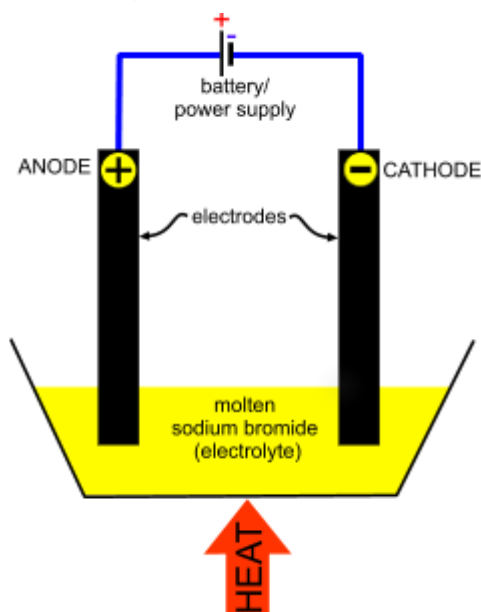


1 Sketch a labelled diagram of an electrolytic cell for electrolysis of molten sodium bromide.



2 Explain why a salt must be molten for electrolysis to occur

In the solid state the ions are held tightly in a lattice structure – they are not free to move. When molten, the ions are free to move so the electrolyte conducts electricity

3 Explain how current is conducted in an electrolytic cell.

In the external circuit electrons move through wires; in the electrolyte ions move between the electrodes

4 For electrolysis of each of the molten salts in the table:

Molten sodium chloride	Molten aluminium oxide
Molten zinc bromide	Molten magnesium chloride

- state the products at the anode and the cathode
- write half-equations for the reactions at the electrodes

	anode (+)	cathode (-)
Molten sodium chloride	Chlorine $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$	Sodium $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$
Molten zinc bromide	Bromine $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$	Zinc $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$
Molten aluminium oxide	Oxygen $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$	Aluminium $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$
Molten magnesium chloride	Chlorine $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$	Magnesium $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$

Sometime state symbols are required - in each case here the state symbol of the ions will be (l)  
e.g.  $2\text{Cl}^-(\text{l}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$

4 Explain at which electrodes oxidation and reduction occur in an electrolytic cell

Oxidation always occurs at the anode in electrolysis: negative ions will lose electrons (get oxidised) at the positive electrode.

Reduction always occurs at the cathode in electrolysis: positive ions will gain electrons (get reduced) at the negative electrode.

5 State two differences between an electrolytic cell and a voltaic cell

electrolytic cell	voltaic cell
electrical energy converted to chemical energy	chemical energy converted to electrical energy
non-spontaneous reaction	spontaneous reaction
Anode positive and cathode negative	Anode negative and cathode positive

6 For electrolysis, using platinum electrodes, of each of the aqueous solutions in the table:

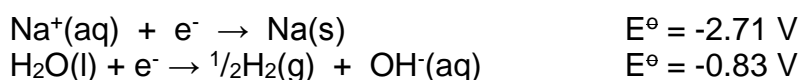
Concentrated NaCl(aq)	CuSO <sub>4</sub> (aq)
H <sub>2</sub> SO <sub>4</sub> (aq) (acidified water)	MgBr <sub>2</sub> (aq)

- state the products at the anode and the cathode
- write half-equations for the reactions at the electrodes

Half-equations are in the data booklet

	anode (+)	cathode (-)
Concentrated NaCl(aq)	Chlorine $2\text{Cl}^-(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$	Hydrogen $\text{H}_2\text{O}(\text{l}) + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2(\text{g}) + \text{OH}^-(\text{aq})$
H <sub>2</sub> SO <sub>4</sub> (aq) (acidified water)	Oxygen $\text{H}_2\text{O}(\text{l}) \rightarrow \frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$	Hydrogen $\text{H}_2\text{O}(\text{l}) + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2(\text{g}) + \text{OH}^-(\text{aq})$
CuSO <sub>4</sub> (aq)	Oxygen $\text{H}_2\text{O}(\text{l}) \rightarrow \frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$	Copper $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$
MgBr <sub>2</sub> (aq)	Bromine $2\text{Br}^-(\text{aq}) \rightarrow \text{Br}_2(\text{l}) + 2\text{e}^-$	Hydrogen $\text{H}_2\text{O}(\text{l}) + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2(\text{g}) + \text{OH}^-(\text{aq})$

