrode.

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(4)

(b) A conventional representation of a lithium cell is given below.
This cell has an e.m.f. of +2.91 V

$$Li(s) \mid Li^+(aq) \parallel Li^+(aq) \mid MnO_2(s) , LiMnO_2(s) \mid Pt(s)$$

Write a half-equation for the reaction that occurs at the positive electrode of this cell.

Calculate the standard electrode potential of this positive electrode.

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(2)

(c) Suggest what reactions occur, if any, when hydrogen gas is bubbled into a solution containing a mixture of iron(II) and iron(III) ions. Explain your answer.

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(2)

(d) A solution of iron(II) sulfate was prepared by dissolving 10.00 g of $FeSO_4.7H_2O$ (M_r = 277.9) in water and making up to 250 cm^3 of solution. The solution was left to stand, exposed to air, and some of the iron(II) ions became oxidised to iron(III) ions. A 25.0 cm^3 sample of the partially oxidised solution required 23.70 cm^3 of 0.0100 mol dm^{-3} potassium dichromate(VI) solution for complete reaction in the presence of an excess of dilute sulfuric acid.

Calculate the percentage of iron(II) ions that had been oxidised by the air.

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(6)
(Total 14 marks)

Q3.(a) When a solution containing iron(II) ions is treated with a slight excess of a solution containing ethanedioate ions a bright yellow precipitate of hydrated iron(II)

ethanedioate, $FeC_2O_4.2H_2O$, is formed. The precipitate is filtered off, washed with propanone and then allowed to dry. A typical yield of the solid is 95%.

(i) Propanone boils at 56 °C and is miscible with water in all proportions. Suggest **two** reasons why washing with propanone is an effective method for producing a pure, dry precipitate.

Reason 1 .....

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Reason 2 .....

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(2)

(ii) By suggesting a simple test tube reaction, state how the filtrate could be tested to show that all of the iron(II) ions have been removed from the solution. State what you would observe.

Test .....

Observation .....

(2)

(iii) Suggest **one** reason why the typical yield of iron(II) ethanedioate is less than 100%.

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(1)

(iv) Calculate the mass of hydrated iron(II) ethanedioate, $FeC_2O_4.2H_2O$ that can be formed from 50.0 cm^3 of a 0.50 mol dm^{-3} solution of iron(II) sulfate when the yield of the reaction is 95%. Show your working.

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(3)

(v) The identity of the precipitate can be confirmed by dissolving it in sulfuric acid and titrating the mixture with potassium manganate(VII).

Deduce the number of moles of iron(II) ethanedioate that would react with one mole of potassium manganate(VII) in acidic solution.

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(1)

(b) Ethanedioate ions can be used to remove calcium ions from blood plasma. A precipitate of calcium ethanedioate is formed. Write an ionic equation for the reaction of ethanedioate ions with calcium ions.

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(1)

(c) Ethanedioic acid is used to clean marble, a form of calcium carbonate. Suggest **one** reason why the reaction between ethanedioic acid and marble stops after a short time.

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(1)

(d) Tea leaves contain ethanedioic acid. Suggest **one** reason why tea drinkers do **not** suffer from ethanedioic acid poisoning.

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(1)

(e) Ethanedioic acid is produced by the oxidation of carbon monoxide in a multi-step process. The equation which summarises the reactions taking place is shown below.

$$4CO + 4NaOH + O_2 + 4HCl \rightarrow 2H_2C_2O_4 + 4NaCl + 2H_2O$$

Calculate the percentage atom economy for the formation of ethanedioic acid in this reaction. Show your working.

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(2)
(Total 14 marks)

Q4. Nitric acid is manufactured from ammonia in a process that involves several stages.

(a) In the first stage, ammonia is converted into nitrogen monoxide and the following equilibrium is established.

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \qquad \Delta H = -905 \text{ kJ mol}^{-1}$$

The catalyst for this equilibrium reaction is a platinum–rhodium alloy in the form of a gauze. This catalyst gauze is heated initially but then remains hot during the reaction.

