9.1 Redox Processes

Question Paper

Course	DP IB Chemistry
Section	9. Redox Processes
Topic	9.1 Redox Processes
Difficulty	Medium

Time allowed: 90

Score: /73

Percentage: /100

Question 1a

a) Common household bleach is a cleaning product which smells like chlorine gas and is therefore, also called chlorine bleach.
It contains a mixture of sodium chlorate ($NaOCl$), sodium chloride and water and can be made by dissolving chlorine gas in a solution of sodium hydroxide.

i) Write a balanced equation with state symbols for this reaction.

ii) Deduce the oxidation number of chlorine in all of the chlorine-containing reactants and products.

[3 marks]

Question 1b

b) The mixing of household bleach with ammonia during cleaning should be avoided, as a redox reaction between the ammonia and the chlorate(I) ions in bleach will generate toxic chlorine gas and hydrazine (N_2H_4).

The overall redox reaction for this reaction is shown below.

$$2NH_3\,(aq) + 2ClO^-\,(aq) \rightarrow N_2H_4\,(aq) + Cl_2\,(g) + 2OH^-\,(aq)$$

i) What are the oxidation numbers of the nitrogen atom in NH_3 and in N_2H_4?

ii) What is the oxidizing agent in this reaction? Explain your answer.

iii) Why is the hazard of the toxic chlorine gas being produced greater than the hazard of hydrazine?

[4 marks]

Question 1c

c) Due to the risks associated with chlorine-based bleach, alternative bleaches are often used instead. These bleaches are based on peroxides such as hydrogen peroxide.

Manganate(VII) ions oxidize hydrogen peroxide to oxygen gas. The reaction is carried out with both species under acidic conditions.

i) Identify the oxidizing and reducing agents in this reaction.

ii) Write the half-equation for the oxidation of hydrogen peroxide to oxygen gas.

iii) The manganate(VII) ions themselves get reduced to manganese(II) ions. Write down the half-equation for the reduction of manganate(VII) ions.

iv) Deduce the overall redox equation for this reaction.

[5 marks]

Question 1d

d) Explain how the oxidation number of the oxygen atom in H_2O_2 is different from its oxidation state in other compounds.

[2 marks]

Question 2a

a) Metals can often be seen written as a list, from the most reactive metal to the least reactive metal. This list is known as the reactivity series of metals and can be used to predict the feasibility of a reaction.

Below is a section of the reactivity series of metals, ordered from most to least reactive:

Calcium
Magnesium
Aluminium
Zinc
Iron
Tin
Lead

A piece of zinc was placed into a solution of iron(II) sulfate and a solution of magnesium sulfate.

Predict, giving a reason, whether a reaction would occur in each solution.

[2 marks]

Question 2b

b) Copper is below lead on the reactivity series shown in part (a). A piece of zinc was placed into a solution of copper(II) sulfate.

Write the half equation for the zinc and identify the type of reaction taking place.

[2 marks]

Question 2c

c) Many chemical reactions are redox reactions as they involve the transfer of electrons.

i) Explain the role of the oxidizing agent in a redox reaction in terms of electron transfer.

ii) State the most common oxidation number of an oxygen atom when in a compound.

iii) Which oxygen compounds are an exception to your answer in part (ii)? Explain your answer.

[4 marks]

Question 2d

d) The following reaction is an example of a common redox reaction:

$$5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

Deduce the oxidation numbers of iron and manganese in the above reaction, both as reactants and as products.

State which substance is reduced.

[3 marks]

Question 2e

e) The amount of iron in some dietary iron supplements was analyzed by redox titration. Four tablets were crushed and dissolved in 50.0 cm^3 of 2.00 mol dm^{-3} sulfuric acid. The solution was then transferred to a 250 cm^3 volumetric flask and made up to 250 cm^3 with distilled water.

A 25.0 cm^3 sample of the iron tablets solution was titrated against 0.00500 mol dm^{-3} potassium manganate(VII) and 25.8 cm^3 was needed for complete reaction.

Determine the amount of iron, in mol, in **one** tablet.

[4 marks]

Question 3a

a) Halide ions, such as chloride, Cl^-, can be identified using chemical tests. If an unknown compound is dissolved in dilute nitric acid, and then silver nitrate solution is added, a precipitate will form if the unknown solution contains halide ions. The precipitate formed will be a silver halide.

