Electrolysis

- **Electrolysis** is a process that involves using an electrical current to trigger the decomposition (breaking down) of ionic compounds into simpler substances. To understand electrolysis, it is important to understand the components involved in the process:

1. **Electrolyte** – A substance that can conduct electricity when it is either molten or dissolved in a solution. The electrolyte is the substance which gets decomposed during electrolysis.
2. **Electrode** – A conductor of electricity that is usually a metal or graphite rod. It is connected to the power supply and can be either an anode or a cathode.
3. **Anode** – The positive electrode that is connected to the positive terminal of a battery.
4. **Cathode** – The negative electrode that is connected to the negative terminal of a battery.
5. **Electricity** is the flow of charged particles. Electricity occurring not only through metal wires, some liquids can also transfer electricity, behaving like wires, due to the presence of ions (charged particles) that carry the charge.

How Electrolysis works:

- Electrolysis involves the movement of electrons (negatively charged particles), which is driven by an electrical current. Here's how it works:

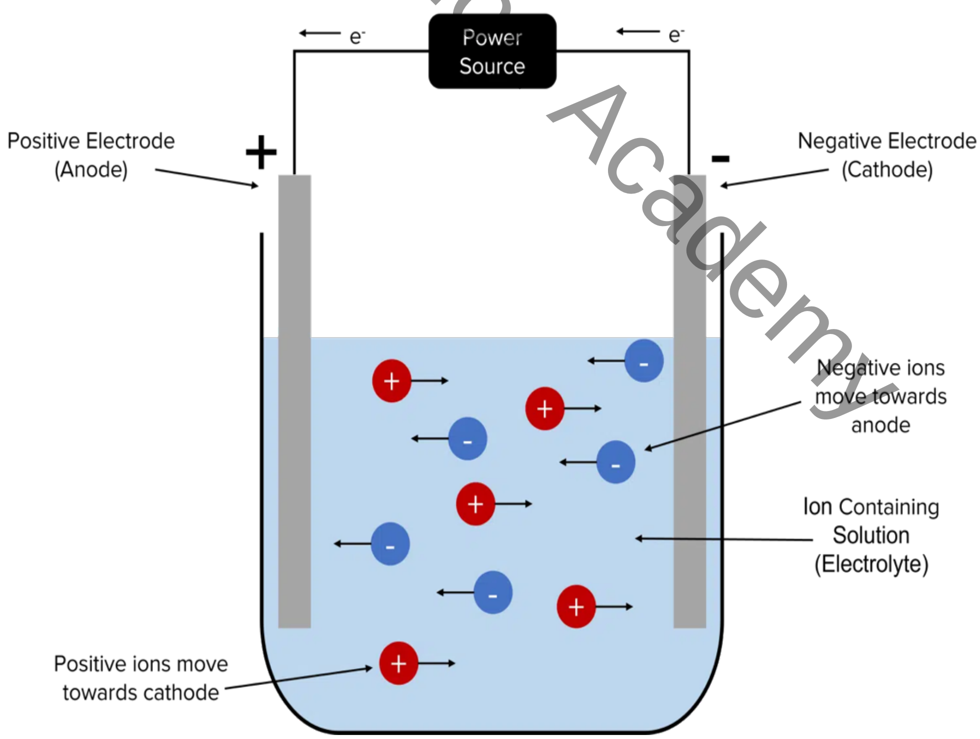
1. The cathode is negatively charged, which attracts positively charged cations. At the cathode, the cations are reduced, which means they gain electrons.

2. The anode is positively charged, which attracts negatively charged anions. At the anode, the anions are oxidised, which means they lose electrons.

