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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(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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(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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(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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(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.

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

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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(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.

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(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 ....

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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.

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

(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.

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

$$I^- \longrightarrow \frac{1}{2}I_2 \quad + \quad e^-$$
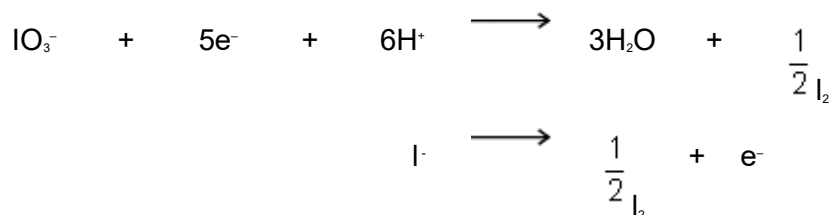
Use these half-equations to deduce an overall ionic equation for the production of iodine by this process. Identify the oxidising agent.

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.

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(1)

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

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(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.

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(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.

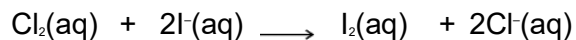...
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(1)

(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*.

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(1)

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

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(1)

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

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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 ....

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Use for sulfur dioxide ....

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(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.

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(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.

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

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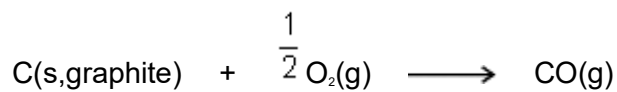
(1)

(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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(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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(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)