(i) In terms of redox, state what happens to the ammonia in the forward reaction.

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(1)

(ii) Suggest a reason why the catalyst must be hot.

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(1)

(iii) Suggest a reason why the catalyst remains hot during the reaction.

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(1)

(iv) State how a catalyst increases the rate of a reaction.

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(2)

(b) In the second stage, nitrogen monoxide is converted into nitrogen dioxide. The equation for the equilibrium that is established is shown below.

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = -113 \text{ kJ mol}^{-1}$$

Explain why the equilibrium mixture is cooled during this stage of the process.

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(2)

(c) In the final stage, nitrogen dioxide reacts with water as shown by the following equation.

$$2NO_2(g) + H_2O(l) \rightarrow H^+(aq) + NO_3^-(aq) + HNO_2(aq)$$

Give the oxidation state of nitrogen in each of the following.

NO_2 .....

NO_3^- .....

HNO_2 .....

(3)
(Total 10 marks)

Q1. Which one of the following is the electron arrangement of the strongest reducing agent?

A $1s^2\ 2s^2\ 2p^5$

B $1s^2\ 2s^2\ 2p^6\ 3s^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

(Total 1 mark)

Q2. Which one of the following is **not** a redox reaction?

A $Br_2 + SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2Br^-$

B $SnCl_2 + HgCl_2 \rightarrow Hg + SnCl_4$

C $Cu_2O + H_2SO_4 \rightarrow CuSO_4 + Cu + H_2O$

D $2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$

(Total 1 mark)

Q3. Hydrogen gas is used in the chemical industry.

(a) Tungsten is extracted by passing hydrogen over heated tungsten oxide (WO_3).

(i) State the role of the hydrogen in this reaction.

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(1)

(ii) Write an equation for this reaction.

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(1)

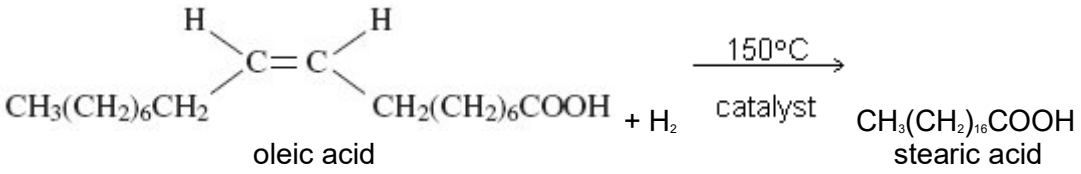
(iii) State **one** risk of using hydrogen gas in metal extractions.

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(1)

(b) Hydrogen is used to convert oleic acid into stearic acid as shown by the following equation.



oleic acid stearic acid

(i) Use your knowledge of the chemistry of alkenes to deduce the type of reaction that has occurred in this conversion.

.....

(1)

(ii) State the type of stereoisomerism shown by oleic acid.

.....

(1)

(c) Hydrogen reacts with nitrogen in the Haber Process. The equation for the equilibrium that is established is shown below.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(i) State Le Chatelier's principle.

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(1)

(ii) Use Le Chatelier's principle to explain why an increase in the total pressure of this equilibrium results in an increase in the equilibrium yield of ammonia.

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(2)

(d) Hydrogen reacts with oxygen in an exothermic reaction as shown by the following equation.

$$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g) \qquad \Delta H = -242 \text{ kJ mol}^{-1}$$

Use the information in the equation and the data in the following table to calculate a value for the bond enthalpy of the H–H bond.

	O–H	O=O
Mean bond enthalpy / kJ mol^{-1}	+ 463	+ 496

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(3)

(Total 11 marks)

Q4. Copper is extracted from the ore chalcopyrite ($CuFeS_2$) in a three-stage process.

(a) In the first stage of this extraction, the chalcopyrite is heated with silicon dioxide and oxygen.

(i) Balance the following equation for this first stage in which copper(I) sulfide is formed.

