Electrochemistry: Electrolysis & Simple Cell (Part 1)

Learning Objectives

- describe electrolysis as the conduction of electricity by an ionic compound (an electrolyte), when molte dissolved in water, leading to the decomposition of the electrolyte
- describe electrolysis as evidence for the existence of ions which are held in a lattice when solid but wh are free to move when molten or in solution
- describe, in terms of mobility of ions present and the electrode products, the electrolysis of molten sodiu chloride, using inert electrodes
- predict the likely products of the electrolysis of a molten binary compound
- apply the idea of selective discharge based on
  - (i) cations: linked to the reactivity series
  - (ii) anions: halides, hydroxides and sulfates (e.g. copper(II) sulfate & dilute sodium chloride)
  - (iii) concentration effects (as in the electrolysis of concentrated and dilute aqueous sodium chloride) (us inert electrodes)
  - (iv) effect of electrode (copper electrode vs carbon electrodes in electrolysis of copper(II) sulfate solutio
- predict the likely products of the electrolysis of an aqueous electrolyte, given relevant information
- construct ionic equations for the reactions occurring at the electrodes during the electrolysis, given relevant information
- describe the electrolysis of aqueous copper(II) sulfate with copper electrodes as a means of purifying coppe
- describe the electroplating of metals, e.g. copper plating, and state one use of electroplating
- describe the production of electrical energy from simple cells (i.e. two electrodes in an electrolyte) linked t the reactivity series and redox reactions (in terms of electron transfer)
- describe hydrogen, derived from water or hydrocarbons, as a potential fuel, reacting with oxygen to generat electricity directly in a fuel cell (details of the construction and operation of a fuel cell are not required).

---

1 CONDUCTORS OF ELECTRICITY

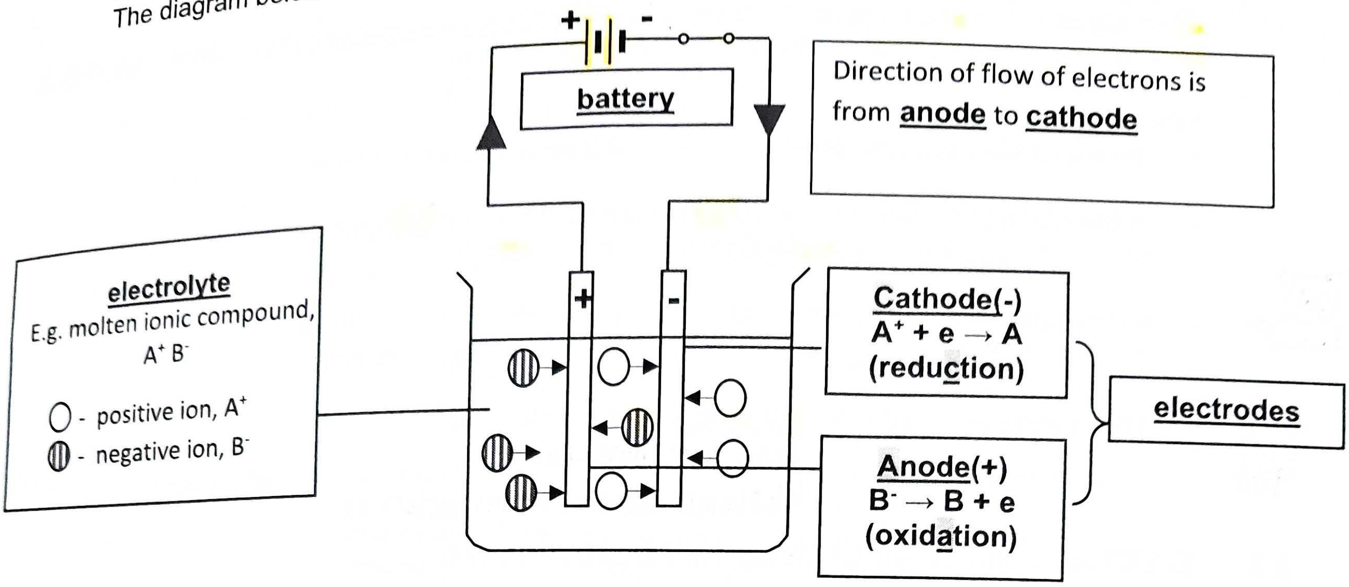
Recall that the following substances can conduct electricity:

- All metals, alloys and graphite conduct electricity.
- All ionic compounds conduct electricity when molten or in aqueous solutions.
- Some covalent substances which dissolve and ionise in water (e.g. acids: HCl, H_2SO_4, etc.)

Substances	Arrangement of particles	Particles that conduct electricity
Metals	positive ions in a 'sea' of free-moving valence electrons	free-moving valence electrons
Graphite	only three out of four valence electrons per carbon atom are used in covalent bonding, the remaining valence electron becomes free-moving	free-moving valence electrons
Ionic compounds	positively and negatively charged ions held together by strong electrostatic forces of attraction	free-moving ions (when compound is molten or in aqueous solution)

2 ELECTROLYTIC CELL

An electrolytic cell consists of the battery, the electrodes, and an electrolyte.
The diagram below shows an electrolytic cell:



2.1 BATTERY & ELECTRODES

The battery acts as an 'electron pump', drawing electrons away from the **electrode** which becomes **positively** charged. Electrons are supplied to the **electrode** which becomes **negatively** charged.

An electrode is a conductor made up of a rod or plate (i.e. graphite/platinum) in which electrons enter or leave the electrolyte.

