2024 IP4 Chemistry

Topic 11: Electrochemistry Notes

TEMASEK JUNIOR COLLEGE

Learning objectives

At the end of this notes, you should be able to:

- ☐ **(a)** Describe electrolysis as the conduction of electricity through an ionic compound (an electrolyte), when molten or dissolved in water, leading to chemical changes (including decomposition) at the electrodes
- ☐ **(b)** Describe electrolysis as evidence for the existence of ions which are held in a lattice when solid but which are free to move when molten or in solution
- ☐ **(c)** Describe, in terms of the mobility of ions present and the electrode products, the electrolysis of molten sodium chloride, using inert electrodes
- ☐ **(d)** Predict the likely products of the electrolysis of a molten binary ionic compound using inert electrodes
- ☐ **(e)** Apply the idea of selective discharge based on
  - (i) cations: linked to the reactivity series
  - (ii) anions: halides, hydroxides and sulfates (e.g. aqueous copper(II) sulfate and dilute sodium chloride solution (as essentially the electrolysis of water))
  - (iii) concentration effects (as in the electrolysis of concentrated and dilute aqueous sodium chloride).(In all cases above, inert electrodes are used.)
- ☐ **(f)** Predict the likely products of the electrolysis of an aqueous electrolyte, given relevant information
- ☐ **(g)** Construct ionic equations for the reactions occurring at the electrodes during the electrolysis, given relevant information
- ☐ **(h)** Describe the electrolysis of aqueous copper(II) sulfate with copper electrodes as a means of purifying copper (no technical details are required)
- ☐ **(i)** Describe the electroplating of metals, e.g. copper plating, and state one use of electroplating
- ☐ **(j)** Describe the production of electrical energy from simple cells (from two electrodes in an electrolyte) linked to the reactivity series and redox reactions (in terms of electron transfer)
- ☐ **(k)** Describe hydrogen, derived from water or hydrocarbons, as a potential fuel, reacting with oxygen to generate electricity directly in a fuel cell (details of the construction and operation of a fuel cell are not required)

Topic Outline

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References:

Chemistry Matters	*Tan Yin Toon & Chen Ling Kwong (pg 261-282)*
Chemistry Matters Workbook	*Tan Yin Toon & Chen Ling Kwong (pg 94-102)*

1 Introduction to Electrolysis

1.1 Definition

- **Electrolysis** is the process of using **electricity** to break down or decompose a compound.
- Electrolysis takes place in an electrolytic cell, which consists of a battery, electrodes and electrolyte.

1.2 Components of an electrolytic cell

1.2.1 Electrolyte

- **Electrolytes** are **molten (liquid) compounds or aqueous solutions** which can **conduct electricity**.
- In molten or aqueous state, there are **free mobile ions** that can act as charge carriers to conduct electricity.

