FACTFILE: GCSE CHEMISTRY: UNIT 2.7

Rewarding Learning

Electrochemistry

Learning outcomes

Student should be able to:

2.7.1 explain the meaning of the terms electrolysis, inert electrode, anode, cathode and electrolyte and explain conduction in an electrolyte in terms of ions moving and carrying charge;

2.7.2 predict the products of electrolysis of molten salts including lithium chloride and lead(II) bromide using graphite electrodes and state appropriate observations at the electrodes;

2.7.3 **interpret and write half equations for the reactions occurring at the anode and cathode for the electrolysis processes listed in 2.7.2, for other molten halides and in the extraction of aluminium;**

2.7.4 recall the products of electrolysis of dilute sulfuric acid using inert electrodes **and the half equations for the reactions occurring at the anode and the cathode;**

2.7.5 describe the industrial extraction of aluminium from alumina and demonstrate knowledge and understanding that the alumina has been purified from the ore bauxite and demonstrate knowledge and understanding of the need to replace the anodes periodically;

2.7.6 demonstrate knowledge and understanding that recycling aluminium uses only a fraction of the energy needed to extract it from bauxite and saves waste.

Electrolysis

The term **electrolysis** can be split to aid understanding – '*electro*' refers to electricity and '*lysis*' refers to breaking down.

There are a number of key definitions required for this topic:

Electrolysis – the decomposition of a liquid electrolyte using a direct current of electricity.

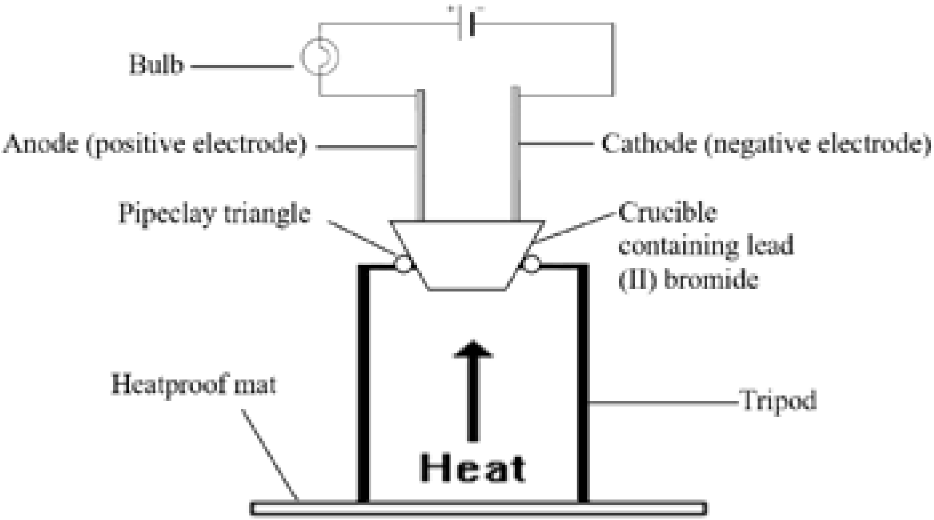
Electrolyte – an ionic compound that is molten or dissolved in water that conducts electricity and is decomposed by it.

Inert electrode – an unreactive material which can conduct electricity usually graphite or platinum.

Cathode – the negative electrode (attached to the negative terminal of the battery).

Anode – the positive electrode (attached to the positive terminal of the battery).