There are two types of electrodes: **Cathode** and **Anode**

Cathode (-) **(negative electrode)**	**Anode (+)** **(positive electrode)**
negative electrode that is connected to the **negative** terminal of the battery.	positive electrode that is connected to the **positive** terminal of the battery
electrons ENTER electrolyte	electrons LEAVE electrolyte
Cations (positively charged ions) are attracted and migrate to the cathode where they gain electrons from the cathode	Anions (negatively charged ions) are attracted and migrate to the anode where they lose their electrons to the anode

2.2 ELECTROLYTES

An electrolyte is a compound which when molten or in aqueous solution, conducts electricity with its decomposition at the electrodes. Examples: acids, bases, salts

Examples of non-electrolytes, weak and strong electrolytes:

Non-electrolytes	**Weak electrolytes**	**Strong electrolytes**
Ethanol $C_2H_5OH(l)$ Pure water $H_2O(l)$ Sugar solution $C_{12}H_{22}O_{11}(aq)$	Weak acids and alkalis: Ethanoic acid $CH_3COOH(aq)$ Aqueous ammonia $NH_3(aq)$	Strong acids and alkalis, salt solutions: Dilute sulfuric acid $H_2SO_4(aq)$ Dilute hydrochloric acid $HCl(aq)$ Aqueous sodium hydroxide $NaOH(aq)$ Aqueous copper(II) sulfate $CuSO_4(aq)$

3 ELECTROLYSIS

3.1 WHAT HAPPENS DURING ELECTROLYSIS

Electrolysis is the chemical decomposition of a compound (electrolyte), by passing electri through the compound, when molten or in aqueous solution.

When the current in the external circuit is switched on,

- the electrodes become positively charged (**anode**) and negatively charged (**cathode**).
- in the electrolyte, positive ions (**cations**) migrate to the **cathode**, while the negative ions (**anions**) migrate to the **anode**
- the **cations** gain electrons and undergo **reduction** at the **cathode**, while the **anions** lose electrons and undergo **oxidation** at the **anode**.
- These **ions** are said to be **discharged** at the electrodes.

↳ no longer have charges

↳ becomes neutral atoms / molecules
metal / non-metal

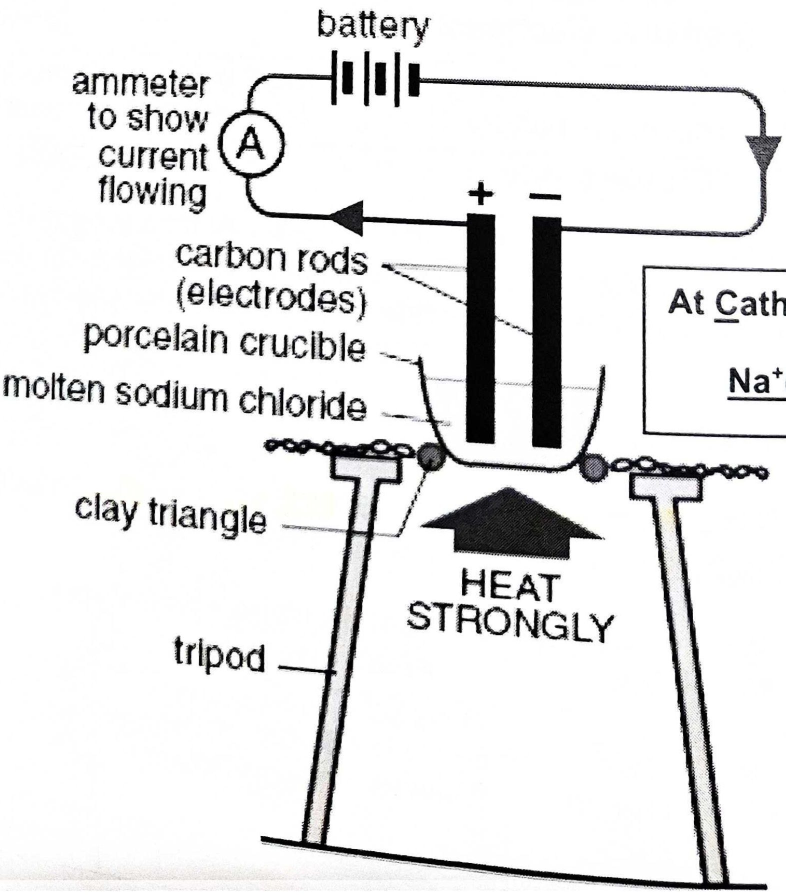
3.2 ELECTROLYSIS OF MOLTEN BINARY ELECTROLYTE

When a molten ionic compound (***binary** electrolyte) is electrolysed, the **metal** is produced at the **cathode**, while the **non-metal** is produced at the **anode**.

(Note: a ***binary** compound is a compound containing only two elements)

Example 1: Molten sodium chloride, Na*Cl*

Ions present in electrolyte: Na^+, Cl^-



At Anode (oxidation):

$2Cl^-(l) \rightarrow Cl_2(g) + 2e$

At Cathode (reduction):

$Na^+(l) + e \rightarrow Na(l)$

Reaction at CATHODE (Reduction)	Reaction at ANODE (Oxidation)
Sodium ions (cations) migrate / move towards the cathode. Sodium ions are <u>discharged</u> where they <u>gain electrons</u> to form <u>sodium metal</u>. Half equation: **$Na^+(l) + e \rightarrow Na(l)$** Observation: **<u>Silvery globules</u>** of **<u>molten sodium</u>** are found floating on the surface of electrolyte.	Chloride ions (anions) migrate / move towards the anode. Chloride ions are <u>discharged</u> where they <u>lose electrons</u> to form <u>chlorine gas</u>. Half equation: **$2Cl^-(l) \rightarrow Cl_2(g) + 2e$** Observation: **<u>Effervescence</u>** of **<u>pale greenish yellow, pungent gas</u>**

At cathode: $Na^+(l) + e \rightarrow Na(l)$ - (1) multiply by 2

At anode: $2Cl^-(l) \rightarrow Cl_2(g) + 2e$ - (2)