.....$CuFeS_2$ + ...SiO_2 + ...O_2 → Cu_2S + ...$FeSiO_3$ + ...SO_2

(1)

(ii) Give **one** environmental reason why the SO_2 gas formed in this reaction is not allowed to escape into the atmosphere.

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.....

(1)

(iii) State **one** use for the sulfur dioxide formed in this reaction.

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(1)

(b) In the second stage of this extraction, the copper(I) sulfide is converted into copper(II) oxide. This occurs by roasting the sulfide with oxygen at high temperature.
Write an equation for this reaction.

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(1)

(c) In the third stage of this extraction, copper(II) oxide is reduced to copper by its reaction with carbon. Write an equation for this reaction.

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(1)

(d) Scrap iron can be used to extract copper from dilute aqueous solutions containing copper(II) ions.

(i) Explain why this is a low-cost method of extracting copper.

.....

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(1)

(ii) Write the **simplest ionic** equation for the reaction of iron with copper(II) ions in aqueous solution.

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(1)
(Total 7 marks)

Q5. Oxidation and reduction can de defined in terms of electron transfer.

(a) Define the term *reduction* in terms of electrons.

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(1)

(b) The oxide of nitrogen formed when copper reacts with nitric acid depends upon the concentration and the temperature of the acid. The reaction of copper with cold, dilute acid produces NO as indicated by the following equation.

$$3Cu + 8H^+ + 2NO_3^- \rightarrow 3Cu^{2+} + 4H_2O + 2NO$$

In warm, concentrated acid, NO_2 is formed.

Oxidation states can be used to understand electron transfer in these reactions.

(i) Give the oxidation states of nitrogen in NO_3^-, NO and NO

Oxidation state in NO_3^- .....

Oxidation state in NO_2 .....

Oxidation state in NO .....

(ii) Identify, as oxidation or reduction, the formation of NO_2 from NO_3^- ions in the presence of H^+ ions. Deduce the half-equation for the reaction.

NO from NO_3^- .....

Half-equation .;.....

(iii) Deduce the half-equation for the formation of NO_2 from NO_3^- ions in the presence of H+ ions.

.....

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(iv) Deduce the overall equation for the reaction of copper with NO_3^- ions and H^+ ions to produce Cu^{2+} ions, NO_2 and water.

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(8)
(Total 9 marks)

Q6. This question is about the extraction of metals.

(a) Coke is mainly carbon and is a raw material used in the extraction of iron from iron(III) oxide.

(i) Write an equation for the formation of carbon monoxide from carbon.

.....

(1)

(ii) Write an equation for the reduction of iron(III) oxide to iron by carbon monoxide.

.....

(1)

(iii) The Earth's resources of iron(III) oxide are very large and commercial ores have a high iron content. Give **one** economic and **one** environmental reason for recycling scrap iron and steel.

Economic reason .....

.....

Environmental reason .....

.....

(2)

(b) Pure titanium is extracted by the reduction of titanium(IV) chloride, but not by the direct reduction of titanium(IV) oxide using carbon.

(i) Write an equation for the conversion of titanium(IV) oxide into titanium(IV) chloride.

.....

(2)

(ii) Write an equation for the extraction of titanium from titanium(IV) chloride.

.....

(2)

(iii) State why titanium is not extracted directly from titanium(IV) oxide using carbon.

.....

(1)

(c) Aluminium is extracted by the electrolysis of a molten mixture containing aluminium oxide.

(i) State why the electrolysis needs to be of a *molten* mixture.

.....

(1)

(ii) Write an equation for the reaction of oxide ions at the positive electrode during the electrolysis.

.....

(1)

(iii) State why the positive electrodes need frequent replacement.

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(1)

(iv) Give the major reason why it is less expensive to recycle aluminium than to extract it from aluminium oxide by electrolysis.

.....

(1)

(Total 13 marks)

Q7. In the past 150 years, three different processes have been used to extract bromine from potassium bromide. These processes are illustrated below.

