Edexcel International Chemistry A Level

CP12 - Investigating Electrochemical Cells

(A Level only)

Flashcards

What must a half cell contain?

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An element in 2 oxidation states

Define the terms oxidation and reduction

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Oxidation - loss of electrons

Reduction - gain of electrons

Why must metal electrodes be cleaned with sandpaper before creating an electrochemical cell?

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To remove any metal oxide that has formed on the surface and improve electrical conductivity.

Describe the movement of electrons in an electrochemical cell

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Electrons flow through the wire from the positive electrode to the negative electrode.

Why is a salt bridge used in an electrochemical cell?

Why is a salt bridge used in an electrochemical cell?

To maintain the charge balance and complete the circuit.

This is because negative electrons are moving from one half cell to another. Without the salt bridge, positive charge would build up in the half cell containing the anode and negative charge would build up in the half cell containing the cathode. This would cause the reaction to stop.

Why must an inert salt be used in the salt bridge?

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So the salt doesn't react with the solutions and alter their concentrations. If a reactive salt was used, the cell potential would change.

What moves across the salt bridge?

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Ions

What are standard conditions?

What are standard conditions?

Solution concentrations of 1 mol dm^{-3}

Pressure of 101 kPa

Temperature of 298 K

What does EMF mean?

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Electromotive force.

The voltage when no current flows.

What must the cell potential value be, for a process to be feasible?

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Cell potential must be greater than 0

How can cell potential be calculated?

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$$E^{\theta}_{cell} = E^{\theta}(\text{positive terminal}) - E^{\theta}(\text{negative terminal})$$

Why wouldn't silver nitrate solution be used in a standard half cell?

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Silver nitrate is highly oxidising making it too dangerous to be used at a standard concentration of 1 mol dm^3

Why wouldn't potassium be used to make an inert electrode for a half cell?

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Potassium is highly reactive. It would react with water in the solution, producing a lilac flame, hydrogen and potassium hydroxide.

Why might theoretical cell potential values be different to values obtained experimentally?