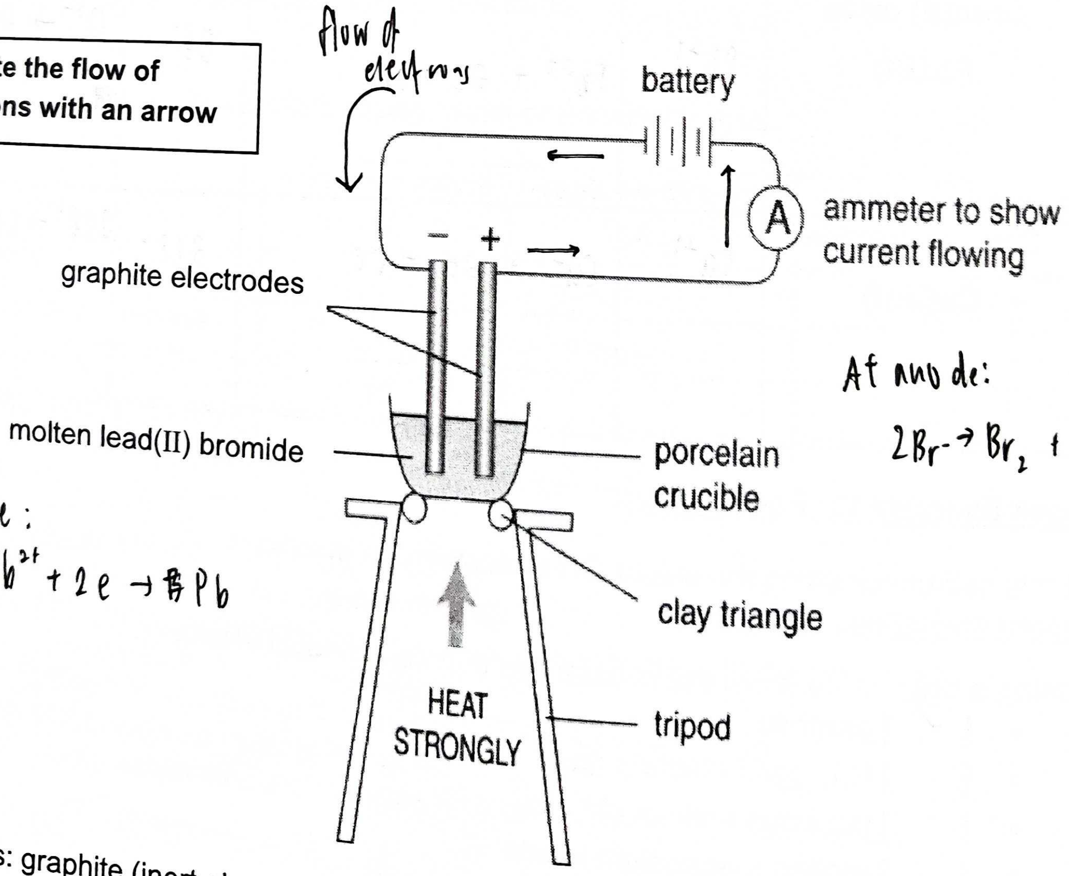
$2Na^+(l) + 2Cl^-(l) \rightarrow 2Na(l) + Cl_2(g)$

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

Part I Notes - Quick Check 1:

(1) In the set-up shown, molten lead(II) bromide is electrolysed.

Indicate the flow of electrons with an arrow



At cathode:

$Pb^{2+} + 2e \rightarrow Pb$

At anode:

$2Br^- \rightarrow Br_2 + 2e$

Electrodes: graphite (inert electrode)

Ions present in electrolyte: Pb^{2+}, Br^-

molten

Reaction at CATHODE (Reduction)	**Reaction at ANODE (Oxidation)**
(i) Half equation $Pb^{2+}(l) + 2e \rightarrow Pb(l)$	(i) Half equation $2Br(l) \rightarrow Br_2(g) + 2e$
(ii) What type of reactions occur at the cathode? Reduction	(ii) What type of reactions occur at the anode? Oxidation.
(iii) At the end of the experiment, shiny silvery globules was found in the remaining lead(II) bromide. Suggest its identity. lead	(iii) What you would see around the anode during the electrolysis? Effervescence of red-brown, pungent gas

Overall reaction: $PbBr_2(l) \rightarrow Pb(l) + Br_2(g)$

(2) Predict the products formed at the cathode and the anode. Write half equations at the electrodes.

Molten Electrolyte	Product at cathode	Half equation (with state symbols) at **cathode**	Product at anode	Half equation (with state symbols) at **anode**
Sodium iodide **NaI(l)**	Na	$Na^+ + e \rightarrow Na$	I_2	$I^- \rightarrow I_2 + e$
Lead(II) oxide **PbO(l)**	Pb	$Pb^{2+} + 2e \rightarrow Pb$	O_2	$O^{2-} \rightarrow O_2 + 2e$
Calcium chloride **CaCl₂(l)**	Ca	$Ca^{2+} + 2e \rightarrow Cl$	Cl_2	$2Cl^- \rightarrow Cl_2 + 2e$

Self Check Exercise for Part 1 notes:

Attempt this self-check using the link or QR to answer the questions:
https://forms.gle/oeZvigpz25HdxfJw9

1 Using a tick, '√' to show the substances that can conduct electricity.

- [✓] graphite
- [ ] hydrogen chloride gas
- [ ] aqueous sodium chloride
- [ ] molten magnesium iodide
- [ ] silver chloride
- [✓] ethanol
- [✓] sulfuric acid

2 Use '√ ' to indicate the correct particles which help conduct current in the named substance.

particles / named substance	free-moving ions	free-moving valence electrons	positive ions and free-moving valence electrons	free-moving cations and anions
metal		✓		
graphite		✓		
aqueous copper(II) chloride				
molten potassium chloride				

