The **topics** and **types of questions** examined in this Achievement Standard. Use this sheet to plan and organise your study so that you cover everything that is required.

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3.3 OXIDATION REDUCTION AS 90696

Describe oxidation-reduction processes

3.3 1. Oxidation-Reduction

- appearance and state of oxidants and reductants and the product they are converted to
- oxidants: O_2, Cl_2, I_2, Fe^{3+}, dilute acid (with metals), H_2O_2, MnO_4^- (acidic, basic or neutral), Cu^{2+}, $Cr_2O_7^{2-}/H^+$, OCl^-, conc. HNO_3, IO_3^-, MnO_2
- reductants: metals, C, CO, H_2, Fe^{2+}, Br^-, I^-, H_2S, SO_2, SO_3^{2-}, $S_2O_3^{2-}$, H_2O_2, $H_2C_2O_4$

(a) Sodium thiosulfate solution reacts with dilute acid to form a precipitate of sulfur, and sulfur dioxide.

$S_2O_3^{2-} + 2H^+ \rightarrow S + SO_2 + H_2O$

(i) Give the oxidation numbers for sulfur in the species below.

$S_2O_3^{2-}$ S SO_2 .

(ii) Write balanced half-equations for the oxidation and reduction reactions.

(b) Oxalic acid ($H_2C_2O_4$) reacts with acidified potassium permanganate solution (MnO_4^- / H^+).

(i) Describe what would be observed.

(ii) Write balanced half-equations for the oxidation and reduction reactions, and then write a balanced equation for the overall reaction.

(c) Iodate ions, IO_3^-, react with iodide ions, I^-, in the presence of acid to form iodine, I_2.

Use oxidation numbers to identify the oxidant and the reductant in this reaction, and justify your answer.

(d) If a beaker of hydrogen peroxide is left at room temperature for a period of time, small bubbles appear on the bottom and sides of the beaker. Identify the gas in the bubbles and justify your decision by determining the $E°_{cell}$ for spontaneous reaction that occurs, using the $E°$ values given below. Include a balanced equation for the formation of the gas in the bubbles.

$E°$ (H_2O_2 / H_2O) = + 1.77 V $E°$ (O_2 / H_2O_2) = + 0.68 V

Hydrogen peroxide (H_2O_2) and iron (II) ions (Fe^{2+}) can act as both oxidant and reductant.

(a) What is the product formed from Fe^{2+} when it is

(i) oxidised? (ii) reduced?

(b) Hydrogen peroxide can undergo autoxidation-reduction. This means that one molecule of H_2O_2 acts as the oxidant while another molecule of H_2O_2 acts as the reductant. Using the ion-electron method write the half-equations and the overall balanced equation for the reaction that occurs.

(c) Identify the strongest oxidant from the electrochemical data below. Give a reason for your selection.

$E°(Fe^{3+} / Fe^{2+})$ = +0.77 V $E°(Fe^{2+} / Fe)$ = –0.44 V

$E°(H_2O_2 / H_2O)$ = +1.78 V $E°(O_2 / H_2O_2)$ = +0.68 V

(d) (i) Using the standard electrode potentials given in (c) above, show that a reaction can occur between H_2O_2 and Fe^{2+}.

(ii) Write a balanced equation for this reaction.

(iii) Describe what would be **observed** as the reaction proceeds.

Lead storage batteries have been used in cars for the past 85 years. The spontaneous reaction for each cell in the battery is:

$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(\ell)$

(a) (i) Give the oxidation numbers for lead in:

Pb _____ PbO_2 _____ $PbSO_4$ _____

(ii) Show how these oxidation numbers can be used to identify the oxidant and reductant in the cell reaction.

(b) Write balanced half-equations for the reactions occurring at the anode and the cathode of each cell in the battery.

3.3 2. Electrochemical cells

- use of reduction potentials
- spontaneity of oxidation-reduction reactions
- calculations may include oxidation numbers and mole ratios

The table below gives some standard electrode potentials that may be used in the questions that follow.

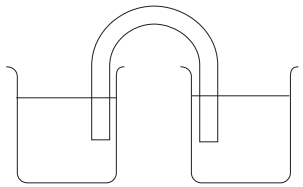
Reduction half equation	$E°$/V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	–2.37
$NO_3^-(aq) + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O$	+0.96

(a) Identify the species from the table that is the strongest reductant.

(b) The following electrochemical cell is required:

$Mg(s) / Mg^{2+}(aq) // Cu^{2+}(aq) / Cu(s)$

(i) Complete and label the diagram below to show how such a cell would be constructed.



(ii) The salt bridge could be made using a piece of paper soaked in a solution of potassium chloride.

Explain the role of the salt bridge in the cell.

(iii) Write the net ionic equation for the overall cell reaction.

(iv) Calculate the voltage of the cell in standard conditions.

(c) Copper metal reacts with dilute nitric acid but, unlike magnesium, has no reaction with dilute sulfuric acid.

(i) Account for the above observations using the standard reduction potentials in the table above.

(ii) Write a balanced equation for the reaction of copper metal with dilute nitric acid.