Apparatus for electrolysis of a molten compound



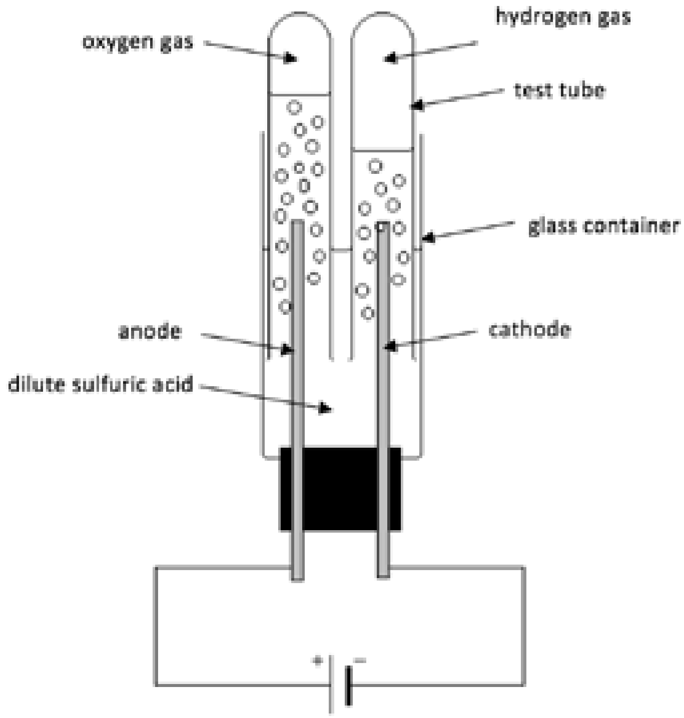
The electrolytes are molten ionic compounds and can conduct electricity as the **ions** are free to move and carry charge. The positive ions (**cations**) are attracted to cathode and gain electrons to form atoms or molecules – gain of electrons is **reduction**. The negative ions (**anions**) are attracted to the anode and lose electrons to form atoms or molecules – loss of electrons is **oxidation**.

The table below summarises the observations and half equations for two named ionic compounds:

MOLTEN IONIC COMPOUND	Lead(II) bromide	Lithium chloride
Anode observation	red-brown vapour pungent smell	yellow-green vapour pungent smell
Anode half equation	$2Br \longrightarrow Br_2 + 2e^-$	$2Cl^- \longrightarrow Cl_2 + 2e^-$
Cathode observation	silver grey metallic bead	silver grey metallic bead
Cathode half equation	$Pb^2 + 2e^- \longrightarrow Pb$	$Li^+ + e^- \longrightarrow Li$

Electrolysis of dilute sulfuric acid

Dilute sulfuric acid conduct electricity as it contains ions. The products of the electrolysis of the acid are gases and can be collected using the apparatus shown below. Inverted test tubes filled with acid are placed over platinum electrodes. The gases produced fill the test tubes.



Dilute sulfuric acid contains hydrogen (H^+) ions from water and sulfuric acid, sulfate (SO_4^{2-}) ions from sulfuric acid and hydroxide (OH^-) ions from water.

Cathode

The hydrogen ions (H^+) are attracted to the negative cathode and gain electrons to form hydrogen gas. Bubbles of gas can be observed on the electrode.

$$2H^+ + 2e^- \rightarrow H_2 \qquad \text{(gain of electrons = REDUCTION)}$$

The gas produced at the cathode will burn with a pop confirming it is hydrogen.

Anode

The negative sulfate ions (SO_4^{2-}) and hydroxide ions (OH^-) are attracted to the positive anode. The sulfate ion is too stable and nothing happens to it, it remains in solution. The hydroxide ions lose electrons to form oxygen gas. Bubbles of gas can be observed on the electrode.

$$4OH^- \rightarrow 2H_2O + O_2 + 4e^- \qquad \text{(loss of electrons = OXIDATION)}$$

The gas produced at the anode will relight a glowing splint confirming it is oxygen.

For every 4 electrons which flow through the circuit 2 molecules of hydrogen and one molecule of oxygen are produced – the gases therefore are produced in a 2:1 ratio, so the volume of hydrogen produced will be twice the volume of oxygen produced.

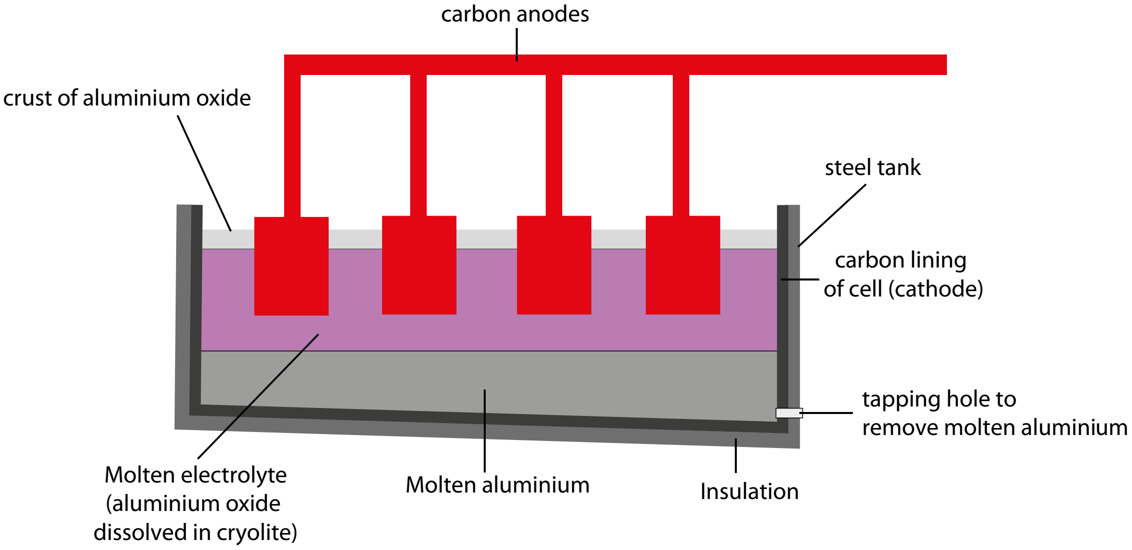
Extraction of aluminium from its ore by electrolysis