3 Use '√ ' to indicate the statement(s) that is/are true of the cathode in an electrolytic cell.

- [ ] it is a conductor electricity
- [ ] it is positively charged
- [ ] electrons leave the cathode and enter the electrolyte
- [✓] cations are attracted to the cathode
- [ ] reduction occurs here

4 Use '√ ' to indicate the statement(s) is/are true of an electrolyte.

- [ ] it is a conductor electricity
- [✓] it decompose
- [ ] it contain free-moving ions or electrons

5 Use '√ ' to classify the following into strong, weak or non-electrolyte.

	non-electrolyte	**weak electrolyte**	**strong electrolyte**
ethanoic acid		✓	
sulfuric acid			✓
aqueous sodium chloride	✓		
glucose solution	✓		
aqueous ammonia		✓	

6 Molten magnesium chloride is electrolysed.
Use '√ ' to show the option(s) is / are **not** correct.

- [ ] Magnesium chloride must be heated continuously to decompose it.
- [ ] Magnesium ions are reduced.
- [ ] The half equation at the anode is $2Cl(l) \rightarrow Cl_2(g) + 2e$
- [✓] Chloride ions migrate to the cathode.
- [ ] Magnesium ions and chloride ions are found in the electrolyte.

Electrochemistry: Electrolysis & Simple Cell (Part 2)

3.3 ELECTROLYSIS OF AQUEOUS ELECTROLYTE

Recall that in the electrolysis of molten electrolytes, there are only two kinds of ions.

However, in **aqueous solutions**, there are **four kinds of ions**; two come from the electrolyte and two come from the water.

Besides the free-moving **cations** and **anions** from the **electrolyte**, there are also a **low concentration** of **hydrogen ions, $H^+(aq)$** and **hydroxide ions, $OH^-(aq)$** from **water**.

Water undergoes **self-ionisation** and hence produces a **low concentration of H^+ and OH^- ions in solution:** $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

Both the **cations** from the electrolyte and the **hydrogen ions** from water would be attracted to the **cathode**.

Both the **anions** from the electrolyte and the **hydroxide ions** from water would be attracted to the **anode**.

State the ions produced by the following electrolytes:

Electrolyte	Positive ion(s) present	Negative ion(s) present
Dilute hydrochloric acid	$H^+(aq)$	$Cl^-(aq)$ $OH^-(aq)$
Sodium hydroxide solution	$Na^+(aq)$ $H^+(aq)$	$OH^-(aq)$
Copper(II) sulfate solution	$Cu^{2+}(aq)$ $H^+(aq)$	$SO_4^{2-}(aq)$ $OH^-(aq)$

One type of ion will be **preferentially discharged** at each electrode.

The type of ion selected to be discharged, and hence the product(s) formed at each electrode depend on the following factors:

FACTORS AFFECTING THE DISCHARGE OF IONS DURING ELECTROLYSIS

(i) Relative position of ions in the electrochemical series

(ii) Relative concentration of ions

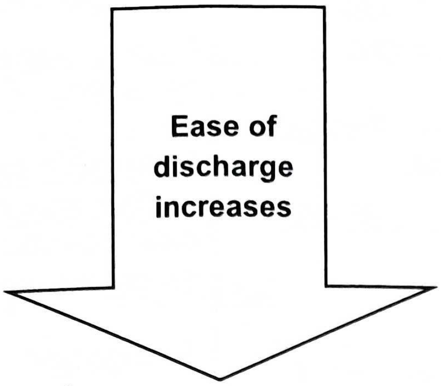
(iii) Nature of electrode

(i) Relative position of ions in the electrochemical series

Cations

K^+
*Ca^{2+}
Na^+
Mg^{2+}
Al^{3+}
Zn^{2+}
Fe^{2+}
Pb^{2+}
H^+
Cu^{2+}
Ag^+
Au^+

greater chance of being discharged



Anions

F^-
SO_4^{2-}
NO_3^-
Cl^-
Br^-
I^-
OH^-

usually not discharged

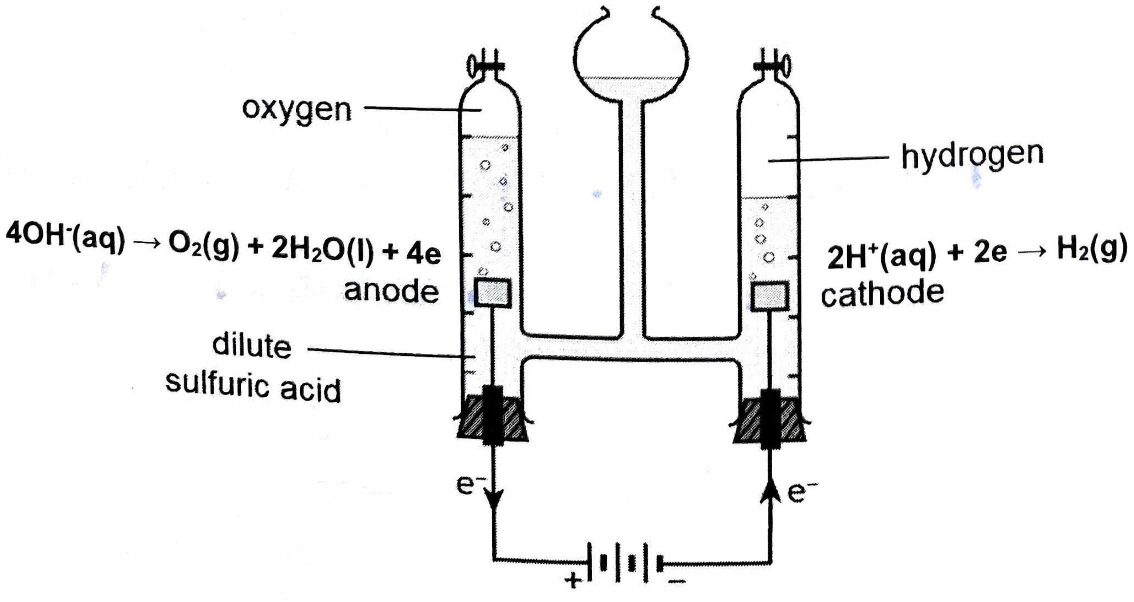
If all the other factors (concentration and nature of electrode) remain constant, the ions lower in the series will be discharged in preference to those above it.
i.e. An ion lower in the electrochemical series will show a greater tendency to be discharged.