Extraction Process **1**

$2KBr + MnO_2 + 2H_2SO_4 \rightarrow MnSO_4 + K_2SO_4 + 2H_2O + Br_2$

Extraction Process **2**

The reaction of solid potassium bromide with concentrated sulfuric acid.

Extraction Process **3**

The reaction of aqueous potassium bromide with chlorine gas.

(a) Write a half-equation for the conversion of MnO_2 in acid solution into Mn^{2+} ions and water. In terms of electrons, state what is meant by the term *oxidising agent* and identify the oxidising agent in the overall reaction.

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(3)

(b) Write an equation for Extraction Process **2** and an equation for Extraction Process **3**.
Calculate the percentage atom economy for the extraction of bromine from potassium bromide by Extraction Process **3**. Suggest why Extraction Process **3** is the method in large-scale use today.

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(5)

(c) Bromine has been used for more than 70 years to treat the water in swimming pools.
The following equilibrium is established when bromine is added to water.

$$Br_2 + H_2O \rightleftharpoons HBrO + HBr$$

Give the oxidation state of bromine in HBr and in HBrO

Deduce what will happen to this equilibrium as the HBrO reacts with micro-organisms in the swimming pool water. Explain your answer.

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(4)
(Total 12 marks)

Q1. In which one of the following reactions is the role of the reagent stated correctly?

	Reaction	**Role of reagent**
A	$TiO_2 + 2C + 2Cl_2 \rightarrow TiCl_4 + 2CO$	TiO_2 is an oxidising agent
B	$HNO_3 + H_2SO_4 \rightarrow H_2NO_3^+ + HSO_4^-$	HNO_3 is a Brønsted-Lowry acid
C	$CH_3COCl + AlCl_3 \rightarrow CH_3CO^+ + AlCl_4^-$	$AlCl_3$ is a Lewis base
D	$2CO + 2NO \rightarrow 2CO_2 + N_2$	CO is a reducing agent

(Total 1 mark)

Q2. Which one of the following statements is **not** correct?

A The first ionisation energy of iron is greater than its second ionisation energy.

B The magnitude of the lattice enthalpy of magnesium oxide is greater than that of barium oxide.

C The oxidation state of iron in $[Fe(CN)_6]^{3-}$ is greater than the oxidation state of copper in $[CuCl_2]^-$

D The boiling point of C_3H_8 is lower than that of CH_3CH_2OH

(Total 1 mark)

Q3. In which one of the following reactions does hydrogen **not** act as a reducing agent?

A $H_2 + Ca \rightarrow CaH_2$

B $2H_2 + O_2 \rightarrow 2H_2O$

C $H_2 + CH_2{=}CH_2 \rightarrow CH_3CH_3$

C $2H_2 + CH_3COCH_3 \rightarrow CH_3CH_2CH_3 + H_2O$

(Total 1 mark)

Q4. Chlorine and bromine are both oxidising agents.

(a) Define an *oxidising agent* in terms of electrons.

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(1)

(b) In aqueous solution, bromine oxidises sulphur dioxide, SO_2, to sulphate ions, SO_4^{2-}

(i) Deduce the oxidation state of sulphur in SO_2 and in SO_4^{2-}

SO_2 .....

SO_4^{2-} .....

(ii) Deduce a half-equation for the reduction of bromine in aqueous solution.

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(iii) Deduce a half-equation for the oxidation of SO_2 in aqueous solution forming SO_4^{2-} and H^+ ions.

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(iv) Use these two half-equations to construct an overall equation for the reaction between aqueous bromine and sulphur dioxide.

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(5)

(c) Write an equation for the reaction of chlorine with water. Below each of the chlorine-containing products in your equation, write the oxidation state of chlorine in that product.

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(3)

(d) Give a reason why chlorine is not formed when solid potassium chloride reacts with concentrated sulphuric acid.

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(1)

(e) Write an equation for the reaction between solid potassium chloride and concentrated sulphuric acid.

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(1)

(f) Solid potassium bromide undergoes a redox reaction with concentrated sulphuric acid.

(i) Give the oxidation product formed from potassium bromide.

