Sample Task - answers provided for part Two

Achievement Standard Chemistry 91393

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

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.

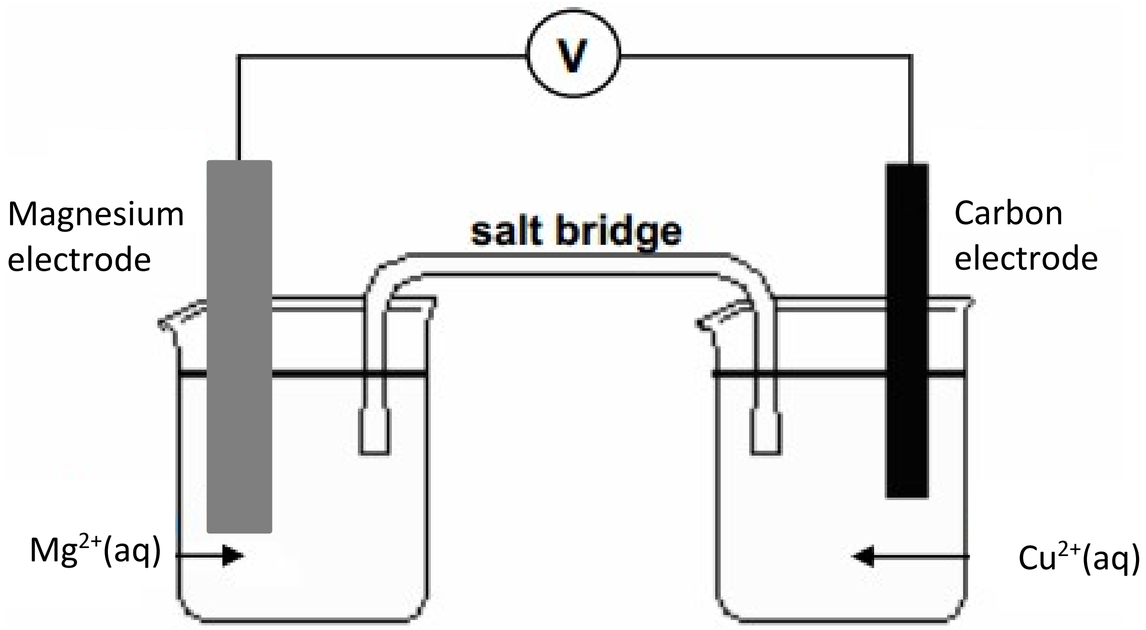
You will be provided with a chart of redox species and their appearance.

The following standard reduction potentials may be useful: NOT ALL WILL BE USED.

O_2 / H_2O	1.23 V	Mn^{2+} / Mn	-1.18 V
H^+ / H_2	0.00 V	Mg^{2+}/ Mg	-2.37 V
H_2O / H_2	-0.83 V	Cl_2 / Cl^-	+ 1.36 V
Cu^{2+}/ Cu	0.34 V	Co^{2+}/ Co	-0.28 V

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

An electrochemical cell was set up.



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

Include in your report:

- Identify the anode and cathode electrodes with their respective charges
- Describe the expected observations at each electrode and in each beaker, clearly linked to the all the species involved – if there is no observable change, explain why.
- Describe the redox process that occurs at each electrode, identifying the species oxidised and reduced by name or formula.
- Write balanced half equations for the both oxidation and reduction processes.
- Write a fully balanced redox equation.
- Justify both redox processes occurring using oxidation numbers **and/or** the loss or gain of electrons. (You must state how many electrons are lost or gained).
- Write the cell diagram for this cell and determine potential difference of the cell.
- Explain the spontaneity of the cell with reference to the standard reduction potentials.

Part 2 – model answer.

At the grey Mg electrode, Mg (atoms) will be oxidised to Mg^{2+} ions. The grey Mg electrode will (start to) dissolve. There will be no colour change to the colourless solution as $Mg^{2+}(aq)$ ions are colourless.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

Oxidation is the loss of electrons. This is oxidation as each Mg atom has lost <u>two</u> electrons as it forms Mg^{2+} (OR Oxidation is an increase in oxidation number. This is oxidation as the oxidation number of Mg has increased from 0 in Mg to +2 in Mg^{2+}).

At the (black) carbon electrode, a layer of pinky orange copper would be seen and the blue colour of the solution would (begin to) fade. This is because blue Cu^{2+} ions are reduced to pinky orange coloured Cu.

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Reduction is the gain of electrons. This is reduction as each Cu^{2+} ions gains <u>two</u> electrons. (OR reduction is the decrease in oxidation number. This is reduction as the oxidation number of Cu in Cu^{2+} has decreased from +2 to 0 in Cu.)

Overall fully balanced redox equation is: $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$

Cell diagram (not assessed) $Mg \mid Mg^{2+} \parallel Cu^{2+} \mid Cu$

$E^o cell = E^o(red) - E^o(ox) = 0.34 - -2.37 = +2.71$ V

As this value is positive, the reaction $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$ is spontaneous.

This is because the reduction potential for Cu^{2+}/Cu is 0.34V which is more positive than the reduction potential of Mg^{2+}/Mg at -2.37V. This means Cu^{2+} is a stronger oxidising agent (or oxidant) than Mg^{2+}, and so the spontaneous reaction is the reduction of Mg to Mg^{2+} and the oxidation of Cu^{2+} to Cu. (as seen above)