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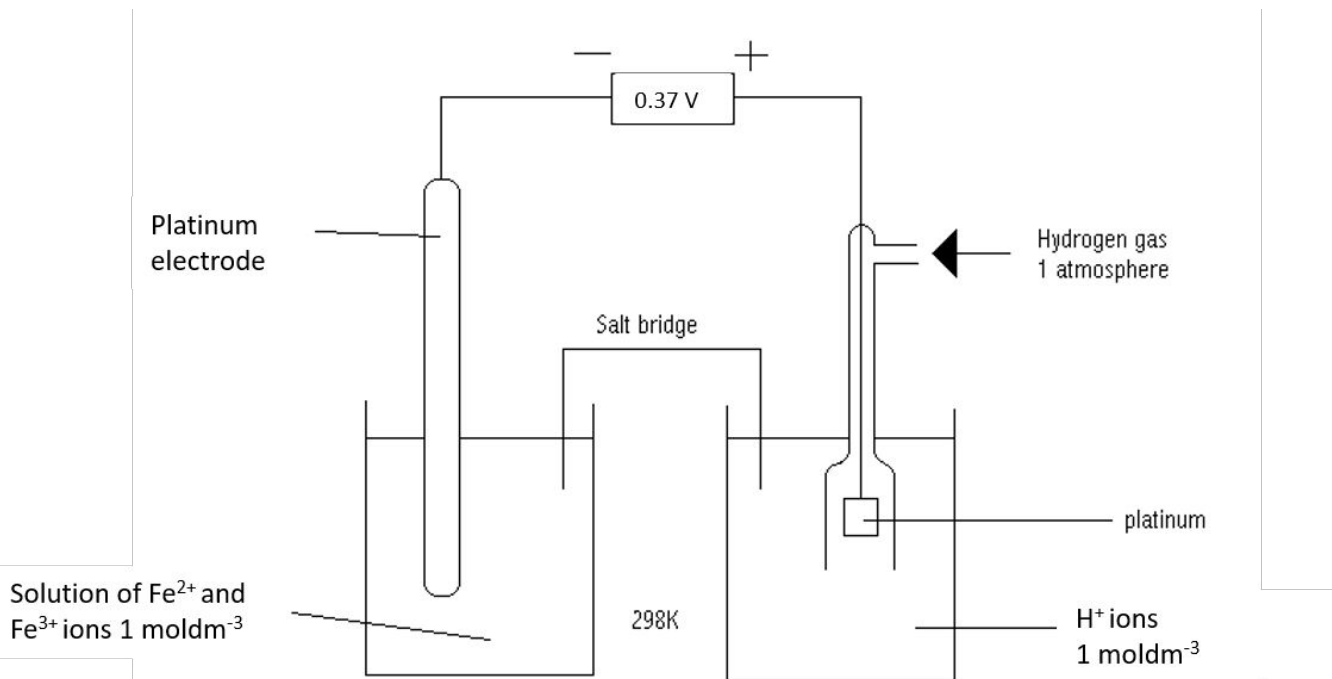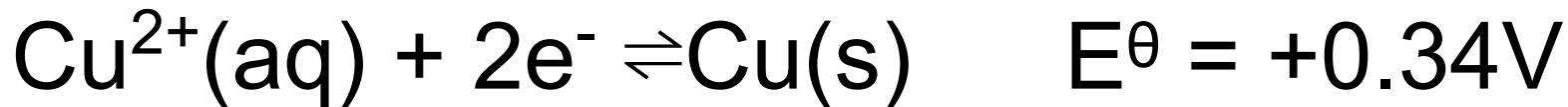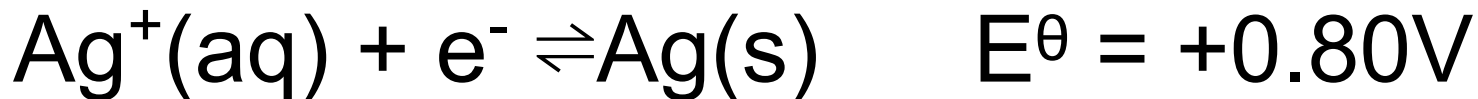
Conditions may be non-standard

Draw a diagram showing a standard Fe^{2+}/Fe^{3+} cell

Draw a diagram showing a standard Fe^{2+}/Fe^{3+} cell

− +

0.37 V

Platinum electrode

Hydrogen gas 1 atmosphere

Salt bridge

platinum

Solution of Fe^{2+} and Fe^{3+} ions 1 moldm^{-3}

298K

H^+ ions 1 moldm^{-3}

A cell is made up of the following half cells:

$$Ag^+(aq) + e^- \rightleftharpoons Ag(s) \qquad E^\theta = +0.80V$$
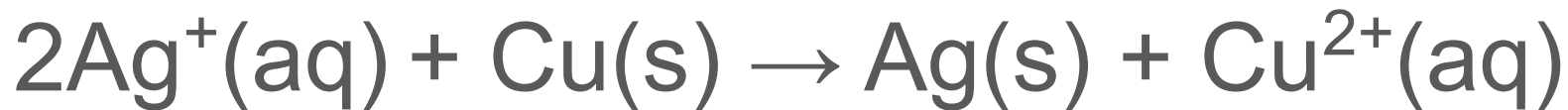
$$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s) \qquad E^\theta = +0.34V$$

Write the overall cell equation and calculate the standard cell potential

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$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ $E^\theta = +0.80V$

$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ $E^\theta = +0.34V$

Write the overall cell equation and calculate the standard cell potential

$$2Ag^+(aq) + Cu(s) \rightarrow Ag(s) + Cu^{2+}(aq)$$

$$E^\theta_{cell} = +0.80 - (+0.34) = 0.46V$$

In an electrochemical cell, is the more negative half cell oxidised or reduced?

In an electrochemical cell, is the more negative half cell oxidised or reduced?

Oxidised