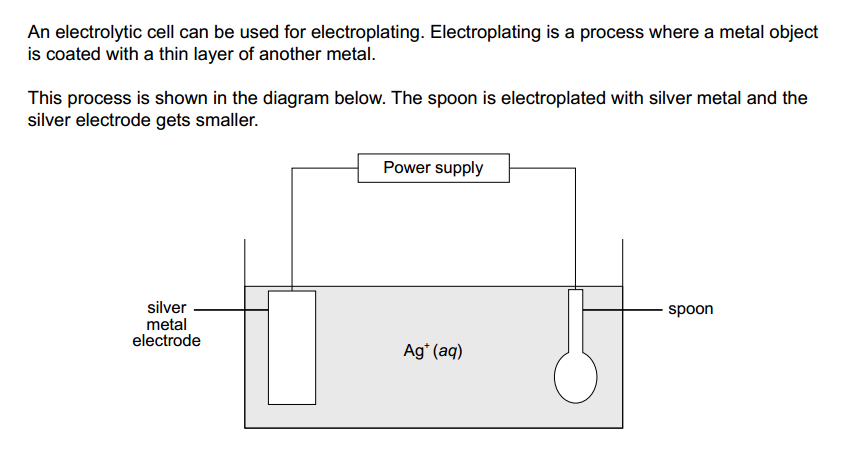
Analysing ELECTROLYTIC cells - Answers

For each cell described below:

* + Describe the observations and relate your observations to the species involved
  + Describe what is being oxidised and what is being reduced in each cell. Justify the oxidation and reduction half reactions in terms of e transfer and/or oxidation number
  + Write half equations and then a balanced net ionic equation
  + Calculate cell potentials using your data sheet, state whether the reaction is spontaneous or not and relate this to energy input/output

# Cell 1

**Cell 1**



E°cell = E°(R) - E°(O), = 0.80 – 0.80 = 0.00V

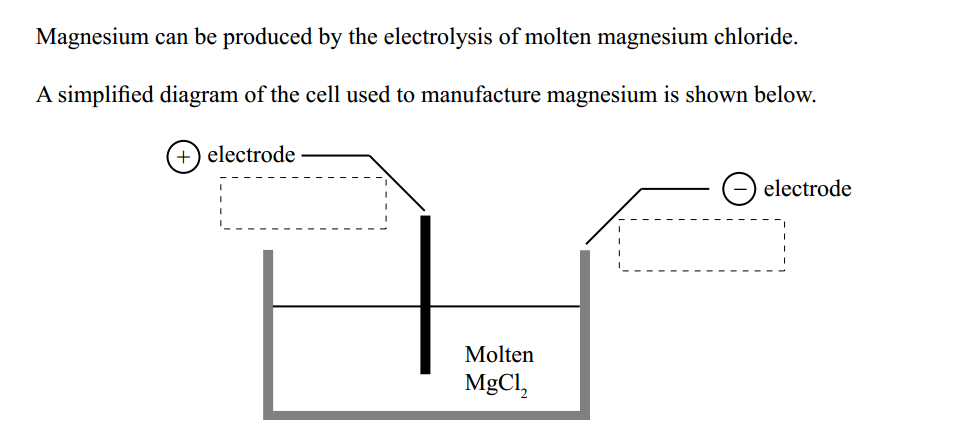
Reduction: Ag+ + e → Ag (reduction as ON of Ag is decreasing from +1 in Ag+ to 0 in Ag, decrease in ON = reduction)

Oxidation: Ag → Ag+ + e (oxidation as ON of Ag is increasing from to 0 in Ag to +1 in Ag+, increase in ON = oxidation)

Observations: No colour change in the solution as Ag+ is colourless and is being produced by oxidation and removed by reduction. The anode (silver electrode) would decrease in mass as Ag is oxidised and the spoon would be coated in a grey/silver solid of Ag and would increase in mass as Ag+ is reduced to Ag at the cathode.

The E°cell is zero, therefore the reaction is non-spontaneous. For this reaction to occur, an input of electrical energy is required from the power supply.

# Cell 2



Mg2+ must be reduced. Cl- must be oxidised.