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(ii) Give the reduction product formed from sulphuric acid.

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(2)

(Total 13 marks)

Q5. The vanadium does **not** have an oxidation state of +3 in

A $[V(H_2O)_6]^{3+}$

B $[V(C_2O_4)_3]^{3-}$

C $[V(OH)_3(H_2O)_3]$

D $[VCl_4]^{3-}$

(Total 1 mark)

Q6. (a) The term oxidation was used originally to describe a reaction in which a substance gained oxygen. The oxygen was provided by the oxidising agent. Later the definition of oxidation was revised when the importance of electron transfer was recognised.

An aqueous solution of sulfur dioxide was reacted in separate experiments as follows.

Reaction **1** with HgO

$$H_2O + SO_2 + HgO \rightarrow H_2SO_4 + Hg$$

Reaction **2** with chlorine

$$2H_2O + SO_2 + Cl_2 \rightarrow H_2SO_4 + 2HCl$$

(i) In Reaction **1**, identify the substance that donates oxygen and therefore is the oxidising agent.

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(ii) Show, by writing a half-equation, that this oxidising agent in reaction **1** is an electron acceptor.

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(iii) Write a half-equation for the oxidation process occurring in reaction **2**.

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(iv) Write a half-equation for the reduction process occurring in reaction **2**.

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(4)

(b) Use the standard electrode potential data given in the able below to answer the questions which follow.

		E / V
$V^{3+}(aq) + e^-$	$\rightarrow V^{2+}(aq)$	–0.26
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$	$\rightarrow H_2SO_3(aq) + H_2O(l)$	+0.17
$VO^{2+}(aq) + 2H^+(aq) + e^-$	$\rightarrow V^{3+}(aq) + H_2O(l)$	+0.34
$Fe^{3+}(aq) + e^-$	$\rightarrow Fe^{2+}(aq)$	+0.77
$VO_2^+(aq) + 2H^+(aq) + e^-$	$\rightarrow VO^{2+}(aq) + H_2O(l)$	+1.00
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions

(i) An excess of potassium manganate(VII) was added to a solution containing $V^{2+}(aq)$ ions. Determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from $V^{2+}(aq)$.

Vanadium species present at the end of the reaction .....

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Oxidation state of vanadium in the final species .....

.....

Half-equation .....

(ii) The cell represented below was set up under standard conditions.

$$Pt|H_2SO_3(aq),SO_4^{2-}(aq),H^+(aq)||Fe^{3+}(aq),Fe^{2+}(aq)|Pt$$

Calculate the e.m.f. of this cell and state, with an explanation, how this e.m.f. will change if the concentration of $Fe^{3+}(aq)$ ions is increased.

Cell e.m.f. .....

Change in cell e.m.f. .....

Explanation .....

.....

(7)

(c) Consider the cell below

$$\overset{-}{Pt|H_2(g)|H^+(aq)||O_2(g)|OH^-(aq)|Pt}\overset{+}{}$$

(i) Using half-equations, deduce an overall equation for the cell reaction.

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(ii) State how, if at all, the e.m.f. of this cell will change if the surface area of each platinum electrode is doubled.

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(3)

(d) Currently, almost all hydrogen is produced by the high-temperature reaction between methane, from North Sea gas, and steam. Give one economic and one environmental disadvantage of this method of producing hydrogen.

Economic disadvantage .....

Environmental disadvantage .....

(2)

(e) Hydrogen can also be produced by the electrolysis of acidified water using electricity produced using solar cells. Give one reason why this method is not used on a large scale.

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(1)

(Total 17 marks)

Q7. (a) By referring to electrons, explain the meaning of the term *oxidising agent*.

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(1)

(b) For the element **X** in the ionic compound **MX**, explain the meaning of the term *oxidation state*.

.....

(1)

(c) Complete the table below by deducing the oxidation state of each of the stated elements in the given ion or compound.