The general equation for the precipitation reaction of halide ions with silver nitrate solution is:

$$AgNO_3\,(aq) + X^-\,(aq) \rightarrow AgX\,(s) + NO_3^-\,(aq)$$

i) Deduce the oxidation number of silver in $AgNO_3$ and AgX and deduce the oxidation number of the halide in X^- and in AgX.

ii) Is the precipitation of silver halides a redox reaction? Explain your answer.

[4 marks]

Question 3b

b) Halide ions can also react with other halogens in aqueous solutions. Chlorine reacts in a redox reaction with an aqueous solution of sodium bromide, to form sodium chloride and bromine.

$$Cl_2\,(aq) + NaBr\,(aq) \rightarrow NaCl\,(aq) + Br_2\,(aq)$$

i) State what type of redox reaction this is.

ii) Using the overall redox reaction above, deduce the half-equation for chlorine. State whether chlorine is oxidized or reduced.

iii) Using the overall redox reaction above, deduce the half-equation for bromine. State whether bromine is oxidized or reduced.

iv) Use the reaction above and your knowledge of the halogens, to explain whether chlorine or bromine is a stronger oxidizing agent.

[7 marks]

Question 3c

c) Chlorine also oxidizes sulfur dioxide (SO_2) in aqueous solutions to sulfate ions (SO_4^{2-}) under acidic conditions.

i) Deduce the half-equation for the reduction of chlorine in aqueous solution.

ii) Deduce the half-equation for the oxidation of sulfur dioxide in aqueous solution.

[2 marks]

Question 3d

d) Use the two half-equations from part (c) to construct the overall redox equation for this reaction.

[1 mark]

Question 4a

a) The iron of railway lines rusts when it comes into contact with water and oxygen. The overall redox equation for the rusting of iron is as follows:

$$4Fe\ (s) + 3O_2\ (g) + 6H_2O\ (g) \rightarrow 4Fe(OH)_3\ (s)$$

Define the term *reduction*.

[1 mark]

Question 4b

b) State, with a reason, the oxidizing agent in this reaction in part (a).

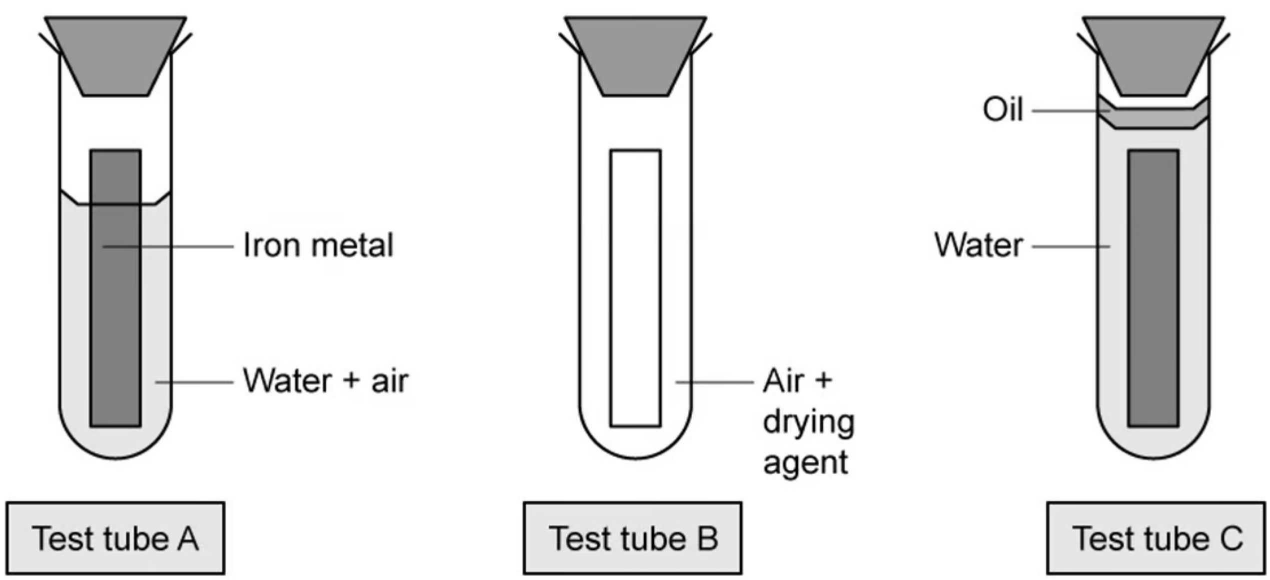
[2 marks]

Question 4c

c) A student investigates the rate of rusting of a piece of iron under different conditions.

