20. Electrochemistry

There are two areas of electrochemistry:

1. Using chemical reactions to produce an electric current.

2. Using electric currents to carry out chemical reactions.

An electrochemical cell is a device that converts chemical energy into electrical energy or electrical energy into chemical energy through redox reactions.

1. **Galvanic cell:** A cell in which a chemical reaction results in the production of an electric current (a flow of charge).

2. **Electrolytic cell:** A cell that uses electricity to drive a chemical reaction.

Galvanic
chemical → electrical
chemical ← electrical
Electrolytic

Galvanic Cells

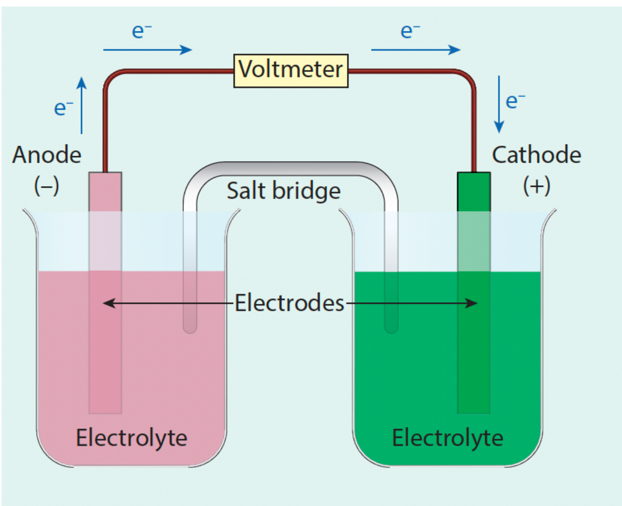
A cell in which a chemical reaction results in the production of an electric current is called a **galvanic cell.**

- Spontaneous redox reactions take place and chemical energy is converted into electrical energy.
- Galvanic cells have two compartments, each called a **half-cell**. Each half-cell contains an **electrode** and an **electrolyte.**
- Electrodes are electrical conductors.
- Electricity flows through electrodes into or out of a medium that conducts electricity (e.g. an electrolyte).

An electrolyte is a substance that will conduct an electric current when molten or dissolved.

- In one half-cell, oxidation occurs; in the other, reduction. Together, the two half-cells carry out a **redox reaction.**
- A wire connects the two electrodes so the **electrons can flow from one half-cell to the other.** This flow of electrons is an **electric current**.
- Oxidation occurs at the anode, which (in a galvanic cell) is the negative electrode.
- Reduction occurs at the cathode, which (in a galvanic cell) is the positive electrode.

Anode
e^- move **a**way



e^- **c**ome
to the **c**athode

- The solutions in each half-cell must **remain electrically neutral** for a galvanic cell to work.
- To do this, a **salt bridge** is used. Examples of salts used include KCl and KNO_3.
- **The cations and anions from the salt replace the positive and negative ions being lost in each half-cell and maintain a neutral charge overall.**

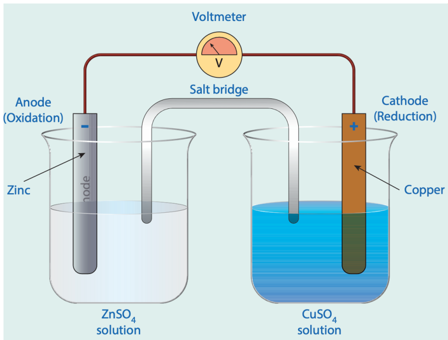
Experiment 19: Create a simple copper-zinc galvanic cell

- Chemical → Electrical (a chemical reaction occurs, resulting in an electric current)
- Since zinc is above copper in the electrochemical series, zinc is oxidised and copper is reduced.

Half-cell 1

Electrolyte = solution of zinc sulfate $ZnSO_4$
Anode (negatively charged) = zinc electrode
Oxidation occurs: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

Each zinc atom loses two electrons, which flow through the wire from the zinc electrode to the copper electrode. Flow of e^- = electric current



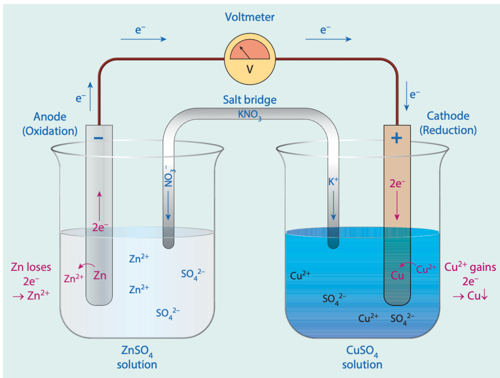
Half-cell 2

Electrolyte = solution of copper sulfate ($CuSO_4$)
Cathode (positively charged) = copper electrode
Reduction occurs: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu\downarrow_{(s)}$

Overall: $\mathbf{Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu\downarrow_{(s)}}$

As a result of these redox reactions, the zinc solution would become very positive, and the copper solution would become very negative.