	Oxidation state
Carbon in CO_3^{2-}	
Phosphorus in PCl_4^{+}	
Nitrogen in Mg_3N_2	

(3)

(d) In acidified aqueous solution, nitrate ions, NO_3^-, react with copper metal forming nitrogen monoxide, NO, and copper(II) ions.

(i) Write a half-equation for the oxidation of copper to copper(II) ions.

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(ii) Write a half-equation for the reduction, in an acidified solution, of nitrate ions to nitrogen monoxide.

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(iii) Write an overall equation for this reaction.

.....

(3)
(Total 8 marks)

Q8. Photochromic glass contains silver ions and copper ions. A simplified version of a redox equilibrium is shown below. In bright sunlight the high energy u.v. light causes silver atoms to form and the glass darkens. When the intensity of the light is reduced the reaction is reversed and the glass lightens.

$$Cu^+(s) + Ag^+(s) \rightleftharpoons Cu^{2+}(s) + Ag(s)$$

clear glass dark glass

When the photochromic glass darkens

A the Ag^+ ion is acting as an electron donor.

B the Cu^+ ion is acting as a reducing agent.

C the Ag^+ ion is oxidised.

D the Cu^+ ion is reduced.

(Total 1 mark)

Q9. Refer to the following reaction

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad \Delta H^{\ominus} = -11 \text{ kJ mol}^{-1}, \qquad \Delta S^{\ominus} = +20 \text{ J K}^{-1} \text{ mol}^{-1}$$

Which one of the following statements is correct?

A This is a redox reaction.

B The reaction is **not** feasible below 298 K

C At equilibrium, the yield of hydrogen iodide is changed by increasing the pressure.

D At equilibrium, the yield of hydrogen iodide increases as the temperature is increased.

(Total 1 mark)

Q10. At high temperatures, nitrogen is oxidised by oxygen to form nitrogen monoxide in a reversible reaction as shown in the equation below.

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad \Delta H^{\ominus} = +180 \text{ kJ mol}^{-1}$$

(a) In terms of electrons, give the meaning of the term *oxidation*.

.....

(1)

(b) State and explain the effect of an increase in pressure, and the effect of an increase in temperature, on the yield of nitrogen monoxide in the above equilibrium.

Effect of an increase in pressure on the yield .....

Explanation .....

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Effect of an increase in temperature on the yield .....

Explanation .....

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(6)

(c) Nitrogen monoxide, NO, is formed when silver metal reduces nitrate ions, NO_3^- in acid solution.

(i) Deduce the oxidation state of nitrogen in NO and in NO_3^-

NO.....

NO_3^- .....

(ii) Write a half-equation for the reduction of NO_3^- ions in acid solution to form nitrogen monoxide and water.

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(iii) Write a half-equation for the oxidation of silver metal to $Ag^+(aq)$ ions.

.....

(iv) Hence, deduce an overall equation for the reaction between silver metal and nitrate ions in acid solution.

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(5)

(Total 12 marks)

Q1.(a) The following is an equation for a redox reaction.

$2NO + 12H^+ + 10I^- \rightarrow 2NH_4^+ + 2H_2O + 5I_2$

(i) Define *oxidation* in terms of electrons.

.....

(ii) Deduce the oxidation state of nitrogen in NO and of nitrogen in NH_4^+

Oxidation state of nitrogen in NO .....

Oxidation state of nitrogen in NH_4^+ .....

(iii) Identify the species formed by oxidation in this reaction.....

(4)

(b) When chlorine gas is bubbled into an aqueous solution of sulphur dioxide, hydrogen ions, sulphate ions and chloride ions are formed.

(i) Write a half-equation for the formation of chloride ions from chlorine.

.....

(ii) Write a half-equation for the formation of hydrogen ions and sulphate ions from sulphur dioxide and water.

.....

(iii) Hence, deduce an overall equation for the reaction which occurs when chlorine is bubbled into aqueous sulphur dioxide.

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(3)
(Total 7 marks)

Q2.

Summarised directions for recording responses to multiple completion questions			
A (i), (ii) and (iii) only	**B** (i) and (iii) only	**C** (ii) and (iv) only	**D** (iv) alone

In which of the following conversions is the copper reduced?

