Practice task 3

Achievement Standard Chemistry 91393 (v2)

Demonstrate understanding of oxidation-reduction processes

Electrochemistry

Level 3 **Credits:** 3

Recommended time to complete: 1 hour

Assessment conditions: Closed book

Achievement	Achievement with Merit	Achievement with Excellence
Demonstrate understanding of oxidation-reduction processes.	Demonstrate in-depth understanding of oxidation-reduction processes.	Demonstrate comprehensive understanding of oxidation-reduction processes

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Student instructions

Introduction

This activity requires you to write a report demonstrating your understanding of oxidation-reduction in the context of electrolytic and electrochemical cells.

You are required to answer **both** questions.

You will be assessed on how comprehensive your understanding of the oxidation-reduction processes is demonstrated in this report.

Throughout your report, use correct chemical vocabulary, symbols and conventions.

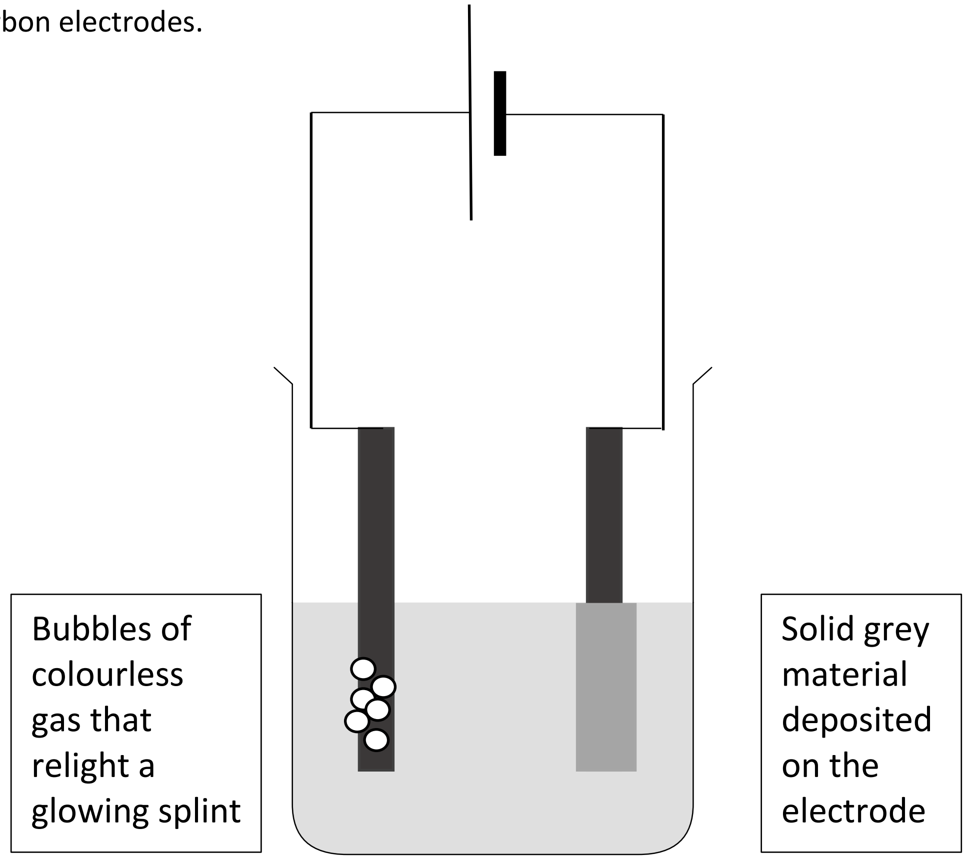
You will be provided with a chart of redox species and their appearance.

The following standard reduction potentials may be useful: not all will be used.

O_2 / H_2O	1.23 V	Zn^{2+} / Zn	-0.76 V
H^+ / H_2	0.00 V	Pb^{2+}/Pb	-0.13 V
H_2O / H_2	-0.83 V	MnO_4^-/Mn^{2+}	1.51 V
H_2O_2 / H_2O	1.78 V	Fe^{3+} / Fe^{2+}	0.77 V

Part One **Electrolysis**

This diagram shows an electrolytic cell. Electrolysis of lead nitrate $Pb(NO_3)_2$ (aq) is carried out using 2 carbon electrodes.



Write a report on the oxidation-reduction processes occurring in this electrolytic cell.

Colourless Pb^{2+} are attracted to the cathode (– electrode) where they are reduced to Pb, forming the grey layer. As Pb^{2+} ions are colourless the solution does not change colour.

$$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$$

This is reduction as each Pb^{2+} gains 2 electrons OR it is also reduction because the oxidation number of lead decreases from +2 in Pb^{2+} to 0 in Pb.

At the anode (+ electrode) water is oxidised to oxygen gas which is seen as bubbles of colourless gas.