7 Explain why hydrogen is produced at the cathode when aqueous sodium chloride is electrolysed but copper is produced at the cathode when aqueous copper(II) sulfate is electrolysed



The reduction of water to hydrogen is more favourable (more positive/less negative) electrode potential than the reduction of sodium ions to sodium

8 Explain how concentration of a sodium chloride solution can affect the products of electrolysis

At very low concentrations electrolysis of the sodium chloride solution produces oxygen at anode



Oxidation so opposite signs to those in the data booklet

The oxidation of water to oxygen is more favourable than the oxidation of chloride ions

At higher concentrations chlorine becomes the major product at the anode

As Cl<sup>-</sup> concentration increases the forward reaction

becomes more favourable – oxidation becomes more favourable.

The actual reason is more complicated than this but this is sufficient for exams

9 Explain two observations when copper(II) sulfate solution is electrolysed with platinum electrodes

- Pink-brown metal produced at cathode – Cu<sup>2+</sup>(aq) ions reduced to copper
- Blue colour of solution fades – Cu<sup>2+</sup>(aq) ions are responsible for the blue colour – as these are removed from the solution at the cathode, the concentration of Cu<sup>2+</sup>(aq) decreases.
- Bubbles of gas at anode – water oxidized to oxygen
- pH of solution decreases – concentration of H<sup>+</sup> ions increases as water is oxidized at the anode:  $\text{H}_2\text{O}(\text{l}) \rightarrow \frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$

10 Explain two observations when copper sulphate solution is electrolysed with copper electrodes

- Pink-brown metal produced at cathode – Cu<sup>2+</sup>(aq) ions reduced to copper
- Anode gets smaller as Cu<sup>2+</sup> ions go into solution – no bubbles of gas – oxidation of copper is more favourable than oxidation of water  

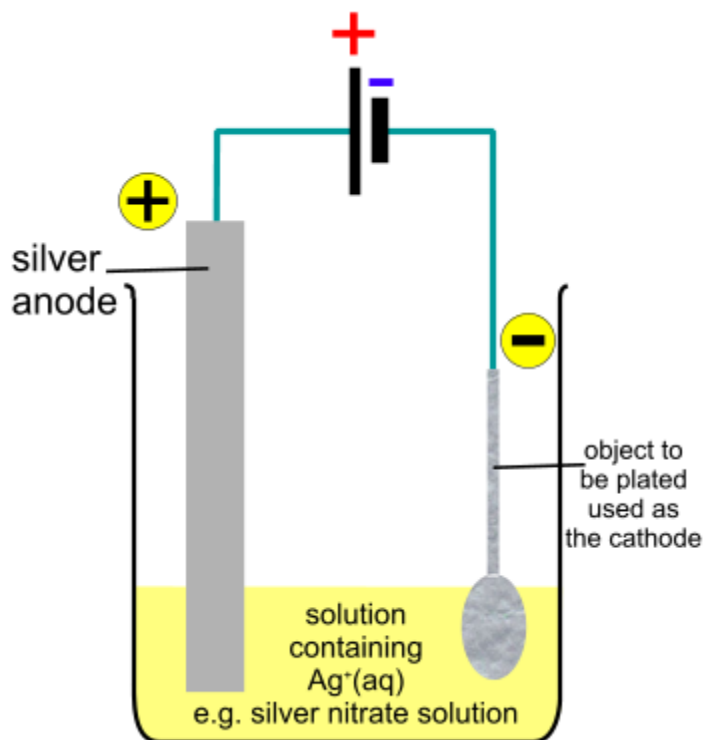
$$\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \quad \text{potential for oxidation} = -0.34 \text{ V}$$

$$\text{H}_2\text{O}(\text{l}) \rightarrow \frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \quad \text{potential for oxidation} = -1.23 \text{ V}$$
- Blue colour of solution stays constant – Cu<sup>2+</sup>(aq) ions are responsible for the blue colour – these are removed from the solution at the cathode but replaced at the anode - the concentration of Cu<sup>2+</sup>(aq) stays constant.
- pH of solution stays constant – concentration of H<sup>+</sup> ions stays constant because water is not oxidized at the anode.