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

(ii) $[Cu(H_2O)_6]^{2+} \rightarrow Cu(H_2O)_4(OH)_2$

(iii) $Cu \rightarrow CuCl_2$

(iv) $[Cu(H_2O)_6]^{2+} \rightarrow CuCl$

(Total 1 mark)

Q3. Which one of the following contains the metal with the lowest oxidation state?

A CrO_2F_2

B $[Cr_2O_7]^{2}$

C $[MnCl_6]^{2-}$

D $[Mn(CN)_6]^{3-}$

(Total 1 mark)

Q4. (a) Samples of solid sodium fluoride, sodium chloride, sodium bromide and sodium iodide are each warmed separately with concentrated sulphuric acid. All four compounds react with concentrated sulphuric acid but only two can reduce it.

(i) Identify the **two** halides which do **not** reduce concentrated sulphuric acid. Write an equation for the reaction which does occur with **one** of these two halides.

(ii) Identify the **two** halides which reduce concentrated sulphuric acid to sulphur dioxide. Using half-equations for the oxidation and reduction processes, deduce an overall equation for the formation of sulphur dioxide when concentrated sulphuric acid reacts with **one** of these halides.

(iii) In addition to sulphur dioxide, two further reduction products are formed when one of these two halides reacts with concentrated sulphuric acid. Identify the two reduction products and write a half-equation to show the formation of **one** of them from concentrated sulphuric acid.

(9)

(b) How would you distinguish between separate solutions of sodium chloride, sodium bromide and sodium iodide using solutions of silver nitrate and ammonia?

(6)
(Total 15 marks)

Q5. In which one of the following reactions does the metal species undergo reduction?

A $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

B $[Cu(H_2O)_6]^{2+} + 4Cl^- \rightarrow [CuCl_4]^{2-} + 6H_2O$

C $CrO_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$

D $TiO_2 + 2C + 2Cl_2 \rightarrow TiCl_4 + 2CO$

(Total 1 mark)

Q6. The compound lithium tetrahydridoaluminate(III), $LiAlH_4$, is a useful reducing agent. It behaves in a similar fashion to $NaBH_4$. Carbonyl compounds and carboxylic acids are reduced to alcohols. However, $LiAlH_4$ also reduces water in a violent reaction so that it must be used in an organic solvent.

Which one of the following concerning the violent reaction between $LiAlH_4$ and water is **false**?

A A gas is produced.

B The activation energy for the reaction is relatively high.

C The reaction has a negative free-energy change.

D Aqueous lithium ions are formed.

(Total 1 mark)

Q7. Which one of the following is the electronic configuration of the strongest reducing agent?

A $1s^2\ 2s^2\ 2p^5$

B $1s^2\ 2s^2\ 2p^6\ 3s^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

(Total 1 mark)

Q8. Which one of the following is a redox reaction?

A $2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$

B $3Cl_2 + 6OH^- \rightarrow 5Cl^- + ClO_3^- + 3H_2O$

C $HNO_3 + 2H_2SO_4 \rightarrow NO_2^+ + H_3O^+ + 2HSO_4^-$

D $CaCO_3 + SiO_2 \rightarrow CaSiO_3 + CO_2$

(Total 1 mark)

Q9. (a) In terms of electron transfer, what does the reducing agent do in a redox reaction?

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(1)

(b) What is the oxidation state of an atom in an uncombined element?

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(1)

(c) Deduce the oxidation state of nitrogen in each of the following compounds.

(i) NCl_3 .....

(ii) Mg_3N_2 .....

(iii) NH_2OH .....

(3)

(d) Lead(IV) oxide, PbO_2, reacts with concentrated hydrochloric acid to produce chlorine, lead(II) ions, Pb^{2+}, and water.

(i) Write a half-equation for the formation of Pb^{2+} and water from PbO_2 in the presence of H^+ ions.

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(ii) Write a half-equation for the formation of chlorine from chloride ions.