- **Sulfate ions** and **nitrate ions** are usually **not** discharged.
- In **dilute aqueous halide solution**, **hydroxide ions** are **preferentially discharged** at the anode to chloride, bromide and iodide ions.
- **Hydroxide ions** from water are discharged to form **oxygen** and **water**.

Equation for **discharge of hydroxide ions,** OH^-: $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$

Example 1: Electrolysis of Dilute Sulfuric Acid (using Inert Electrodes)

Ions from electrolyte: $H^+(aq)$, $SO_4^{2-}(aq)$ and $OH^-(aq)$



Reaction at CATHODE **(Reduction)**	**Reaction at ANODE** **(Oxidation)**
Hydrogen ions are attracted to the cathode. Hydrogen ions are discharged where they gain electrons to produce **hydrogen gas**. Half equation: **$2H^+(aq) + 2e \rightarrow H_2(g)$** Observation: **Effervescence of colourless, odourless gas.**	Sulfate ions and hydroxide ions are attracted to the anode. **Hydroxide ions** are **preferentially discharged** as they lose electrons more readily (lower in the electrochemical series than sulfate ions) to produce **water and oxygen gas**. Half equation: **$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e$** Observation: **Effervescence of colourless, odourless gas.**

At the cathode: $4H^+(aq) + 4e^- \rightarrow 2H_2(g)$ *half equation 1*
At the anode: $4OH^-(aq) \rightarrow 2H_2O\ (l) + O_2(g) + 4e^-$ *half equation 2*

$4H^+(aq) + 4OH^-(aq) \rightarrow 2H_2(g) + 2H_2O(l) + O_2(g)$ *half equations 1+2*
$4H_2O(l) \rightarrow 2H_2(g) + 2H_2O(l) + O_2(g)$

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

Thus, the electrolytic products are **2 moles** of **hydrogen** at the cathode and **1 mole** of **oxygen** at the anode for every 2 moles of water electrolysed.

Based on Avogadro's Law:
Equal volume of gases at same temperature and pressure contain the same number of gas particles.

Thus, we can also say that the electrolytic products are **2 volumes** of **hydrogen** at the cathode and **1 volume** of **oxygen** at the anode.

Changes in the solution:
During the electrolysis, **only water decomposes**. The **number of moles of sulfuric acid remains unchanged** at the end of the electrolysis.

Due to the **removal of water molecules**, the **concentration** of **sulfuric acid solution increases.**

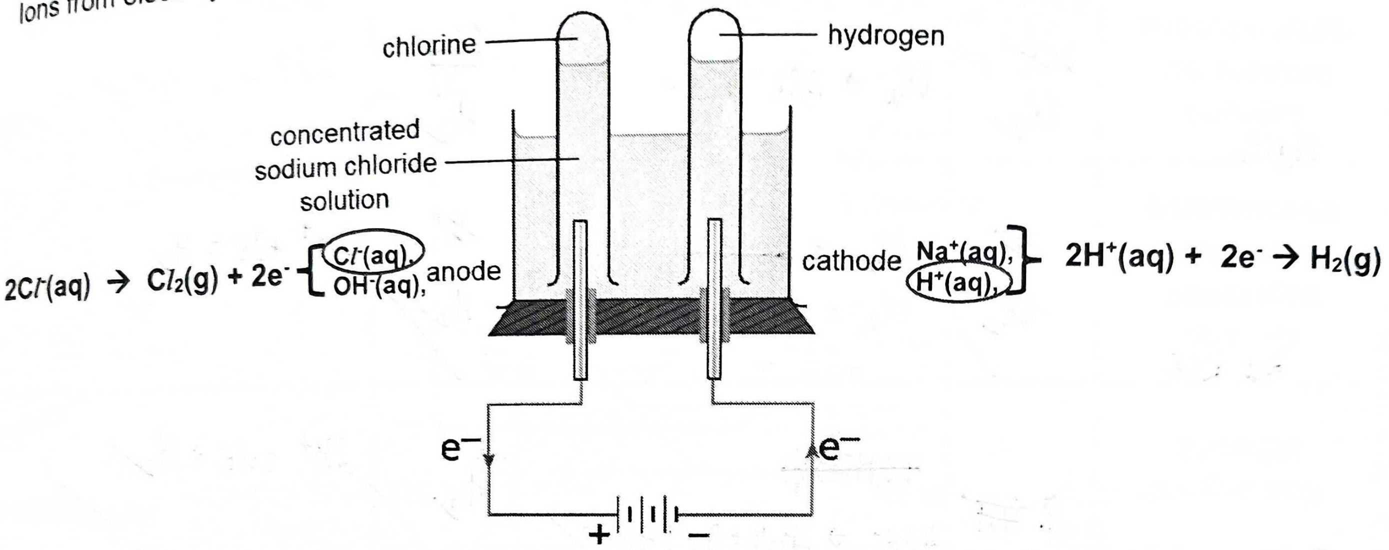
(ii) Relative concentration of ions

High concentration of an **ion** tends to promote its **discharge.**

Influence of concentration is **usually secondary** to that of the position of ions in the electrochemical series, and it is **effective only** when the **two 'rival' ions are closely positioned** in the series. The **further apart** they are, the **less important** is the effect of concentration.