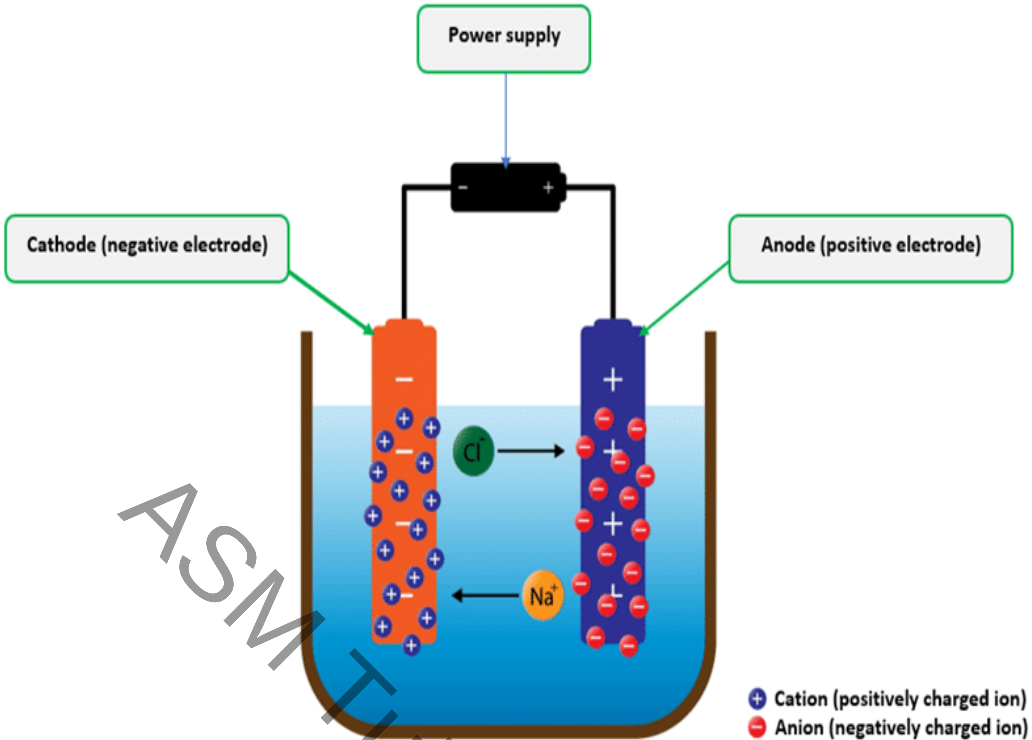
Electrolysis of Molten ionic solids from element

- **Electrolysis** is a chemical process that can be used to **separate ionic compounds** into their constituent ions, when in **solution** or when **molten**, by using an **electrical current** to drive **chemical reactions**.
- Electrolysis makes use of **electrostatic attractions** to pull the ions in a solution in opposite directions.
- Substances **must be in solution or molten** to undergo electrolysis as the ions of the substance need to be able to **move**.

- If the substances were **solid** then the ions within them would be bound in place and unable to separate. Molten ionic compounds can be electrolyse because the ions can move freely and conduct electricity.
- To **electrolyse** an ionic compound a **positive electrode**, known as the **anode**, and a **negative electrode**, known as the **cathode**, should be immersed in a **solution of the compound**, known as the **electrolyte**.
- The electrodes are then connected to an electrical power source, allowing a circuit to be set up. In most cases these electrodes are **inert** (meaning that they do not react). This prevents the **electrolyte** from **reacting with the electrodes**.
- A **current** will flow from the power source to the electrodes, with the **electrolyte carrying charge between them** to complete the circuit. This causes the ions of the electrolyte to **move** towards them, causing the compound to **decompose**.
- The **positive ions** of the electrolyte will be **attracted to the negative charge of the cathode**. This causes the ions to move through the solution towards the **cathode**. The positive ions are then **reduced at the cathode** by electrons flowing in from the external circuit.
- The **negative ions** of the electrolyte will be **attracted to the positive charge of the anode**. This causes the ions to move through the solution towards the **anode**. The negative ions are then **oxidised at the anode**.
- The electrons that are released by this process will then be collected by the cathode to flow into the external circuit.
- Once an ion has been oxidised or reduced it is **discharged** from the solution. Discharged species will often leave the solution either as **solid** sediments or as **gasses**.



- **Example:** if molten sodium chloride is subjected to electrolysis, the positively charged sodium ions (Na+) are attracted to the cathode, where they gain electrons. Meanwhile, the negatively charged chloride ions (Cl–) are attracted to the anode, where they lose electrons.