If these charges build up, the cell wouldn't be able to continue producing an electric current.



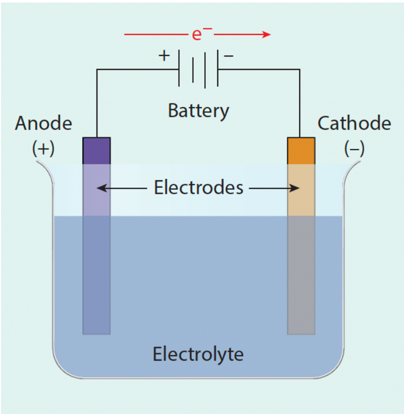
The salt bridge is filled with a KNO_3 solution to maintain a neutral charge overall.

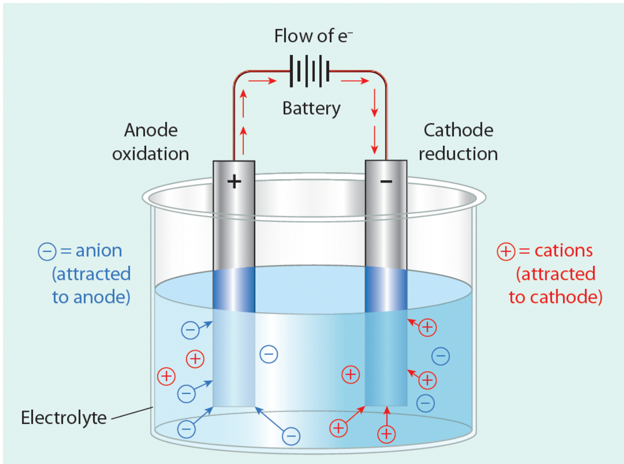
Electrolytic cells

Electrolysis is the use of electricity to drive a chemical reaction.

- Electrolysis is carried out in **electrolytic cells.**
- Electrolytic cells use electrical energy from a battery or power supply to force a chemical reaction to occur.
- An electrolytic cell consists of two electrodes placed in an electrolyte solution.
- A battery or power supply is connected to the electrodes. This pushes electrons onto one electrode (making it negatively charged) and pulls electrons from the other (making it positively charged).
- In an electrolytic cell the cathode is negative and the anode is positive (CNAP).
- Chemical reactions occur at each electrode.

- The positive ions in solution are attracted to the cathode and gain electrons: reduction occurs.
- The negative ions in solution are attracted to the anode and lose electrons: oxidation occurs,
- Overall, a redox reaction occurs.





- Usually, the electrodes are inert (unreactive) electrodes. Common materials for inert electrodes are graphite and platinum.

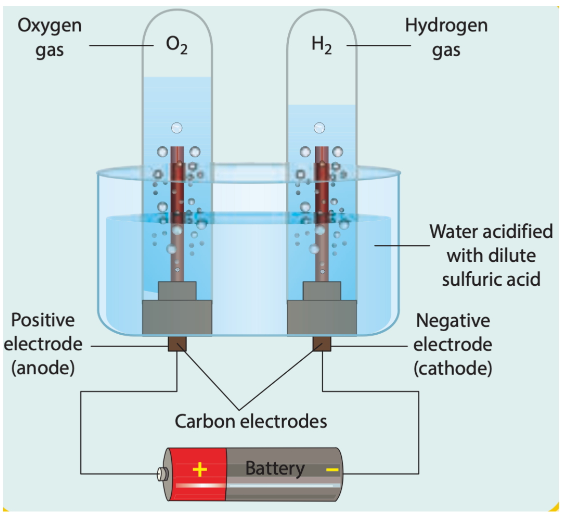
Experiment 20: Electrolysis of acidified water

Uses electrical energy to drive a non-spontaneous chemical reaction.

An electrolytic cell with acidified water and inert electrodes (e.g. platinum) is set up.
- Pure water is a very weak electrolyte (since it ionises to a very small extent).
- Sulfuric acid is added to increase the conductivity (H_2SO_4 is a source of H^+).
- An electric current is passed through acidified water, and oxidation and reduction reactions occur.
- A voltameter with inverted test tubes can be used, or a Hoffman voltameter.

