**Student 1: Excellence**

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Electrochemical cell:

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

Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$ 1

Overall: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

$Cu^{2+}$ is reduced due to the fact that it gains 2 electrons in order to become Cu and reduction is the gain of electrons. The oxidation number for Cu decreases from +2 in $Cu^{2+}$to 0 in Cu which shows that this is a reduction reaction.

Zn is oxidised due to the fact that it loses 2 electrons in order to become $Zn^{2+}$ and oxidation is the loss of electrons. The oxidation number for Zn increases from 0 in Zn to +2 in $Zn^{2+}$ which shows that this is an oxidation reaction.

At the right-hand electrode, the blue $Cu^{2+}$ solution is reduced to form a brown Cu solid which is deposited on the brown Cu electrode causing it to gain mass. The blue solution in the beaker will lighten as $Cu^{2+}$ is reduced.

At the left-hand electrode, the silver Zn solid is oxidised forming the colourless solution of $Zn^{2+}$. This means that the silver/grey Zn electrode will decrease in mass. The solution in the beaker remains colourless.

$E°_{cell}$ = E°reduction – E°oxidation

=+0.34 – -0.76 2

= 1.10 V

Since the $E°_{cell}$ is positive this means that the reaction is spontaneous in the direction shown. Therefore, electrochemical potential energy is converted into electrical energy which is seen through the 1.10 V of electrical energy produced by the cell under standard conditions.

$E°_{cell}$ ($Zn^{2+}$/Zn) = -0.76V $E°_{cell}$ ($Cu^{2+}$/Cu) = +0.34V

($Cu^{2+}$/Cu) > ($Zn^{2+}$/Zn) Therefore $Cu^{2+}$ is reduced because it has a more positive $E°_{cell}$ and the right-hand electrode gains mass and Zn oxidised because it has a more negative $E°_{cell}$ and the left-hand electrode losses mass. Also, because the $Cu^{2+}$ is has a more positive $E°_{cell}$ it means that it is the stronger oxidant and therefore the reactions will occur spontaneously. 3

Electrolytic cell:

Since both the graphite electrodes are inert, then either $Cu^{2+}$ or $H_2O$ is reduced and $SO_4^{2-}$ or $H_2O$ is oxidised.

Since the right-hand electrode gains mass then $Cu^{2+}$ must be reduced.

Since bubbles which relight a glowing splint are observed at the left-hand electrode, this indicates the presence of $O_2$, meaning water must be oxidised. As well as this $SO_4^{2-}$ is fully oxidised already.

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

Oxidation: $2H_2O \rightarrow O^{2-} + H^+ + 4e^-$ 1

Overall: $2Cu^{2+} + 2H_2O \rightarrow 2Cu + O^{2-} + H^+$

$Cu^{2+}$ is reduced because it gains 2 electrons in order to become Cu and reduction is the gain of electrons. The oxidation number for Cu decreases from +2 in $Cu^{2+}$to 0 in Cu which also shows that this is a reduction reaction.

$2H_2O$ is oxidised because it loses 4 electrons in order to become $O_2$ and oxidation is the loss of electrons. The oxidation number for O increases from -2 in $H_2O$ to 0 in $O_2$ which shows that this is an oxidation reaction.

At the right-hand electrode, the blue $Cu^{2+}$ solution is reduced to brown Cu solid which is deposited on the right-hand electrode (cathode) which therefore gain mass. The blue solution in the beaker will lighten as $Cu^{2+}$ is reduced. At the left-hand electrode, colourless $H_2O$ liquid is oxidised to the colourless $O_2$ gas, which is released as bubbles at the anode and these bubbles can relight a glowing splint.

$E°_{cell}$ = E°reduction – E°oxidation

=+0.34 – + 1.23

= -0.89 V 2

Since the $E°_{cell}$ is negative, this means that the reaction is non-spontaneous in the direction shown. This means that in order for the reaction to occur, an input of electrical energy of at least 0.90 V is required from the power supply.

($Cu^{2+}$/Cu) < ($O_2$/$H_2O$) $Cu^{2+}$ has a more negative $E°_{cell}$ and $H_2O$ is more positive. This means that as the reaction is shown it will not proceed because $Cu^{2+}$ is a stronger reductant than $H_2O$. So, because $Cu^{2+}$ is has a more negative $E°_{cell}$ value it means that the reaction will not occur spontaneously. 3

Both the electrochemical and electrolytic cell reactions produce the brown Cu metal at the right-hand electrode through the reduction of $Cu^{2+}$/Cu. However, the way this is achieved is different. The REDOX reaction in the electrochemical cell which forms Cu is a spontaneous reaction where electrochemical potential energy is converted into electrical energy. This reaction produces 1.10 V under standard conditions. In comparison, the REDOX reaction in the electrolytic cell where Cu is produced is a non-spontaneous reaction, and requires the input of electrical energy to force the reaction to occur. The electrolytic cell requires more than 0.89 V of energy to make the reaction proceed. Therefore, electrochemical cells convert electrochemical potential energy into electrical energy to produce Cu, whereas electrolytic cells require electrical energy to produce Cu. 4