Figure 1 shows the set-up of the students' experiment.

Figure 1



Predict in which test tube(s) the iron metal will not rust. Explain your answer.

[3 marks]

Question 4d

d) Deduce the oxidation number of each of the stated elements in the ions and compounds to complete **Table 1** below.

Table 1

Species	**Oxidation number**
Oxygen in Na_2O_2	
Hydrogen in MgH_2	
Nitrogen in NO_3^-	
Chlorine in *Cl*F	

[4 marks]

Question 5a

a) Aluminium is present in the Earth's crust in aluminium ore, called bauxite. A number of processes are done to this ore, to extract the aluminium from it. The bauxite is initially purified to produce aluminium oxide, Al_2O_3. Electrolysis is then carried out on the molten Al_2O_3, to extract the aluminium.

i) Write down the overall equation for the extraction of aluminium from aluminium oxide by electrolysis.

ii) State whether the aluminium oxide is oxidized or reduced in the electrolysis reaction. Explain your answer.

[3 marks]

Question 5b

b) Another ionic compound which can undergo electrolysis is molten lead bromide.

i) Explain, in terms of ions and electrons, what would happen in an electrolytic cell during the electrolysis of lead bromide, using carbon electrodes.

ii) State two different ways in which electrical charge flows in the electrolysis apparatus.

[4 marks]

Question 5c

c) State the products formed at each electrode during the electrolysis of molten lead bromide, giving the equations at each electrode with state symbols.

[3 marks]

Question 5d

d) Draw a labelled diagram of the apparatus suitable to carry out the electrolysis of molten lead bromide. Include the direction of electron flow, the negative electrode (cathode), the positive electrode (anode) and the electrolyte.

[3 marks]

Question 6a

a) The list below shows three metals from the activity series in order of reactivity.

Mg

Zn

Ni

Increasing reactivity

Deduce which of the three metals is the strongest reducing agent.

[1 mark]

Question 6b

b) A voltaic cell can be made by joining two half-cells together, such as Zn/Zn^{2+} and Ni/Ni^{2+}.

Write a balanced equation for the overall reaction taking place when the two half-cells are connected together, and state which species is undergoing oxidation.

[2 marks]

Question 6c

c) Cell diagrams are a way to represent the redox reactions taking place in voltaic cells.

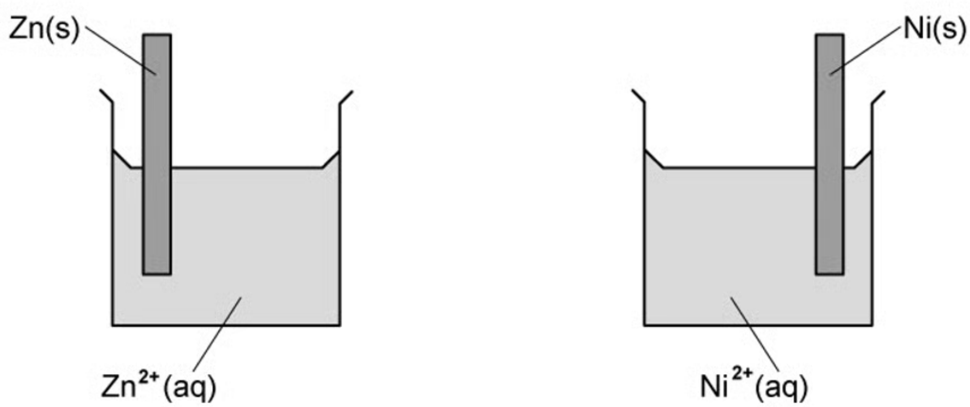
Write a cell diagram for the overall cell reaction taking place in part (b).

[1 mark]

Question 6d

d) Complete the partially labelled diagram in **Figure 1**, of the apparatus used in the voltaic cell in part (b). Show the direction of the movement of the electrons and ions in the cell.

Figure 1



[3 marks]