<u>**Anode (positive electrode)**</u>

The water molecules lose electrons to form oxygen gas and hydrogen ions. Bubbles are
observed oxygen is collected.
Oxidation occurs: $H_2O \rightarrow ½O_2 + 2H^+ + 2e^-$



<u>**Cathode (negative electrode)**</u>

The positive H^+ are attracted to the cathode. The H^+ ions in solution gain electrons to form
atoms. Two atoms combine to form a molecule of hydrogen.
Reduction occurs: $2H^+ + 2e^- \rightarrow H2$

Overall: $H_2O_{(l)} \rightarrow H_{2(g)} + ½O_{2(g)}$

Observations

The volume of gas collected at the cathode (negative electrode) is twice the volume of the gas collected at the anode (positive electrode). (This makes sense when we look at the ratio of hydrogen to oxygen in water, H_2O, which is 2:1.)

Gas at the anode relights a glowing splint = oxygen
Gas at the cathode burns with a 'pop!' = hydrogen

	Galvanic cell	Electrolytic cell
Composed of	Two half-cells each containing an electrode and electrolyte solution, connected by a wire and salt bridge	Two electrodes placed in the same electrolyte solution in the same container A battery or power supply connected by wires to the electrodes
Use	Chemical reactions to produce an electric current	An electric current to drive a chemical reaction
Energy conversion	Chemical → electrical	Electrical → chemical
Spontaneity	Spontaneous reaction	Non-spontaneous reaction
Anode	Negative Oxidation occurs	Positive Oxidation occurs
Cathode	Positive Reduction occurs	Negative Reduction occurs
Uses	Used in AA batteries (what we know as an 'AA battery' in everyday life is, in fact, a galvanic cell)	Electroplating (using an electrical current to deposit a thin layer of metal onto the surface of another material – e.g. a nickel spoon is coated with silver)

Hydrogen fuel cells

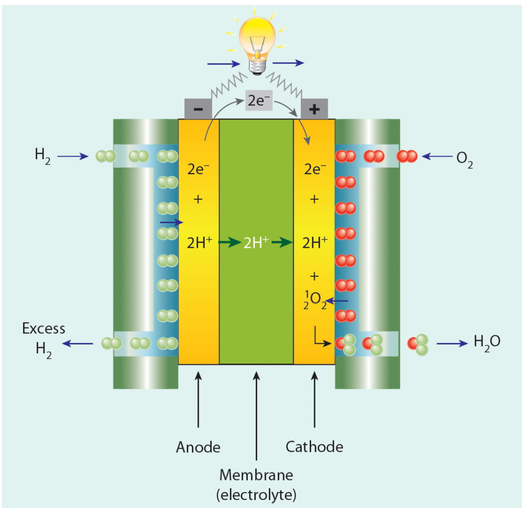
- A hydrogen fuel cell converts chemical energy to electrical energy.
- A fuel cell is different from a chemical cell as it has **constant supply of fuel**, with which it **continuously produces electricity**.
- Fuel = hydrogen.
- Oxidising agent = oxygen.
- A simple hydrogen fuel cell consists of an anode, a cathode, an electrolyte and a catalyst (often platinum).
- The electrolyte allows protons to pass through it, but not electrons.

Step 1: Oxidation of hydrogen

Hydrogen molecules (H_2) split into protons (H^+) and electrons (e^-). This happens in the presence of a catalyst – e.g. platinum.

Step 2: Generation of electricity

The protons pass through the electrolyte from the anode to the cathode (through the electrolyte). However, the **electrons are forced to flow through an external circuit**, creating an electric current.



Step 3: Reduction of oxygen, formation of water
When the electrons flow through the external circuit and reach the cathode, they combine with oxygen (supplied) and the protons (H^+) that have passed through the electrolyte, forming water.

Anode: $2H_2 \rightarrow 4H^+ + 4e^-$
Cathode: $4H^+ + 4e^- + O_2 \rightarrow 2H_2O$
Overall: $2H_2 + O_2 \rightarrow 2H_2O$ + electricity + heat

Advantages of HFCs

- **Environmentally friendly** (if produced without fossil fuels): no greenhouse gas emissions.
- The only by-products are water (which can be used elsewhere or in continuous loop production) and heat.
- **Continuous production** of electricity (as long as the supply of fuel is continuous).
- **High efficiency:** HFCs are more efficient than traditional power-generation methods – e.g. in vehicles, hydrogen fuel cells are more efficient (approximately 50%) than traditional combustion engines.
- **Renewable** source of energy.

Challenges – HFCs

- **Cost** (platinum catalyst used = expensive!)
- **Environmental impact** (only environmentally friendly if hydrogen used is not generated using fossil fuels)
- **Infrastructure** (currently a lack of infrastructure for hydrogen production, storage and distribution)
- **Safety** (hydrogen gas = flammable)
- **Storage and distribution**

Uses – HFCs

- **Transportation** (cars, buses, etc.)
- **Stationary power generation** (used to generate electricity)
- **Portable power** (mobile phones, laptops, cameras, etc.)