Example 2: Electrolysis of Concentrated Sodium Chloride Solution (using Inert Electrodes)

Ions from electrolyte: **$Na^+(aq)$, $Cl^-(aq)$, $H^+(aq)$, $OH^-(aq)$**



Reaction at CATHODE (Reduction)	**Reaction at ANODE (Oxidation)**
Sodium ions and hydrogen ions are attracted to the cathode.	Chloride ions and hydroxide ions are attracted to the anode.
Hydrogen ions are **preferentially discharged** as they gain electrons more readily (lower in the electrochemical series than sodium ions) to produce hydrogen gas.	**Chloride ions** are **preferentially discharged** due to their **much higher concentration**, and they lose electrons to produce **chlorine gas**.
Half equation: **$2H^+(aq) + 2e^- \rightarrow H_2(g)$**	Half equation: **$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$**
Observation: **Effervescence of colourless, odourless gas.**	Observation: **Effervescence of pale greenish yellow pungent gas**

At cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

At anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

Overall equation:

The electrolytic products : **1 volume** of **hydrogen** at the **cathode** and **1 volume** of **chlorine** at the **anode.**

Changes in the solution:

The **concentration** of **hydrogen ions** and **chloride ions decreases** as they are being discharged at the electrodes, while the **sodium ions** and **hydroxide ions** remain in the solution.

Thus, the sodium chloride solution gradually becomes **sodium hydroxide solution** (**alkaline**), where pH of the solution **increases**.

Part 2 Notes - Quick Check 1:

The following electrolytes are electrolysed using **inert electrodes.** Complete the table below.

Electrolytes	oxygen gas Product at anode	Half equation at anode	molten aluminium Product at cathode	Half equation at cathode
molten aluminium oxide Al_2O_3	(aq) ~~$3O^{2-}$~~ O_3	$O_3 \rightarrow 3O^{2-} + 6e$ $2O^{2-} \rightarrow O_2 + 4e$	~~Al^{3+}~~ Al	$Al^{3+} + 3e \rightarrow Al$ ✓
dilute aqueous magnesium chloride $MgCl_2$	Water / oxygen gas ~~$2Cl^-$~~ Cl_2 ✓	$4OH \rightarrow O_2 + 2H_2O + 4e$ $Cl_2 \rightarrow 2Cl^- + 2e$	~~Mg~~ H_2 ✓	$2H^+ + 2e \rightarrow H_2$ ✓
concentrated aqueous potassium chloride ~~K~~ KCl	~~Cl~~ Cl_2 ✓	~~$Cl_2 \rightarrow 2Cl^-$~~ $Cl_2 \rightarrow 2Cl^- + 2e$ $2Cl^- \rightarrow Cl_2 + 2e$	~~H~~ H_2 ✓	$2H^+ + 2e \rightarrow H_2$ ✓ ✓
aqueous zinc sulfate $ZnSO_4$	q ~~OH~~ H_2O ✓	~~$4OH^-(aq)$~~ $4OH^- \rightarrow 2H_2O + O_2 + 4e$ ✓	H_2 ✓	$2H^+ + 2e \rightarrow H_2$ ✓ ✓
molten potassium hydride, KH	H_2 ~~OH~~ H_2O	$2H^-(l) \rightarrow H_2(g) + 2e$ $4OH^- \rightarrow 2H_2O + O_2 + 4e$	potassium H_2 ✓	$K^+ + e \rightarrow K$
concentrated aqueous caesium iodide CsI	~~I~~ I_2 (aq)	$2I_2 \rightarrow I^- + 2e$ ✓	~~Cs~~ Cs ✗ H_2	$Cs^+ + e \rightarrow Cs$ ✗ $2H^+ + 2e \rightarrow H_2$ ✓

CsOH
pH of solution increases.

Cs > K
Cs will ~~done~~ never be discharged.

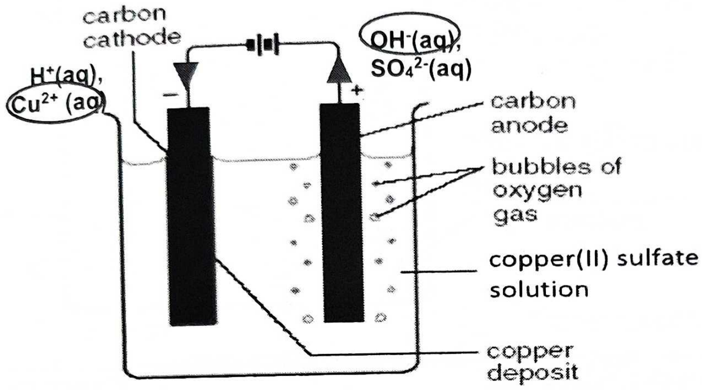
(iii) Nature of electrode

Inert electrodes (made of platinum or graphite) do **not** react in the electrolysis and thus do **not** take part in the electrolytic reaction.

Reactive electrodes (made of metals) take part in the electrolysis and **affect** the **discharge of ions.**

Example 3(a): Electrolysis of Copper(II) Sulfate Solution (using Inert Electrodes)

Ions from electrolyte: $Cu^{2+}(aq)$, $SO_4^{2-}(aq)$, $H^+(aq)$ and $OH^-(aq)$



Reaction at CATHODE (Reduction)	**Reaction at ANODE (Oxidation)**
Copper(II) ions and hydrogen ions are attracted to the cathode. **Copper(II) ions** are **preferentially discharged** as they **gain electrons more readily** (lower in the electrochemical series than hydrogen ions) to produce **copper metal**. Half equation: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ Observation: **Red-brown solid deposited at the cathode.**	Sulfate ions and hydroxide ions are attracted to the anode. **Hydroxide ions** are **preferentially discharged** as they **lose electrons more readily** (lower in the electrochemical series than sulfate ions) to produce **oxygen gas** and **water**. Half equation: $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$ Observation: **Effervescence of colourless, odourless gas.**