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

This is oxidation as each water molecule loses 2 electrons OR it is oxidation as the oxidation number of oxygen increases from –2 in H_2O to 0 in O_2.

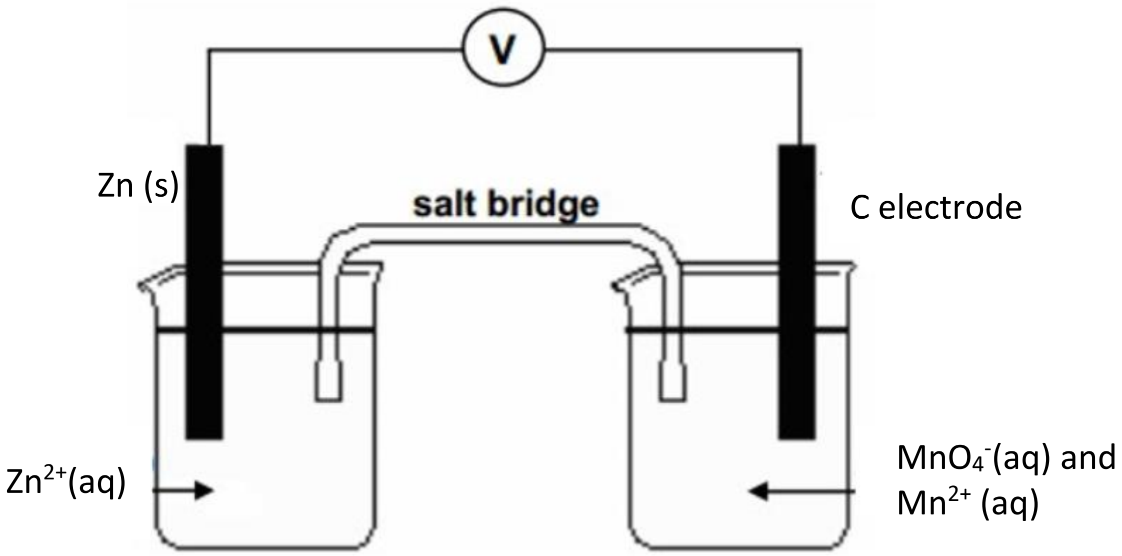
The overall reaction occurring is:

$$2Pb^{2+}(aq) + 2H_2O(l) \rightarrow 2Pb(s) + O_2(g) + 4H^+(aq)$$

$E^o cell = E^o(reduction) - E^o(oxidation)$ $E^o cell = -0.13 - 1.23 = -1.36V$. This reaction is not spontaneous as the value of E^ocell is negative and so requires a voltage of greater than 1.36V to be supplied to make the reaction occur. This is because O_2 is a better oxidising agent than Pb^{2+} and so the spontaneous reaction would be the reduction of oxygen to water and the oxidation of Pb to Pb^{2+}. This is why electrical energy is needed to make the reverse / non-spontaneous reaction occur.

<u>Part Two</u> **Electrochemical cells**

An electrochemical cell was set up as below, with two half cells, one containing Zn^{2+} / Zn and the other MnO_4^-/Mn^{2+} with a C electrode



Write a report on the oxidation-reduction processes occurring in this electrochemical cell.

At the carbon electrode, the cathode (+), reduction would occur. Purple $MnO_4^-(aq)$ would be reduced to colourless $Mn^{2+}(aq)$ and so the colour of the purple solution would fade.

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

Reduction is the gain of electrons. This is reduction as each MnO_4^- gains 5 electrons OR Reduction is the decrease in oxidation number. This is reduction as the oxidation number of Mn has decreased from +7 in MnO_4^- to +2 in Mn^{2+}.

At the zinc electrode, the anode (-), oxidation takes place. The grey Zn electrode is oxidised to colourless $Zn^{2+}(aq)$ and so the electrode gets smaller / starts to dissolve. As the ions are colourless there is no colour change in this beaker.

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$. Oxidation is the loss of electrons. This is oxidation as each Zn atom loses 2 electrons to form $Zn^{2+}(aq)$ OR Oxidation is the

increase in oxidation number. This is oxidation as the oxidation number of zinc increases from 0 in Zn to +2 in Zn^{2+}.
The overall reaction is:

$$2MnO_4^-(aq) + 16H^+(aq) + 5Zn(s) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Zn^{2+}(aq)$$

E^ocell = E^O(reduction) – E^o(oxidation) E^ocell = 1.51 - -0.76 = 2.27V. This reaction is spontaneous as the value of E^ocell is positive.

The standard reduction potential of MnO_4^-/Mn^{2+} is more positive (1.51 V) than that of Zn^{2+}/Zn (-0.76 V) which means that since MnO_4^- is a stronger oxidising agent that Zn^{2+}, the spontaneous reaction is the reduction of MnO_4^- to Mn^{2+} and the oxidation of Zn to Zn^{2+}.