  Molten ionic compound: $NaCl(l) \rightarrow \mathbf{Na^+(l) + Cl^-(l)}$

  Aqueous solution: $HCl(aq) \rightarrow \mathbf{H^+(aq) + Cl^-(aq)}$

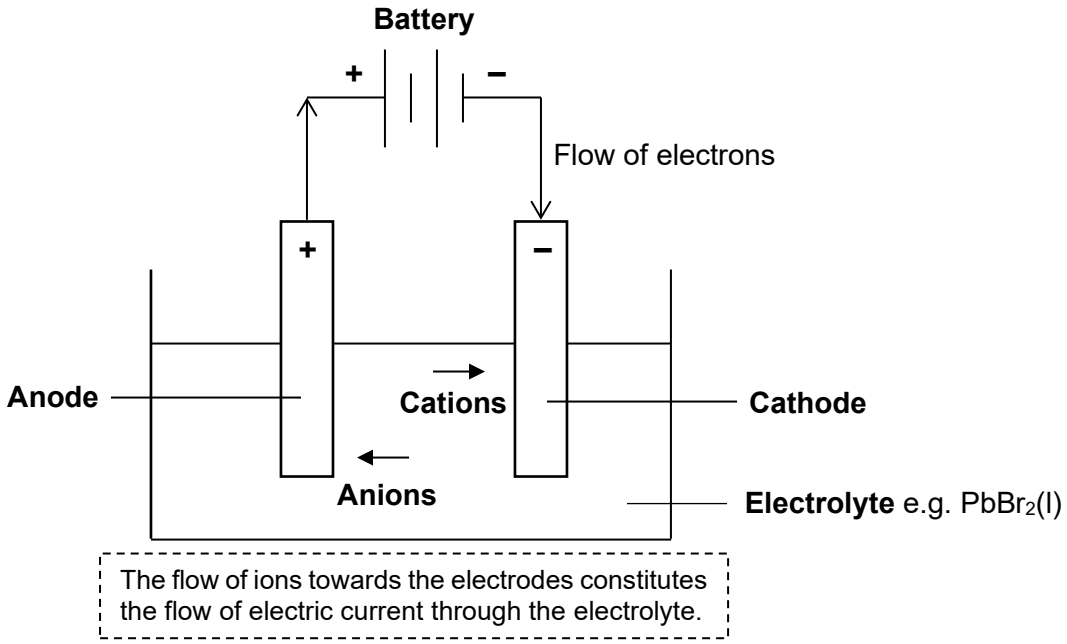
1.2.2 Electrodes

- **Electrodes** are **metal plates or carbon (graphite) rods** which **can conduct electricity**.
- Inert electrodes are often used as they do not take part in any chemical reaction during electrolysis. They provide the surface for electron transfer to occur during electrolytic reactions.
- Carbon and platinum (Pt) electrodes are considered to be inert electrodes.

1.2.3 Set-up of an Electrolytic cell

- **Anode** is the electrode where **oxidation** occurs. Anode is the positive electrode connected to the positive terminal of the battery.
- **Cathode** is the electrode where **reduction** occurs. Cathode is the negative electrode connected to the negative terminal of the battery.
- The process of gaining or losing electrons at the electrodes is called **discharge**.

Set-up of Electrolytic Cell – *Learn how to draw and label*



(a) Anode

- **Positive** electrode (attached to **positive** terminal of battery)
- **Anions** migrate to anode
- oxid**A**tion of **A**nions occurs

Oxidation half-equation: $\mathbf{2Br^-(l) \rightarrow Br_2(g) + 2e^-}$ ----- (1)

(b) Cathode

- **Negative** electrode (attached to **negative** terminal of battery)
- **Cations** migrate to cathode
- redu**C**tion of **C**ations occurs

Reduction half-equation: $\mathbf{Pb^{2+}(l) + 2e^- \rightarrow Pb(l)}$ ----- (2)

(c) Overall equation

The overall equation is (1) + (2): $\mathbf{PbBr_2(l) \rightarrow Pb(l) + Br_2(g)}$

Note that electrons are only present in the half equations. The number of electrons lost at the anode is equal to the number of electrons gained at the cathode. Electrons do not appear in the overall equation.

2 Electrolysis of Molten Ionic Compounds

- When solid ionic compounds are heated strongly, they melt to give the molten form. Molten ionic compounds contain mobile ions to act as charge carriers and can hence act as electrolytes.

- Binary ionic compounds contain only two elements to form the metal cation and non-metal anion. An example of binary compounds is sodium chloride.

- The electrolysis of a molten binary ionic compound is straightforward since the electrolyte contains only type of one cation and one anion.

Watch molten sodium chloride conduct electricity! Click here or scan the QR code or key in the link: https://www.youtube.com/watch?v=d3SDxNhPhtE

2.1 Electrolysis of molten sodium chloride using inert electrode

Ions present in electrolyte: **Na^+(l) and Cl^-(l)**

At the anode:

- Cl^-(l) migrates to the anode
- Cl^-(l) is oxidised to form Cl_2(g)
- Oxidation half-equation: **$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$**
- Observation: **greenish-yellow fumes** of chlorine are produced at the anode.

At the cathode:

- Na^+(l) migrates to the cathode
- Na^+(l) is reduced to form Na(l)
- Reduction half-equation: **$Na^+(l) + e^- \rightarrow Na(l)$**
- Observation: **Grey sodium metal** is produced at the cathode

Overall equation: **$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$**

3 Electrolysis of Aqueous Ionic Compounds

3.1 Electrolysis of water

- An aqueous solution of an ionic compound contains cations, anions and water molecules, H_2O.
- The electrolysis is not as straightforward since water may be involved in the process.
- The equations for oxidation and reduction of water are shown below:

Oxidation of water at the anode:

$$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

- H_2O is oxidised as the oxidation state of O increases from –2 in H_2O to 0 in O_2.
- Since H^+ ions are produced, the solution around the anode becomes more **acidic**.

Reduction of water at the cathode:

$$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$$

- H_2O is reduced as the oxidation state of H decreases from +1 in H_2O to 0 in H_2.
- Since OH^- are produced, the solution around the cathode becomes more **alkaline**.

Watch the electrolysis of water into hydrogen and oxygen gases!
Click here or scan the QR code or key in the link:
https://www.youtube.com/watch?v=OTEX38bQ-2w

3.2 Electrolysis of dilute sodium chloride using inert electrode

Species present: **$Na^+(aq)$, $Cl^-(aq)$, $H_2O(l)$**

At the anode:

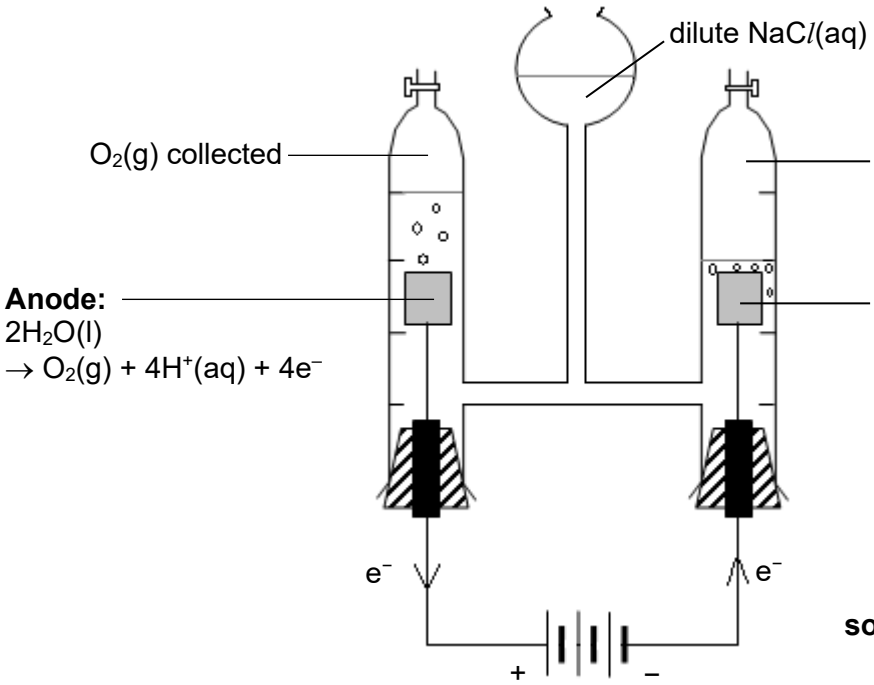
- **Cl^- and H_2O** migrate to the anode
- **H_2O** is preferentially discharged (oxidised) to form **O_2**
- Oxidation half-equation: **$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$**
- Observation: Effervescence of **O_2 gas** observed that relights a glowing splint.

At the cathode:

- **Na^+ and H_2O** migrate to the cathode
- **H_2O** is preferentially discharged (reduced) to form **H_2**
- Reduction half-equation: **$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$**
- Observation: Effervescence of **H_2 gas** observed that extinguishes a lighted splint with a 'pop' sound

Overall equation: **$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$**

Water is preferentially discharged at both electrodes. The overall equation shows that the electrolysis of dilute sodium chloride is *equivalent* to the electrolysis of water. There are 2 volumes of hydrogen gas produced at the cathode compared to 1 volume of oxygen gas produced at the anode. Since water is discharged, the concentration of sodium chloride gradually increases.



Electrolysis of dilute sodium chloride solution

3.3 Electrolysis of concentrated sodium chloride (brine) using inert electrode

***Concept*:**	**The higher the concentration of ions, the easier it will be discharged**

Species present: $Na^+(aq)$, $Cl^-(aq)$, $H_2O(l)$

Watch the electrolysis of brine into hydrogen and chlorine gases! Click here or scan the QR code or key in the link: https://www.youtube.com/watch?v=JwznB5obwiU

At the anode:

- Cl^- and H_2O migrate to the anode
- Cl^- is preferentially oxidised to form Cl_2.

  Reason: There is a much **higher** concentration of Cl^- ions compared to H_2O molecules.
- Oxidation half-equation: **$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$**
- Observation: **greenish-yellow fumes** of chlorine are produced at the anode.

At the cathode:

- Na^+ and H_2O migrate to the cathode
- H_2O is preferentially reduced to give H_2.

  Reason: While there is a high concentration of Na^+ at the cathode, Na^+ is unlikely to gain an electron to give the very reactive sodium metal. Hence, H_2O is still preferentially discharged to form H_2.
- Reduction half-equation: **$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$**
- Observation: Effervescence of H_2 gas observed that extinguishes a lighted splint with a 'pop' sound

Summary of Electrolysis of NaCl

	Reaction at anode	Reaction at cathode
molten NaCl(l)	**$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$**	**$Na^+(l) + e^- \rightarrow Na(l)$**
dilute NaCl(aq)	**$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$**	**$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$**
concentrated NaCl(aq)	**$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$**	**$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$**

3.4 Identification test for gases

Gas	Test for Gas
O_2	O_2 gas relights a glowing splint.
H_2	H_2 gas extinguishes a lighted splint with a 'pop' sound
Cl_2	Greenish-yellow Cl_2 gas turns moist blue litmus paper red then bleaches it

4 Ease of Discharge

4.1 Ease of discharge of cations

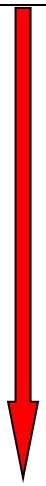
Concept: Cation of less reactive elements are more easily discharged

Reactivity series of metal decreases from K to Ag:

K *most reactive metal*
Na
Ca
Mg
Al
Zn
Fe
Pb
H
Cu
Ag
Au *least reactive metal*

Ease of discharge of cations **increases** from K^+ to Ag^+:

K^+ *least easily discharged*
Na^+
Ca^{2+}
Mg^{2+}
Al^{3+}
H_2O
Zn^{2+}
Fe^{2+}
Pb^{2+}
H^+
Cu^{2+}
Ag^+
Au^+ *most easily discharged*



For example, if solution contains both:

(i) K^+ and H^+ ⇒ H^+ ions will be preferentially discharged. Hence, hydrogen gas is produced at the cathode.