At the cathode: $2Cu^{2+}(aq) + 4e \rightarrow 2Cu(s)$ equation 1

At the anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ equation 2

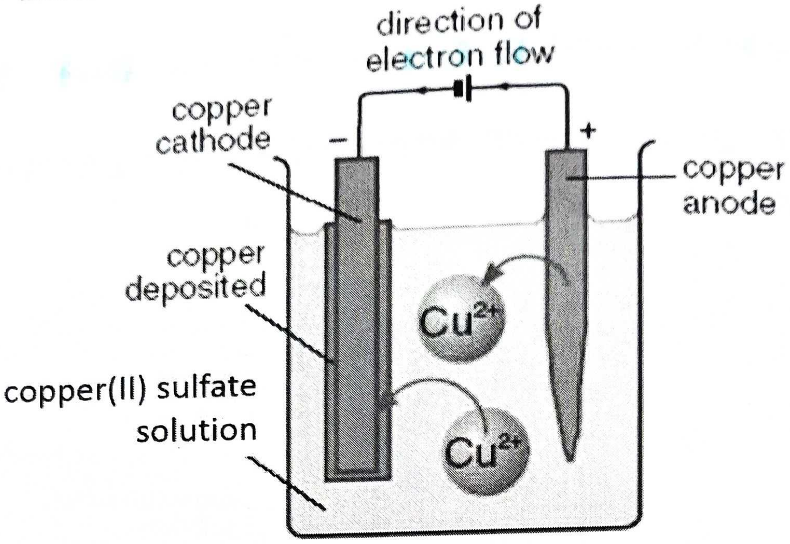
Overall equation: $2Cu^{2+} + 4OH^-(aq) \rightarrow \cancel{4}\cancel{2Cu} 2Cu(s) + O_2(g) + 2H_2O(l)$

Thus, the electrolytic products are **copper deposits** at the **cathode** and **oxygen gas** liberated at the **anode.**

Changes in the solution:
The concentration of copper(II) ions and hydroxide ions **decreases** as they are being discharged at the electrodes while the **hydrogen ions** and **sulfate ions** remain in the solution.
Thus, the **blue colour** of the solution slowly **fades / disappears** due to loss of copper(II) ions, and the copper(II) sulfate solution gradually becomes **sulfuric acid**, where **pH** of the solution **decreases**.

Example 3(b): Electrolysis of Copper(II) Sulfate Solution (using Copper Electrodes)

Ions from electrolyte: $Cu^{2+}(aq)$, $SO_4^{2-}(aq)$, $H^+(aq)$ and $OH^-(aq)$



Reaction at CATHODE **(Reduction)**	**Reaction at ANODE** **(Oxidation)**
Copper(II) ions and hydrogen ions are attracted to the cathode.	Sulfate ions and hydroxide ions are attracted to the anode.
Copper(II) ions. are **preferentially discharged** as they gain electrons more readily (lower in the electrochemical series than hydrogen ions) to produce **copper metal**	Both sulfate ions and hydroxide ions are **not discharged**, instead the **copper anode**, being a **reactive** anode, itself **loses electrons**. Thus, the **copper** in the anode undergoes **oxidation** to produce **copper(II) ions.**
Half equation: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	Half equation: ~~Cu(s) +~~ $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Observation: A **red-brown/pink solid** is deposited on the **copper cathode** which becomes **bigger** in size as the copper metal formed is deposited on it.	Observation: The **copper anode** becomes **smaller in size** as it forms **copper(II) ions** which go into the solution. (Overall, the **decrease** in mass of anode **equals** the **increase** in mass of cathode.)

Changes in the solution:

Although copper(II) ions are discharged **at the cathode**, for **each copper(II) ion discharged**, **one is formed at the anode which goes into the solution**.

Thus, the **blue colour** of the solution **remains the same**, and the **concentration of the copper(II) sulfate solution does not change.**

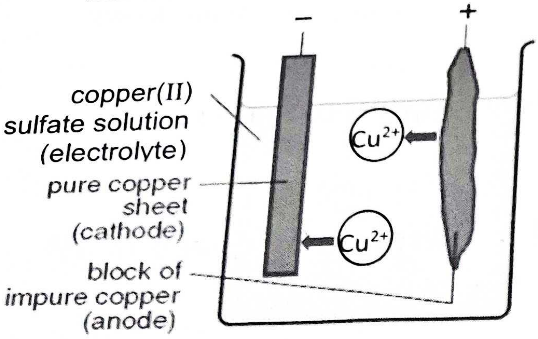
Electrolysis of aqueous copper(II) sulfate
[Comparing Examples 3(a) and 3(b)]

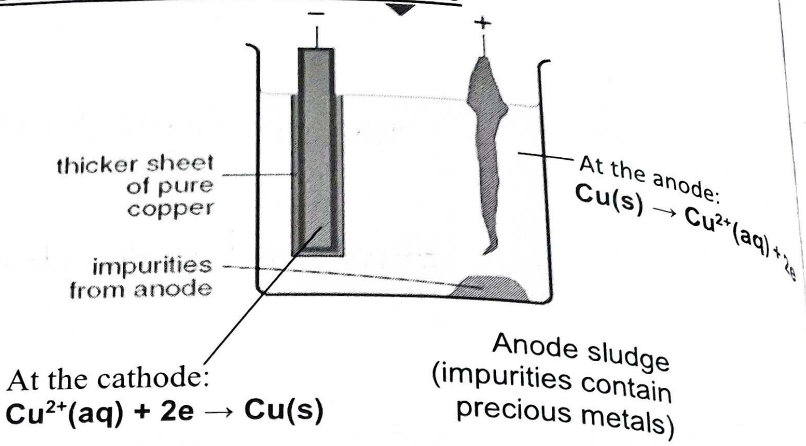
Set-up	**Example 3(a)** **Using graphite electrodes**	**Example 3(b)** **Using copper electrodes**
Equation and observation at anode	Equation: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e$ Observations: Effervescence of colourless, odourless gas.	Equation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e$ Observations: Copper anode decreases in size
Equation and observation at cathode	Equation: **$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$** Observations: **Reddish brown solid deposited on cathode**	Equation: **$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$** Observations: **Reddish brown solid deposited on cathode. Copper cathode** increases **in size.**
Observation of changes to electrolyte (if any)	Blue copper(II) sulfate solution fades	No change to blue copper(II) sulfate solution.