Metal Extraction

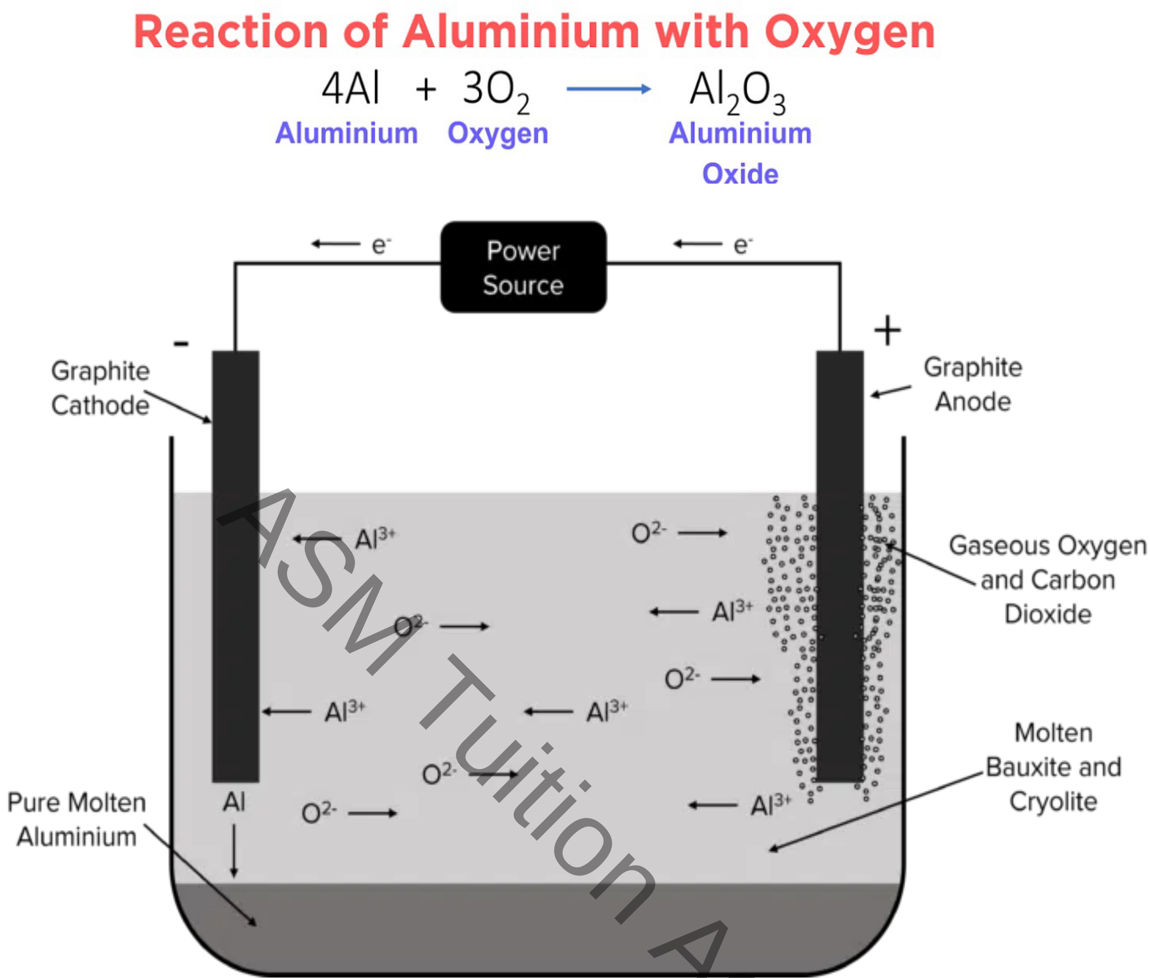
- One useful application of electrolysis is the **extraction of metals** from their ores when these metals are **too reactive** to be **reduced by carbon**.
- Take for example the extraction of **aluminium** from its ore **bauxite**.
- has the chemical formula Al2 O3 and can be **electrolysed when molten**. The melting point of bauxite is very high, so it is mixed with another compound called cryolite which reduces bauxites melting point.
- Once molten, the bauxite is then electrolysed using a pair of **graphite** electrodes.
- At **positive Electrode** Al3+ **ions** are attracted to the **negative graphite cathode**. Here they are **reduced** by electrons flowing in from the external power source. This produces **molten elemental** Al. This molten aluminium is **heavier** than the mixture of bauxite and cryolite and **falls to the bottom** ,of the mixture where it can be collected.