Reduction half-equation: **$2H^+(aq) + 2e^- \rightarrow H_2(g)$**

(ii) Cu^{2+} and H^+ ⇒ Cu^{2+} ions will be preferentially discharged. Hence, copper metal is produced at the cathode.

Reduction half-equation: **$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$**

Note: Very reactive metals, such as potassium and sodium, are obtained from their ores by **electrolysis of a molten compound**.

4.2 Ease of discharge of anions

Concept:	The **ease of discharge of anions** depends on the **concentration of the anion** as **well as the order shown** below.

Ease of discharge of anions **increases** from Cl^- to OH^-:

Cl^-
H_2O
Br^-
I^-
OH^- *most easily discharged*

SO_4^{2-} and NO_3^- are not oxidised / discharged during electrolysis. This is because; the oxidation state of N in NO_3^- is already at the maximum of +5, and the oxidation state of S in SO_4^{2-} is at a maximum of +6. It is unlikely for NO_3^- and SO_4^{2-} to be further oxidised.

The ease of discharge of anions also depends on the relative concentrations of the anions. A higher concentration of an anion tends to promote its discharge.

Example:

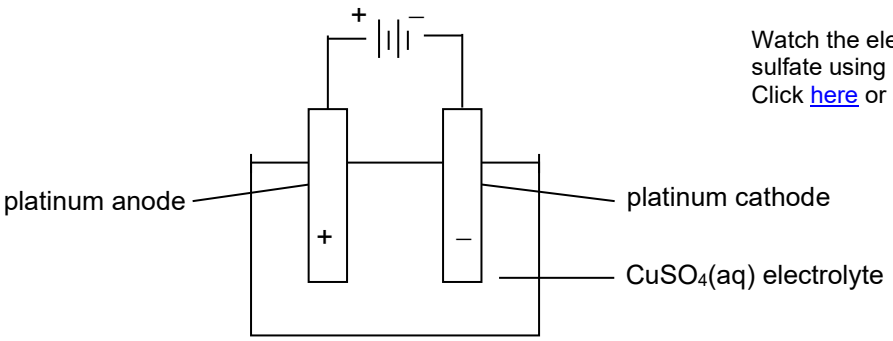
Solution containing *same concentrations* of Cl^- and OH^-	**OH^-** will be preferentially discharged *Explanation : Follow the order in the anion series*
Solution containing *higher concentrations of Cl^-* compared to OH^-	**Cl^-** will be preferentially discharged *Explanation: **Cl^- is preferentially discharged as Cl^- has a higher concentration***

5 Nature of the Electrodes

- Electrodes may be **inert or reactive**.
- An **inert electrode** is an **unreactive** electrode that **cannot be oxidised**.
  Examples of inert electrodes are graphite and platinum electrodes.
- A **reactive electrode** is an electrode that **can be oxidised** in electrolysis.
  Examples of reactive electrodes are some metals such as silver, copper.

5.1 Electrolysis of copper(II) sulfate solution using inert electrodes (such as Pt)

Copper(II) sulfate can be electrolysed using inert platinum electrodes.



Watch the electrolysis of copper(II) sulfate using graphite electrodes! Click here or scan the QR code.

At the **Pt** anode:

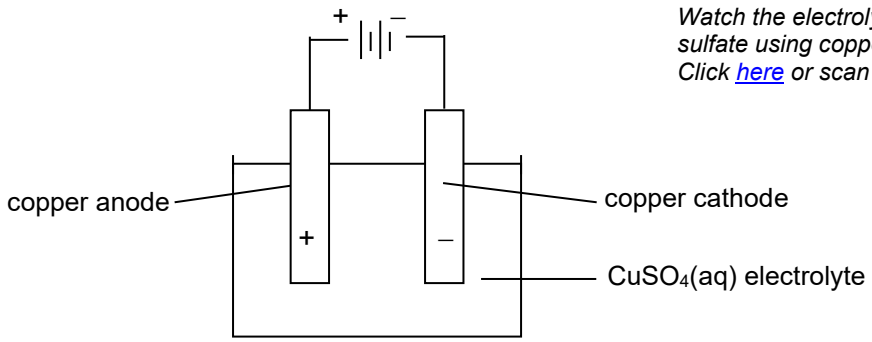
- SO_4^{2-} and H_2O migrate to the anode
- **H_2O** is preferentially discharged (oxidised) to form **O_2**
- Oxidation half-equation: **$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$**
- Observation: **Effervescence of O_2 gas** observed that **relights a glowing splint**

At the **Pt** cathode:

- Cu^{2+} and H_2O migrate to the cathode
- **Cu^{2+}** is preferentially discharged (reduced) to form **Cu**
- Reduction half-equation: **$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$**
- Observations:
  (1) Decolourisation of the blue solution of Cu^{2+} as concentration of Cu^{2+} decreases.
  (2) Pink Cu metal deposited at the cathode.
  (3) Mass of cathode increases since Cu metal is deposited.