E°cell = E°(R) - E°(O), = -2.36 – 1.40 = -3.76V

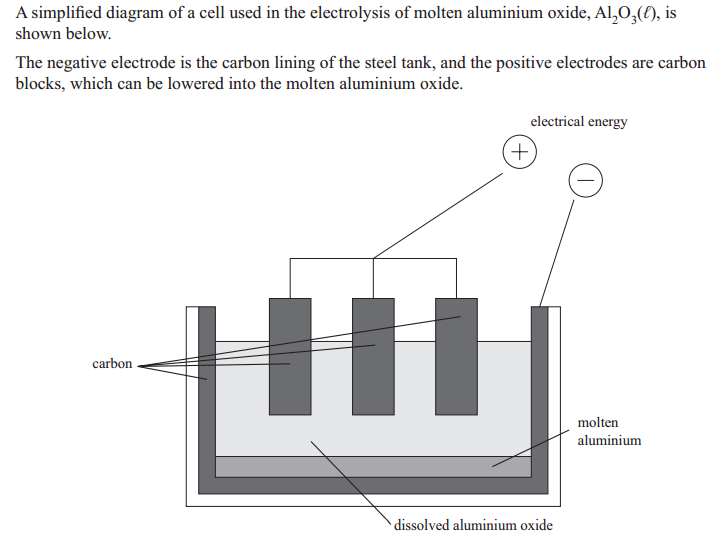
Reduction: Mg+2 + 2e → Mg (reduction as ON of Mg is decreasing from +2 in Mg+2 to 0 in Mg, decrease in ON = reduction)

Oxidation: 2Cl- → Cl2 + 2e (oxidation as ON of Cl is increasing from to -1 in Cl- to 0 in Cl2, increase in ON = oxidation)

Overall equation: Mg+2 + 2Cl- → Mg + Cl2

Observations: No colour change in the solution as Mg2+ is colourless and so is Cl-. At the anode a pale green gas of Cl2 would appear (probably wouldn’t bubble) and the cathode would be coated in a grey/silver solid/liquid of Mg.

The E°cell is negative, therefore the reaction is non-spontaneous. For this reaction to occur, an input of electrical energy is required from the power supply.



# Cell 3

Al3+ must be reduced. O2- must be oxidised.

E°cell = E°(R) - E°(O), = -2.36 – 1.40 = -3.76V

Reduction: Mg+2 + 2e → Mg (reduction as ON of Mg is decreasing from +2 in Mg+2 to 0 in Mg, decrease in ON = reduction)

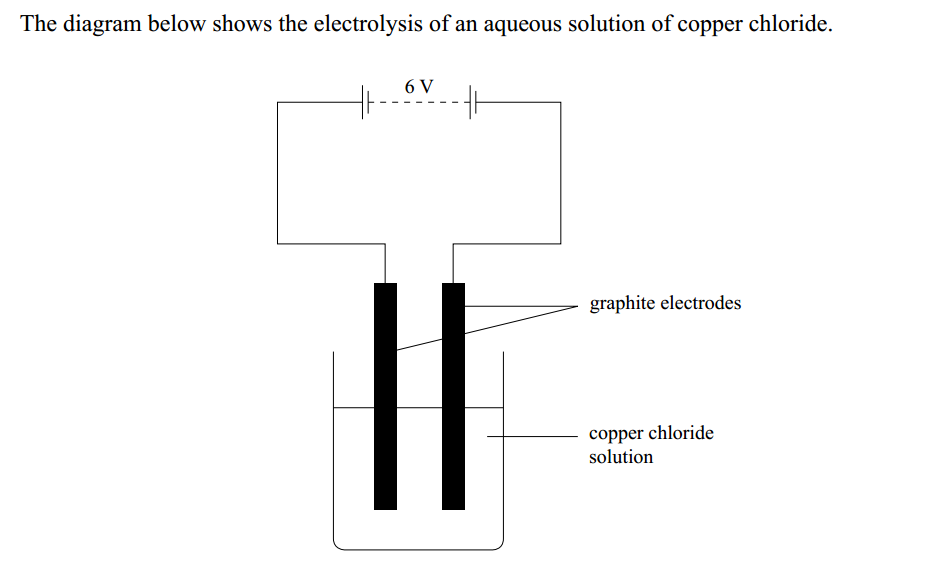
Oxidation: 2Cl- → Cl2 + 2e (oxidation as ON of Cl is increasing from to -1 in Cl- to 0 in Cl2, increase in ON = oxidation)

Overall equation: Mg+2 + 2Cl- → Mg + Cl2

Observations: No colour change in the solution as Mg2+ is colourless and so is Cl-. At the anode a pale green gas of Cl2 would appear (probably wouldn’t bubble) and the cathode would be coated in a grey/silver solid/liquid of Mg.