Al3+(aq) + 3e- → Al(s)…… Reduction: gain of electrons

- At **negative Electrode** O2− **ions** are attracted to the **positive graphite anode**. Here they are **oxidised** to form **gaseous** O2O2 , giving up their electrons to the external circuit. The O2− ions may also react with the carbon of the graphite anode to produce gaseous CO2

$2O^{2-} \rightarrow O_2 + 4e^-$ …….. Oxidation: loss of electrons

Over All Equation:



Electrolysis of Aqueous Solutions

- When **aqueous solutions** are used as electrolytes the picture becomes more complicated than that of the electrolysis of molten ionic compounds.
- When **molten compounds** are electrolysed the **only ions** produced are **those of the molten compound**.
- When **aqueous solutions** are electrolysed, the **water** in the solution will **also electrolyse** to form H+ **and** OH−ions in the following reversible reaction.

$$H2O(l) \rightleftharpoons H\ \ (aq)++OH\ \ (aq$$

Ion Production with Aqueous Electrolytes

When **aqueous electrolytes** are involved, electrolysis is slightly more complicated.

I. **Two ions hydrogen ion** and the **metal ion** are attracted to the cathode. To decide what substance is made at the cathode, we must use the reactivity series.

II. **Metal reactivity is compared to hydrogen.** If the metal element formed during electrolysis is **more** reactive than hydrogen, then hydrogen will be produced at the cathode. If the metal element formed is **less** reactive than hydrogen, then the metal is produced at the cathode.

III. **Two ions hydroxide ions** and **halide ions** are attracted to the anode. If both are present, then a halogen is formed**.** If no halide ions are present, then the hydroxide ions are discharged, and oxygen is formed**.**

Examples of Aqueous Electrolysis

Example 1: Write out the half equations for the electrolysis of copper sulfate.

1.Write out all the ions. before you start the ionic equation.

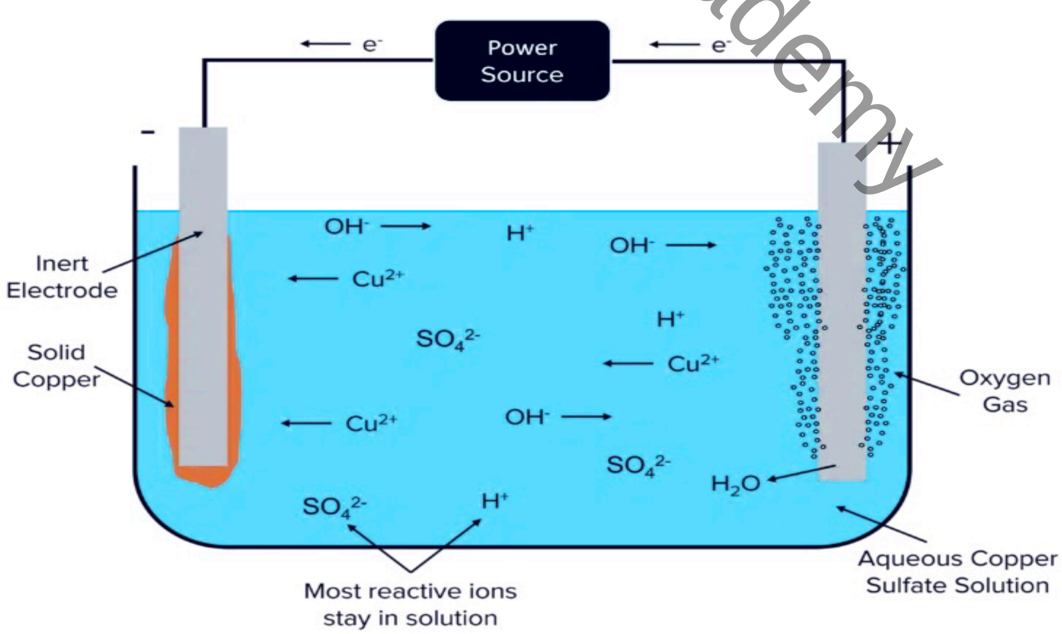
Cu2+, SO42-, H+, OH–

2. Look at the cathode. We can form either Copper or Hydrogen. Copper is less reactive than hydrogen on the reactivity series, so Copper is formed**,** coating the electrode.