5.2 Electrolysis of copper(II) sulfate solution using copper (reactive) electrodes

Using **reactive copper anode**, we need to consider the oxidation of copper too. The copper anode is more preferentially oxidised. Therefore the copper anode gets thinner as electrolysis proceeds.



Watch the electrolysis of copper(II) sulfate using copper electrodes! Click here or scan the QR code.

At the **Cu** anode:

- SO_4^{2-} and H_2O migrate to the anode
- **Cu anode** is preferentially discharged (oxidised) to form **Cu^{2+}(aq)**
- Oxidation half-equation: **$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$**
  anode
- Observation: **The anode becomes thinner or mass of anode decreases**

At the **Cu** cathode:

- Cu^{2+} and H_2O migrate to the cathode
- **Cu^{2+}** is preferentially discharged (reduced) to form **Cu**
- Reduction half-equation: **$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$**
- Observation:
  (1) **Pink Cu metal** deposited at the cathode.
  (2) **Mass of cathode increases** since Cu metal is deposited.

There is no change in the blue colour of the solution as the **concentration of Cu^{2+}(aq)** remains constant. The net result is the transfer of copper from the anode to the cathode.

Summary of the electrolysis of $CuSO_4$(aq)

	Reaction at anode	Reaction at cathode
Inert Pt electrodes	$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
Reactive Cu electrodes	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Quick-check

Which of the following correctly represents the reactions occurring at both inert electrodes in the electrolysis of the following electrolytes?

	Electrolyte	Reaction at cathode	Reaction at anode
A	$AgNO_3(aq)$	$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$	$Ag^+ + e^- \rightarrow Ag$
B	$NaOH(aq)$	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$
C	$HNO_3(aq)$	$2H^+ + 2e^- \rightarrow H_2$	$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$
D	$MgSO_4(aq)$	$Mg^{2+} + 2e^- \rightarrow Mg$	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

Option A **Cathode** : $Ag^+ + e^- \rightarrow Ag$
Anode : $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

Option C **Cathode** : $2H^+ + 2e^- \rightarrow H_2$
Anode : $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

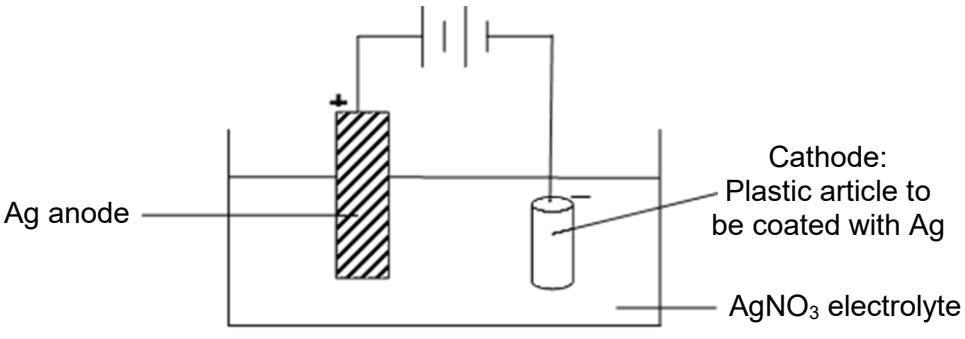
Option D **Cathode** : $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
Anode : $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

6 Applications of Electrolysis

Watch the copper plating of a brass key! Click here or scan the QR code. https://www.youtube.com/watch?v=FnJ0V7B7nKo

6.1 Electroplating

- **Electroplating** is the process of **coating a layer of metal** on an object using electrolysis.
- In electroplating, the object to be plated is made the **cathode** of an electrolytic cell and the anode is the source of the plating metal (e.g. silver).
- The electrolyte is an aqueous solution of a salt of the plating metal (e.g. silver nitrate).
- Use of electroplating:
  - Electroplating is used to get a good decorative finish and to prevent rusting.
  - Silver-plating is used to beautify a metal object
  - Chrome-plating is used to protect a metal object from corrosion.