11 State the meaning of the term *electroplating*

Coating an object with a thin layer of metal using electrolysis

12 Draw a labelled diagram of the experimental set-up to silver plate a metal spoon



13 List the factors which affect the amount of product formed during electrolysis.

Current

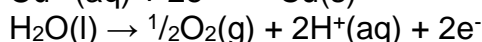
Time of electrolysis

Charge on ion/number of electrons transferred in oxidation/reduction of a species

14 Calculate the relative amounts of products formed during electrolysis

- (a) Copper(II) sulfate solution is electrolysed using platinum electrodes. 0.636 g of copper is produced, calculate the volume of oxygen (measured at STP) collected.
- (b) Dilute sulfuric acid is electrolysed using platinum electrodes. 100.0 cm<sup>3</sup> of hydrogen (measured at STP) is produced, calculate the volume of oxygen (measured at STP) collected.
- (c) A current of x amps is passed for y minutes through a solution containing Cu<sup>2+</sup>(aq) ions and z g of copper were produced. What mass of copper is produced when 2x amps are passed for y/2 minutes?
- (d) Two electrolytic cells are connected together in series. The first contains aqueous copper(II) sulfate and the second contains aqueous lead(II) nitrate. All the electrodes are made of platinum. If 0.16 g of copper are produced in the first cell calculate the mass of lead produced in the second cell.
- (e) How will doubling the current that passes and electrolysing the solutions for 3 times as long affect the mass of copper produced in (d)

- (a) moles of Cu = 0.636/63.55 = 0.0100 mol



Passage of 2 mol e<sup>-</sup> produces 1 mol Cu and 0.5 mol O<sub>2</sub>

Number of moles of O<sub>2</sub> is 0.0100/2 = 5.00x10<sup>-3</sup> mol

Volume of O<sub>2</sub> = 5.00x10<sup>-3</sup> x 22.7 = 0.114 dm<sup>3</sup>

- (b)  $\text{H}_2\text{O}(\text{l}) + \text{e}^{-} \rightarrow \frac{1}{2}\text{H}_2(\text{g}) + \text{OH}^{-}(\text{aq})$   
 $\text{H}_2\text{O}(\text{l}) \rightarrow \frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^{+}(\text{aq}) + 2\text{e}^{-}$  or  $\frac{1}{2}\text{H}_2\text{O}(\text{l}) \rightarrow \frac{1}{4}\text{O}_2(\text{g}) + \text{H}^{+}(\text{aq}) + \text{e}^{-}$

Passage of 1 mol e<sup>-</sup> produces ½ mol H<sub>2</sub> and ¼ mol O<sub>2</sub>

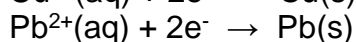
Number of moles of O<sub>2</sub> produced is half the number of moles of H<sub>2</sub>

Volume is proportional to the number of moles, therefore volume of oxygen is 50.0 cm<sup>3</sup>

- (c) 2x current – twice as many electrons flow, therefore twice the number of moles of product  
 y/2 time – half as many electrons flow, therefore half the number of moles of product  
 overall, the mass of Cu will be the same: z g

- (d) Cells in series – same current through each

moles of Cu = 0.16/63.55 = 2.52x10<sup>-3</sup> mol



Same number of moles of lead and copper produced

Mass of lead = 2.52x10<sup>-3</sup> x 207.20 = 0.52 g

- (e) Doubling current – twice as many electrons flow - twice the number of moles of product  
 Three times time – 3 times as many electrons flow - 3 times the number of moles of product  
 3x2 = 6 times the mass of Cu produced      6x0.16 = 0.96 g