Write-on version

Chemistry 3.3 , 2004

90696 Describe oxidation–reduction processes

Credits: Three

A periodic table is provided on Resource Sheet L3-CHEMR.

You should answer ALL the questions in this booklet.

Show working for all calculations.

Achievement Criteria	*For Assessor's use only*	
Achievement	**Achievement with Merit**	**Achievement with Excellence**
Identify and describe oxidation–reduction processes.	Use information about oxidation–reduction processes.	Analyse and interpret information about oxidation–reduction processes.
	Overall Level of Performance	

You are advised to spend 30 minutes answering the questions in this booklet.

Question One: The fuel cell

A hydrogen-oxygen fuel cell consists of an electrolyte solution such as potassium hydroxide and two inert electrodes. Hydrogen and oxygen are bubbled into the cell and the following reactions occur.

$H_2 (g) + 2OH^-(aq) \rightarrow 2H_2O + 2e^-$

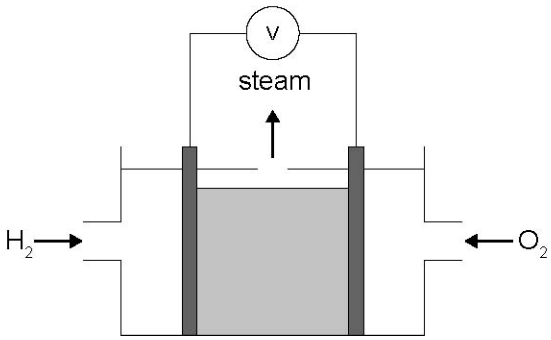
$O_2(g) + 2H_2O + 4e^- \rightarrow 4OH^-(aq)$

a Write the overall equation for the cell reaction.

b On the diagram of the cell (below) clearly label

i the anode and the cathode

ii the oxidant and the reductant.



c The standard cell potential (voltage) of this fuel cell is 1.23 V. Calculate the standard reduction potential for the conversion of H_2O to H_2 (g) and OH^- (aq).

$$E° (O_2 / OH^-) = 0.40 \text{ V}$$

Question Two: Reactions of hydrogen peroxide

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.

Question Three: Titration analysis

The standardisation of a solution of sodium thiosulfate can be done by iodometry. To do this an excess of potassium iodide is added to a solution of iodate, IO_3^-, and the following reaction occurs.

$$IO_3^- (aq) + 5I^-(aq) + 6H_3O^+(aq) \rightarrow 3I_2(aq) + 9H_2O \quad \textit{Reaction 1}$$

The iodine liberated is then titrated with sodium thiosulfate according to the reaction below:

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^- (aq) + S_4O_6^{2-}(aq) \quad \textit{Reaction 2}$$

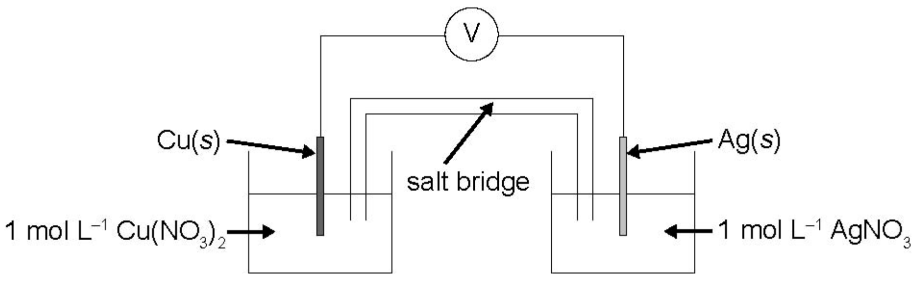
a Describe what would be observed in *Reaction 1* when the iodate is added to the iodide.

b An excess of potassium iodide is added to a solution containing 1.50×10^{-4} moles of IO_3^-. Calculate the amount (mol) of $S_2O_3^{2-}$ that must be added to reach the equivalence point in the titration.

c Using oxidation numbers, show how the species oxidised and the species reduced can be identified in *Reaction 1.*

Question Four: The electrochemical cell

An electrochemical cell is set up as follows:



$$E°(Ag^+/Ag) = +0.80 \text{ V}$$

$$E°(Cu^{2+}/Cu) = +0.34 \text{ V}$$

a What would be the reading on the voltmeter in the above cell?

b Give the standard cell diagram that represents this cell.

The voltmeter is replaced by a piece of wire that allows the flow of electrons.

c Discuss the flow of charge in this cell, both in terms of direction of movement and the species involved. Include the role of the salt bridge in your answer. (You may draw on the above diagram if it helps to clarify the description.)

d **i** Write a balanced equation for the reaction occurring in the electrochemical cell shown at the start of this Question 4.

ii Describe what would be **observed** in each half-cell as the cell discharges and explain these observations in terms of the chemical reactions occurring.

e The cell is allowed to discharge for a period of time during which the mass of the copper electrode changes by 3.20 g. Calculate the mass change that would be expected on the silver electrode. (Assume that the change in mass is only due to the oxidation-reduction reaction occurring as the cell discharges.)

f The Ag^+/Ag electrode is replaced with a standard hydrogen electrode. Fully discuss how this affects the operation of the cell, including aspects such as the charge on the electrodes, the cell voltage, and any changes occurring in the Cu^{2+}/Cu half-cell.