Electroplating of an object

At the **Ag** anode (reactive anode):

- NO_3^- and H_2O migrate to the anode
- Ag anode is preferentially oxidised to form Ag^+
- Oxidation half-equation: **$Ag(s) \rightarrow Ag^+(aq) + e^-$**

At the **Ag coated object** (inert cathode):

- Ag^+ and H_2O migrate to the cathode
- Ag^+ is preferentially reduced to form Ag. This Ag is then coated around the object hence electroplating it with silver.
- Reduction half-equation: **$Ag^+(aq) + e^- \rightarrow Ag(s)$**
  deposited on the object

Quick-check

The heat reflecting shields of space rockets are electroplated with gold. Which of the following are most likely to be used as electrodes and electrolyte to gold plate the heat shield?

	Negative electrode	Positive electrode	Electrolyte
A	carbon	heat shield	gold compound
B	heat shield	gold	gold compound
C	gold	heat shield	platinum
D	heat shield	carbon	platinum

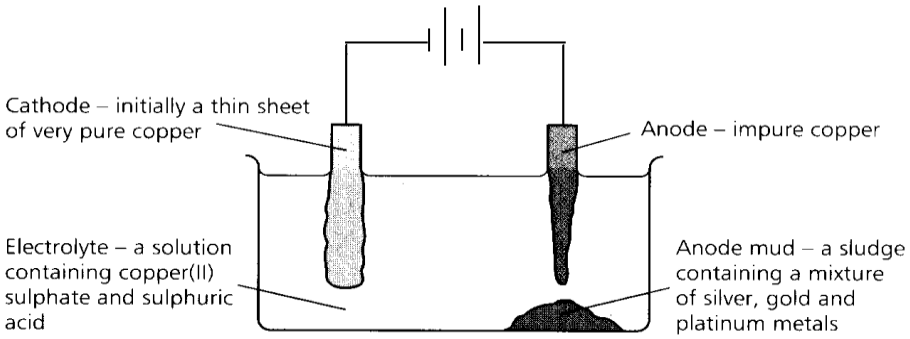
Heat shield is to be gold plated thus it is the cathode (negative electrode) where reduction of gold ions can occur to give gold metal.

Positive electrode is the anode hence it must be made out of gold so gold is oxidized to form gold ions to replenish the electrolyte

The electrolyte used is gold compound.

6.2 Purification of copper

Copper is about 99% pure when first obtained from its ore. The main impurities are silver, platinum, iron, gold and zinc. This level of impurities is sufficient to reduce the electrical conductivity of copper. Thus, copper must be purified (i.e. refined) by electrolysis before it can be used in, for example, electrical wires.



Refining copper using electrolysis

Electrolyte: $CuSO_4(aq)$

Anode: Impure copper

- Impure copper dissolves into the solution.
- Oxidation half-equation: $\mathbf{Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-}$
  impure copper

Cathode: Pure copper

- Thin sheet of very pure copper is placed at the cathode.
- Reduction half-equation: $\mathbf{Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)}$
  pure copper
- Cu^{2+} ions are discharged and deposited on the pure copper cathode.

Net result

- Transfer of Cu from impure anode to pure cathode.
- The reactive impure copper **anode** becomes **thinner** and the pure copper **cathode** becomes **thicker**.
- **The concentration of aqueous copper(II) sulfate electrolyte solution remains unchanged.**
- Impurities such as silver and platinum fall to the bottom of the cell and form the anode mud.

7 Galvanic Cell

7.1 Definition of a Galvanic Cell

Definition:
Galvanic cells (also termed as voltaic cells, electrochemical cell, Daniel Cells) are electrochemical cells in which **spontaneous** reactions occur to convert chemical energy to electrical energy.

Comparing and contrasting between the two main types of cells:

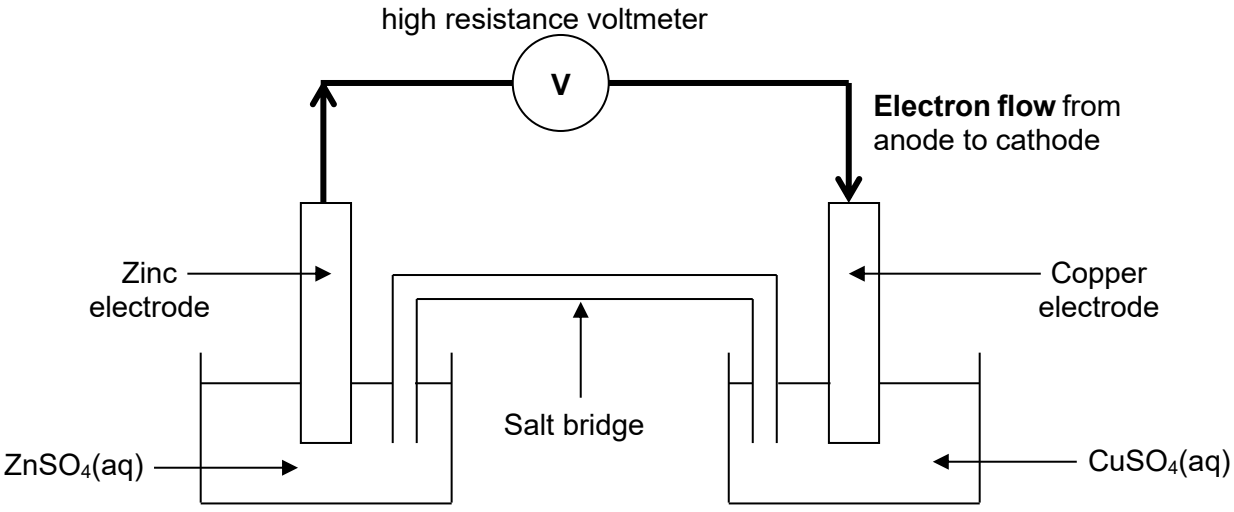
1. Galvanic Cell / Voltaic Cell (acting as batteries)