The E°cell is negative, therefore the reaction is non-spontaneous. For this reaction to occur, an input of electrical energy is required from the power supply.

# Cell 4



As the electrodes are inert, Cu2+ OR H2O must be reduced. Cl- OR H2O must be oxidised.

As E°cell (Cu/Cu2+) = +0.34 V and (H2O/H2) = 0.00V, then copper ions will be reduced as the E°cell is more positive.

As E°cell (Cl‑/Cl2) = +1.40 V and (H2O/O2) = +0.82 V, then water will be oxidised as the E°cell is more negative.

E°cell = E°(R) - E°(O), = + 0.34 V – +0.82 V = -0.50 V

Reduction: Cu+2 + 2e → Cu (reduction as ON of Cu is decreasing from +2 in Cu+2 to 0 in Cu, decrease in ON = reduction)

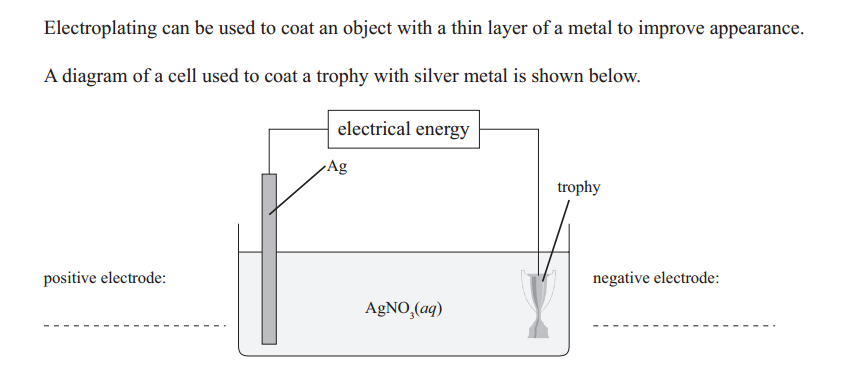
Oxidation: 2H2O → O2 + 4H+ + 4e (oxidation as ON of O is increasing from to -2 in H2O to 0 in O2, increase in ON = oxidation)

Overall equation: 2Cu+2 + 2H2O → 2Cu + O2 + 4H+

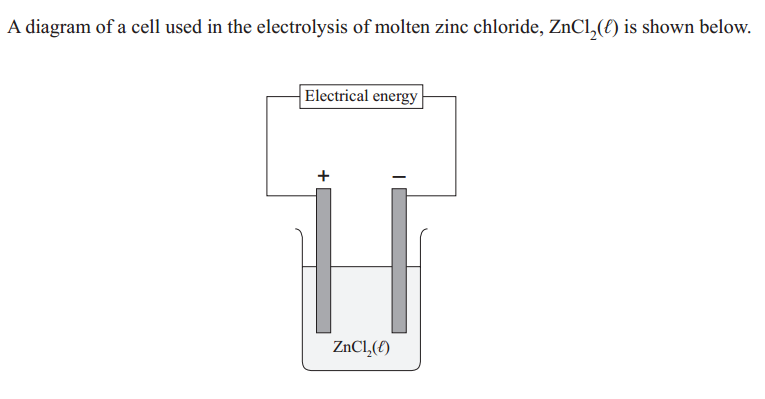
Observations: The blue solution lightens as Cu2+ is reduced to Cu (decrease in concentration of Cu2+). This is reduction therefore occurs at the cathode. In addition the anode increase in mass as a red brown solid of Cu forms on it. At the anode bubbles of a colourless gas appear (O2) as H2O (colourless liquid) is oxidised to O2.

The E°cell is negative, therefore the reaction is non-spontaneous. For this reaction to occur, an input of electrical energy is required from the power supply.

