

Electrochemical cell:



①

$\text{Cu}^{2+}$  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  $\text{Cu}^{2+}$  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  $\text{Zn}^{2+}$  and oxidation is the loss of electrons. The oxidation number for Zn increases from 0 in Zn to +2 in  $\text{Zn}^{2+}$  which shows that this is an oxidation reaction.

At the right-hand electrode, the blue  $\text{Cu}^{2+}$  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  $\text{Cu}^{2+}$  is reduced.

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

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{reduction}} - E^\circ_{\text{oxidation}} \\ &= +0.34 - -0.76 \\ &= 1.10 \text{ V} \end{aligned}$$

②

Since the  $E^\circ_{\text{cell}}$  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.

$$E^\circ_{\text{cell}} (\text{Zn}^{2+}/\text{Zn}) = -0.76\text{V} \qquad E^\circ_{\text{cell}} (\text{Cu}^{2+}/\text{Cu}) = +0.34\text{V}$$

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

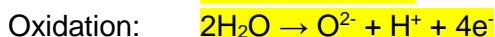
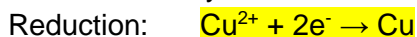
③

Electrolytic cell:

Since both the graphite electrodes are inert, then either  $\text{Cu}^{2+}$  or  $\text{H}_2\text{O}$  is reduced and  $\text{SO}_4^{2-}$  or  $\text{H}_2\text{O}$  is oxidised.

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

Since bubbles which relight a glowing splint are observed at the left-hand electrode, this indicates the presence of  $\text{O}_2$ , meaning water must be oxidised. As well as this  $\text{SO}_4^{2-}$  is fully oxidised already.



①

$\text{Cu}^{2+}$  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  $\text{Cu}^{2+}$  to 0 in Cu which also shows that this is a reduction reaction.

$2\text{H}_2\text{O}$  is oxidised because it loses 4 electrons in order to become  $\text{O}_2$  and oxidation is the loss of electrons. The oxidation number for O increases from -2 in  $\text{H}_2\text{O}$  to 0 in  $\text{O}_2$  which shows that this is an oxidation reaction.

At the right-hand electrode, the blue  $\text{Cu}^{2+}$  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  $\text{Cu}^{2+}$  is reduced. At the left-hand electrode, colourless  $\text{H}_2\text{O}$  liquid

is oxidised to the colourless O<sub>2</sub> gas, which is released as bubbles at the anode and these bubbles can relight a glowing splint.

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{reduction}} - E^{\circ}_{\text{oxidation}} \\ &= +0.34 - +1.23 \\ &= -0.89 \text{ V} \end{aligned}$$

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Since the  $E^{\circ}_{\text{cell}}$  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.

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(Cu<sup>2+</sup>/Cu) < (O<sub>2</sub>/H<sub>2</sub>O) Cu<sup>2+</sup> has a more negative  $E^{\circ}_{\text{cell}}$  and H<sub>2</sub>O is more positive. This means that as the reaction is shown it will not proceed because Cu<sup>2+</sup> is a stronger reductant than H<sub>2</sub>O. So, because Cu<sup>2+</sup> has a more negative  $E^{\circ}_{\text{cell}}$  value it means that the reaction will not occur spontaneously.

Both the electrochemical and electrolytic cell reactions produce the brown Cu metal at the right-hand electrode through the reduction of Cu<sup>2+</sup>/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.

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