- Conversion of **chemical** energy to **electrical** energy
- Redox reaction is a **spontaneous** process

2. Electrolytic Cell (in electrolysis)

- Conversion of **electrical** energy to **chemical** energy
- Redox reaction is a **non-spontaneous** process.

7.2 Set-up of a Galvanic Cell



Anode (–ve terminal)
Oxidation
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Cathode (+ve terminal)
Reduction
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Observations:

1. Light bulb **glows** $\Rightarrow$ **current** is generated
2. Zn electrode gets **thinner**
3. Cu electrode gets **thicker**
4. Blue colour of $CuSO_4$ solution **decolourises**

At the zinc half-cell:

- **Oxidation occurs** hence Zn electrode is the anode

  Anode half-equation: **$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$**

- Zn^{2+} ions enter the solution.
- Electrons are **released** and flow from **Zn anode** to **Cu cathode** electrode through the wire.

In the external circuit:

As electrons travel through the wire, the current generated lights up the bulb.

At the copper half-cell:

- **Reduction occurs** hence Cu electrode is the cathode

  Cathode half-equation: **$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$**

- Cu^{2+} ions **accept** electrons to form Cu atoms plated on the surface of Cu electrode.
- Amount of Cu^{2+} in solution will decrease hence blue solution decolourises.

Electron flow: **Electrons** always flow from **anode** to **cathode**.

Cell reaction or overall reaction:

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Summary:

1. **No direct transfer of electrons** from Zn metal to Cu^{2+} ions as they are separated from one another. Electrons must travel through the wire.
2. **Electrons** always flow from **anode** to **cathode** through the external circuit
3. Conversion of **chemical** energy to **electrical** energy.
4. Electrode where **oxidation** occurs is known as **anode**.
5. Electrode where **reduction** occurs is known as **cathode**.
6. For electrochemical cell, the polarity of **anode is negative** and polarity of **cathode is positive**. *(The reverse of electrolytic cell)*

***Teacher's note*: Description and the 2 functions of Salt Bridge**

It is an inverted U-shaped tube containing salt solutions, e.g. Na_2SO_4 or KCl solution.

2 functions of salt bridge:

1. **Prevent direct mixing of the two solutions but allow ions to pass through them.**

2. **Maintain electrical neutrality** in each half-cell.

Without the salt-bridge, the zinc half-cell would slowly become more positively charged as more Zn^{2+} ions is produced, while the copper half-cell would become negatively charged as more Cu^{2+} is used up.

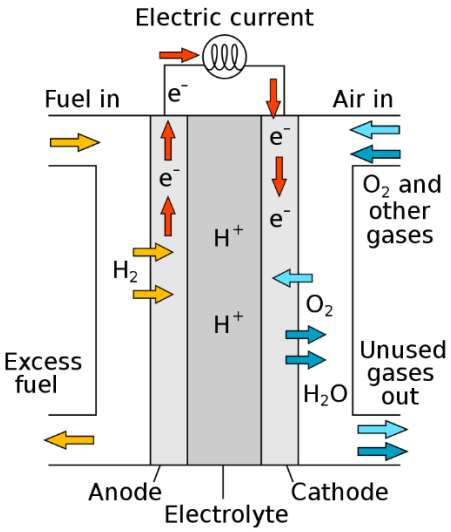
7.3 Reactivity Series and Electromotive Force (e.m.f.)

In a galvanic cell, electrons always flow **from the more reactive metal to the less reactive metal**. Therefore, the more reactive metal becomes the anode and is the negative electrode. The less reactive metal becomes the cathode and is the positive electrode.

The e.m.f. of a magnesium-copper cell is 2.7 V. If magnesium is replaced by zinc, a less reactive metal, the e.m.f. decreases to 1.1. V. This shows that **the further apart the two metals are in the reactivity series, the greater the e.m.f. produced**.

8 Fuel Cells

Combustion reactions are only one means of extracting useful energy from fuels. An alternative is the fuel cell, which is **a device in which a fuel reacts with the oxygen in the air to produce electrical energy directly**. The best-known example is the hydrogen-oxygen fuel cell.