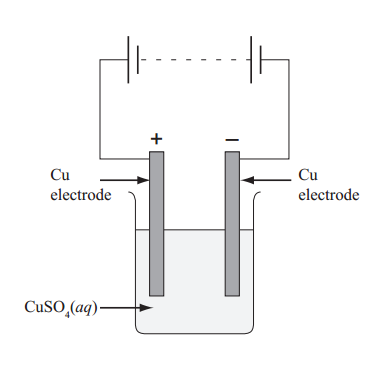
# Cell 5

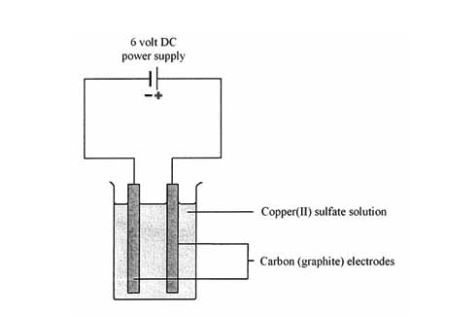


# Cell 6



# Cell 7 What is the same and different about these 2 cells? How does that affect the products?





Same: Both will reduce Cu2+ ions to copper.

Different: In the cell with copper electrodes, the Cu metal can be oxidised to form Cu2+ and therefore the reaction can continue to make Cu (eg in electroplating)

RIGHT CELL: As the electrodes are inert, Cu2+ OR H2O OR SO42- must be reduced. ONLY H2O must be oxidised (as SO42- cannot be oxidised further)

As E°cell (Cu/Cu2+) = +0.34 V, (SO42-/SO2) = +0.16 V and (H2O/H2) = 0.00V, then copper ions will be reduced as the E°cell is the most positive.

E°cell = E°(R) - E°(O), = + 0.34 V – +0.82 V = -0.50 V

Reduction: Cu+2 + 2e → Cu (reduction as ON of Cu is decreasing from +2 in Cu+2 to 0 in Cu, decrease in ON = reduction)

Oxidation: 2H2O → O2 + 4H+ + 4e (oxidation as ON of O is increasing from to -2 in H2O to 0 in O2, increase in ON = oxidation)

Overall equation: 2Cu+2 + 2H2O → 2Cu + O2 + 4H+

Observations: The blue solution lightens as Cu2+ is reduced to Cu (decrease in concentration of Cu2+). This is reduction therefore occurs at the cathode. In addition the cathode increases in mass as a red brown solid of Cu forms on it. At the anode bubbles of a colourless gas appear (O2) as H2O (colourless liquid) is oxidised to O2.

The E°cell is negative, therefore the reaction is non-spontaneous. For this reaction to occur, an input of electrical energy is required from the power supply.

# LEFT CELL:

As the electrodes are reactive, Cu2+ OR H2O OR SO42- must be reduced. H2O OR Cu can be oxidised (as SO42- cannot be oxidised further)

As E°cell (Cu/Cu2+) = +0.34 V, (SO42-/SO2) = +0.16 V and (H2O/H2) = 0.00V, then copper ions will be reduced as the E°cell is the most positive.

As E°cell (Cu/Cu2+) = +0.34 V, (H2O/O2) = +0.82V, then copper will be oxidised as the E°cell is the most negative.

E°cell = E°(R) - E°(O), = + 0.34 V – 0.34 V = 0.00 V

Reduction: Cu+2 + 2e → Cu (reduction as ON of Cu is decreasing from +2 in Cu+2 to 0 in Cu, decrease in ON = reduction)

Oxidation: Cu→ Cu+2 + 2e (oxidation as ON of Cu is increasing from to 0 in Cu to +2 in Cu2+, increase in ON = oxidation)

Overall equation: Cu+2 + Cu → Cu + Cu+2

Observations: The blue solution DOES NOT CHANGE COLOUR as Cu2+ is reduced to Cu and Cu is oxidised to Cu2+, thus the concentration of blue Cu2+ ions remains constant and the colour stays the same blue. This is reduction therefore occurs at the cathode. In addition the cathode increases in mass as a red brown solid of Cu forms on it. The anode decreases in mass as the red brown solid is oxidised to Cu2+.

The E°cell is zero, therefore the reaction is non-spontaneous. For this reaction to occur, an input of electrical energy is required from the power supply.

# Cell 8 Electrolysis of aqueous NaCl