Aluminium is one of the more reactive metals and hence its extraction must come from the process of electrolysis. The ore from which aluminium is extracted is called bauxite. Bauxite is a red mineral made of aluminium oxide and iron oxides and is mined in various places such as Russia and Jamaica. The bauxite must be purified before undergoing electrolysis.

The steps in the purification process are:

1. It is heated with sodium hydroxide, this dissolves the alumina but not the other minerals.
2. The residue is filtered off and disposed of.
3. The filtrate is acidified and heated to produce pure alumina, aluminium oxide powder.

The electrolytic cell used in the extraction of aluminium.



The alumina obtained from the purification has a very high melting point around 2050 °C. This is too high a temperature to make the process economically viable so it is dissolved in molten cryolite which reduces the temperature needed to melt the mixture to approximately 950 °C and also increases its electrical conductivity. A crust of aluminium oxide forms on the surface of the electrolyte which prevents heat loss reducing the energy cost of the process.

Reactions at the electrodes.

ANODE – Oxygen is produced at the anode.

- $2O^{2-} \longrightarrow O_2 + 4e^-$

The graphite (carbon) anode reacts with the oxygen produced resulting in them being burnt away over time and hence then requiring periodic replacement.

- $C + O_2 \longrightarrow CO_2$

CATHODE – Aluminium is produced at the cathode.

- $Al^{3+} + 3e^- \longrightarrow Al$

The molten aluminium produced sinks to the bottom of the cell where it can be tapped off.

Recycling of Aluminium

The production of aluminium through electrolysis requires large amounts of bauxite, cryolite and energy. The process by which aluminium is recycled uses only 5% of this energy by comparison making it much more cost effective. Recycling also has environmental benefits associated with it – carbon dioxide emissions are much lower than those produced in the extraction from bauxite and waste quantities requiring landfill disposal are also reduced.

REVISION QUESTIONS

1. Ionic compounds like lead (II) bromide conduct electricity when molten.

   (a) Explain why ionic compounds conduct electricity when molten.

   ____

   ____ [2]

   (b) Draw a labelled diagram of the apparatus used to carry out the electrolysis of molten lead (II) bromide.

   [5]

(c) Define the term electrolysis.

____

____ [1]

(d) Name the material used to make the electrodes in this electrolysis.

____ [1]

(e) What would be observed at the positive electrode in this experiment?

____

____ [2]

(f) Name the products of this electrolysis.

____

____ [2]

(g) Write the half equation for the reaction which takes place at the cathode.

____ [3]

(h) Why would this electrolysis need to be carried out in a fume cupboard?

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____ [2]

2. Aluminium is extracted from pure aluminium oxide by electrolysis.

(a) Name the ore from which aluminium is extracted.

____ [1]

(b) Name the substance added to the aluminium oxide to reduce its melting point.

____ [1]

(c) Write the half equation for the reaction occurring at the cathode.

____ [3]

(d) Write the half equation for the reaction occurring at the anode.

____ [3]

(e) Why does the anode need to be replaced regularly?

____

____ [1]

(f) How is the aluminium removed from the cell?

____ [1]

3. Sulfuric acid can undergo electrolysis.

(a) What material is used for the electrodes in this apparatus?

____ [1]

(b) Name the gas produced at the cathode.

____ [1]

(c) Describe a test that could be used to confirm the presence of this gas.

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____ [2]

(d) Write the half equation for the reaction that takes place at the anode.

____ [3]

(e) How do the volumes of gases produced compare at each electrode?

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____ [2]