A fuel cell is a chemical cell. At the cathode, oxygen is reduced to form hydroxide ions. At the anode, hydrogen is oxidised to form water.

Cathode: $O_2(g) + 2H_2O(l) + 4e \longrightarrow 4OH^-(aq)$

Anode: $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e$

The overall cell reaction is simply the conversion of hydrogen and oxygen to water, and the equation resembles that of the combustion of hydrogen. Note that a fuel cell differs from an ordinary battery or chemical cell in one important aspect, i.e. the reactants are not contained within the cell but instead are continuously supplied from an external source.

Advantages of the fuel cell:

- Potential alternative fuel and renewable energy resource, as the world's reserves of fossil fuels are rapidly declining
- Pollution-free, as the only by-product formed is water

Disadvantages of the fuel cell:

- Challenging to find a cheap source of hydrogen. Although there is plenty of hydrogen on the planet, it is mostly combined with oxygen in the form of water, and to obtain it, electrolysis of water needs to be done, which is expensive and energy-intensive in itself. Often, hydrogen is obtained from hydrocarbons instead
- Difficult to store hydrogen gas
- Hydrogen gas is highly flammable

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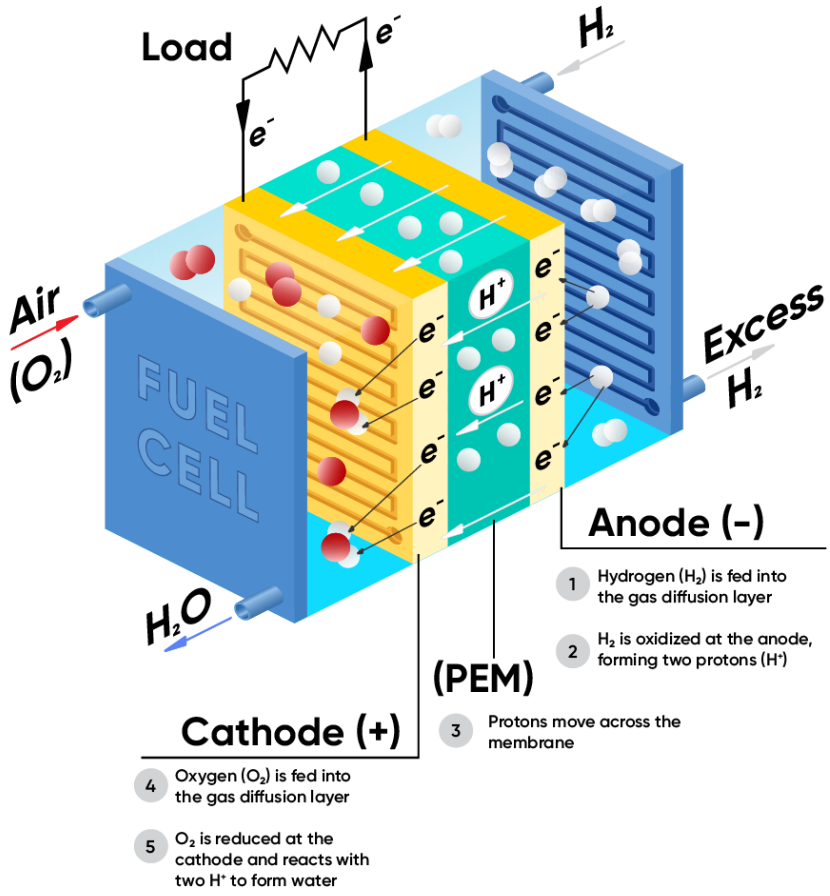
For your information

Electrolysis is a leading hydrogen production pathway, and the hydrogen produced via electrolysis can result in zero greenhouse gas emissions, depending on the source of the electricity used. The electrolytic cell used industrially is often called an electrolyser.

Hydrogen production via electrolysis is being pursued for renewable (wind, solar, hydro, geothermal) and nuclear energy options, because today's power grid is not ideal for providing the electricity required for electrolysis because of the greenhouse gases released and the amount of fuel required due to the low efficiency of the electricity generation process.



Appendix: Interesting Resources for Electrochemistry

1. Electrolysis and extraction of aluminium
   https://www.bbc.co.uk/bitesize/guides/zhk6pbk/revision/4

2. Electroplating
   https://www.britannica.com/technology/electroplating

3. Article: Bauxite in Malaysia: The environmental cost of mining
   https://www.bbc.com/news/world-asia-35340528

4. Article: MIT's accidental smelting discovery unlocks the potential of antimony
   https://www.mining-technology.com/features/featuremit-accidental-smelting-discovery-unlocks-the-potential-of-antimony-5681719/

5. Article: Electrolysis breakthrough could solve the hydrogen conundrum
   https://phys.org/news/2019-09-electrolysis-breakthrough-hydrogen-conundrum.html