Cu2+ + 2e– → Cu

3. Look at the anode. No halide ions are present, so hydroxide ions are discharged, forming water and oxygen.

4OH– → O2 + 2H2O + 4e–



Example 2: Write out the half equations for the electrolysis of sodium chloride.

1.Write out all the ions.

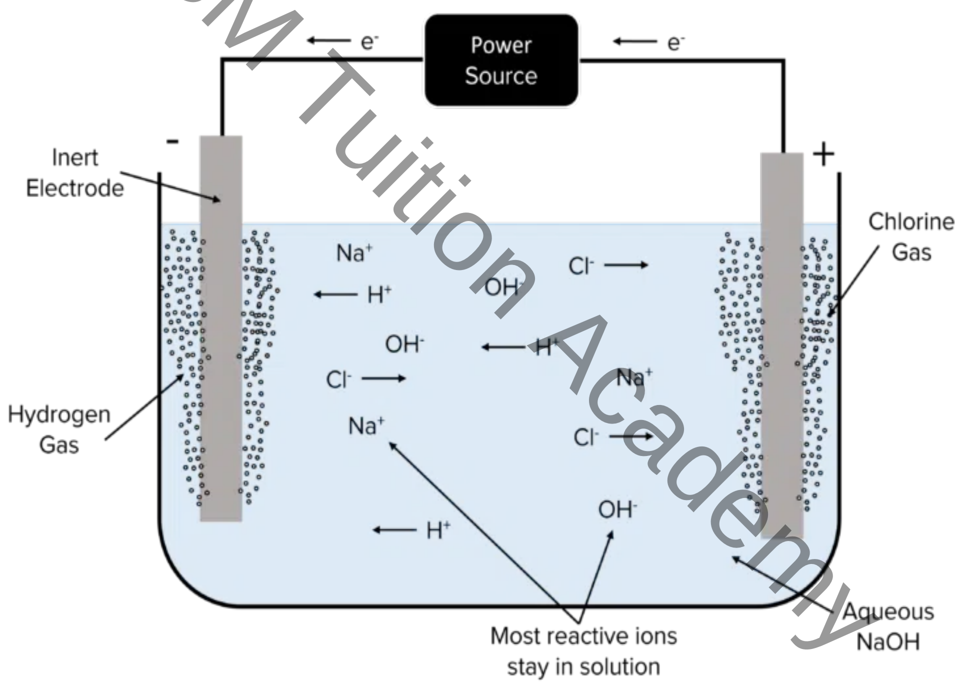
Na^+, Cl^-, H^+, OH^-

2. Look at the cathode. Either sodium or hydrogen can be formed. Since sodium is more reactive than hydrogen, hydrogen is made.

$$2H^+ + 2e^- \rightarrow H_2$$

3. Look at the anode. Anode: either chlorine or hydroxide ions, however halides take precedence, so chlorine is made.

$$2Cl^- \rightarrow Cl_2 + 2e^-$$



Electrode Half Equations

- Electrolysis is the presence of **chemical reactions at each electrode**. These reactions take the form of **redox** half reactions, that contribute to an **overall electrolytic reaction**.
- For example the extraction of aluminium from its oxides. In this process there is an overall decomposition reaction, where by bauxite breaks down into pure aluminium and oxygen:

$$2Al_2O_3(l) \rightarrow 4Al(l) + 3O_2(g)$$

- However, the break down of bauxite by electrolysis does not directly produce either of these products. Instead, ions of aluminium and oxygen are produced and then reduced or oxidised at the electrodes.
- The aluminium ions (Al^{3+}) are **reduced at the cathode** to form pure aluminium. It is possible to write a **half equation** for this reduction to show the formation of elemental aluminium (the term e^- is used to represent the electrons).

$$Al^{3+}(l) + 3e^- \longrightarrow Al\ (l)$$

- The O^{2-} ions are oxidised at the anode to form oxygen gas. It is again possible to write a half equation for this process.

$$2O^{2-}\ (g) \rightarrow O_2(g) + 4e$$

- We can double check that a half equation is correct by checking if **charge is balanced** across the equation. The **charges should be equal on both sides** of the reaction (e.g. above there is a total charge of −4 on the left and a total charge of −4 on the right).