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(iii) Hence deduce an equation for the reaction which occurs when concentrated hydrochloric acid is added to lead(IV) oxide, PbO_2

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(3)
(Total 8 marks)

Q10. (a) Concentrated sulphuric acid can be reduced by some solid sodium halides to H_2S

(i) Give the oxidation state of sulphur in H_2S

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(ii) Give **one** solid sodium halide which will reduce concentrated sulphuric acid, forming H_2S

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(iii) State **one** way in which the presence of H_2S could be recognised.

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(iv) Write a half-equation for the formation of H_2S from sulphuric acid.

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(4)

(b) A different solid sodium halide reacts with concentrated sulphuric acid without reduction forming a halogen-containing product **X**.

(i) Suggest an identity for **X**.

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(ii) Identify the solid sodium halide which produces **X**.

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(iii) State the role of sulphuric acid in the formation of **X**.

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(iv) Write an equation for the reaction with concentrated sulphuric acid in which **X**

is formed.

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(4)
(Total 8 marks)

Q11. (a) In terms of electrons, what happens to an oxidising agent during a redox reaction?

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(1)

(b) Consider the following redox reaction.

$SO_2(aq) + 2H_2O(l) + 2Ag^+(aq) \rightarrow 2Ag(s) + SO_4^{2-}(aq) + 4H^+(aq)$

(i) Identify the oxidising agent and the reducing agent in this reaction.

Oxidising agent .....

Reducing agent .....

(ii) Write a half-equation to show how sulphur dioxide is converted into sulphate ions in aqueous solution.

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(3)

(c) Fe^{2+} ions are oxidised to Fe^{3+} ions by ClO_3^- ions in acidic conditions. The ClO_3^- ions are reduced to Cl^- ions.

(i) Write a half-equation for the oxidation of Fe^{2+} ions in this reaction.

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(ii) Deduce the oxidation state of chlorine in ClO_3^- ions.

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(iii) Write a half-equation for the reduction of ClO_3^- ions to Cl^- ions in acidic conditions.

.....

(iv) Hence, write an overall equation for the reaction.

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(4)

(d) Write an equation to show how sulphur is removed from impure iron obtained from the Blast Furnace. Identify the oxidising agent in this reaction.

Equation .....

Oxidising agent .....

(2)

(Total 10 marks)

Q12. Which equation does **not** involve the reduction of a transition metal compound?

A $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

B $TiO_2 + 2C + 2Cl_2 \rightarrow TiCl_4 + 2CO$

C $Cr_2O_3 + 2Al \rightarrow 2Cr + Al_2O_3$

D $TiCl_4 + 4Na \rightarrow Ti + 4NaCl$

(Total 1 mark)

Q13. In which one of the following reactions do two H ions and one electron have to be added to the left-hand side in order to balance the equation?

A $CH_3CHO \rightarrow CH_3CH_2OH$

B $VO^{2+} \rightarrow V^{3+} + H_2O$

C $NO_3^- \rightarrow HNO_2 + H_2O$

D $HOCl \rightarrow \frac{1}{2}Cl_2 + H_2O$

(Total 1 mark)

Q14. In which one of the following reactions is H_2O_2 behaving as a reducing agent?

A $H_2O_2 + 2I^- + 2H^+ \rightarrow I_2 + 2H_2O$

B $H_2O_2 + 2[Co(NH_3)_6]^{2+} \rightarrow 2[Co(NH_3)_6]^{3+} + 2OH^-$

C $5H_2O_2 + -2MnO_4^- + 6H+ \rightarrow 2Mn^{2+} + 8H_2O + 5O_2$

D $3H_2O_2 + 2[Cr(OH)_6]^{3-} \rightarrow 2CrO_4^{2-} + 8H_2O + 2OH^-$

(Total 1 mark)

Q15. Which one of the following is **not** a redox reaction?

A $TiO_2 + 2Cl_2 + C \rightarrow TiCl_4 + CO_2$

B $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

C $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

D $3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$

(Total 1 mark)