4 INDUSTRIAL APPLICATIONS OF ELECTROLYSIS

Two of the important industrial applications of electrolysis are the following:

(i) Refining/Purification of Copper - Electrolysis of Copper(II) Sulfate Solution using Copper Electrodes





The reactions at the cathode and anode are **similar to Example 3(b)**, except that in the refining/purification process, **pure copper** is made the **cathode**, while **impure copper** is made the **anode**.

Changes in the solution:
Similar to Example 3(b), the blue colour and the concentration of copper(II) sulfate solution does not change.

Note:

- *The **decrease** in mass of anode **equals** the **increase** in mass of cathode **plus the mass of the impurities, assuming impurities consist of only metals less reactive than copper**.*
- As the anode decreases in size, the **impurities fall off** to collect **below the anode** and remain in solution. These **impurities** contain **valuable silver and gold**. (form anode sludge)

The copper produced in this process is 99.9% pure. There are many uses of copper, including electrical wires and cables, water pipes and in alloys such as brass and bronze.

(ii) Electroplating

Electroplating is the **coating** of the **surface** of an object with a thin layer of a **metal** by **electrolysis**.

In **electroplating** an object:
- the object to be plated is made the **cathode**,
- a piece of the metal for electroplating the object is made the **anode**,
- an aqueous solution of a compound of the plating metal is used as the **electrolyte**.

) + 2e

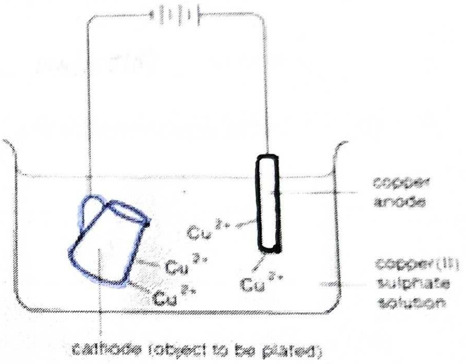
Note: the cathode can be a non-electrical conductor covered with **graphite paste** or a **thin layer of powdered graphite**.

The thickness of the plating may be controlled by adjusting the magnitude of the current and the time duration of the process.

Example 1: Copper Plating

In **copper plating** an object:
- the cathode is **object to be plated (e.g. cup)**
- the anode (reactive) is **the plating metal (e.g. copper)**
- the electrolyte can be **copper(II) sulfate solution / copper(II) nitrate solution / copper(II) chloride solution**

all nitrates are soluble.



Ions from electrolyte: **$Cu^{2+}(aq)$, $SO_4^{2-}(aq)$, $H^+(aq)$ and $OH^-(aq)$**

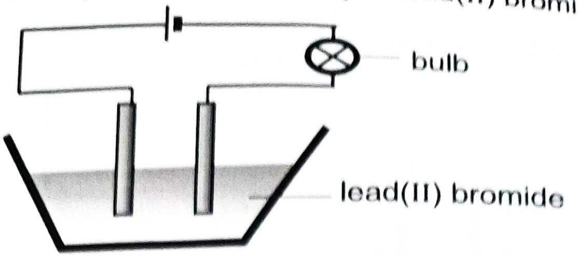
Reaction at CATHODE (Object to be Plated) (Reduction)	Reaction at ANODE (Pure Copper) (Oxidation)
Copper(II) ions and hydrogen ions are attracted to the cathode. **Copper(II) ions** are **preferentially discharged** as they gain electrons more readily (lower in the electrochemical series than hydrogen ions) to produce **copper metal.** Half equation: **$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$** Observation: A **red-brown/pink solid** is deposited on the **cathode which is the object** as the copper metal formed is deposited on it.	Sulfate ions and hydroxide ions are attracted to the anode. Both sulfate ions and hydroxide ions are **not discharged**, instead the copper anode, being a **reactive** anode, itself loses electrons. Thus, the copper atoms in the anode undergo **oxidation** to produce **copper(II) ions.** Half equation: **$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$** Observation: The **copper anode** becomes **smaller** in size as it forms **copper(II) ions** which go into the solution.

Changes in the solution:

Although copper(II) ions are discharged at the cathode, for each copper(II) ion discharged, one is formed at the anode which goes into the solution.
Thus, the **blue colour** of the solution, and the **concentration of copper(II) sulfate solution remain unchanged**.

MCQ Self- Check Exercise for Part 1 & Part 2 notes: Electrolysis

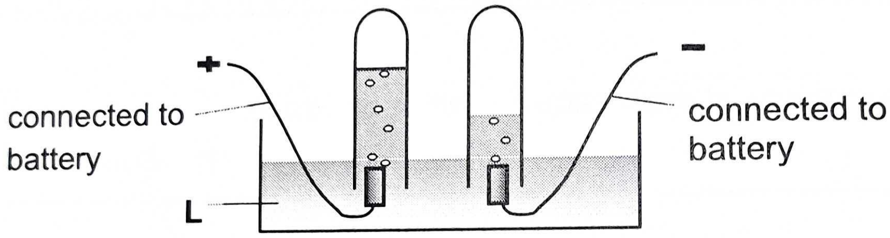
1 The diagram shows an experiment to electrolyse lead(II) bromide using carbon electrodes.



Why does the bulb only glow when the lead(II) bromide is melted?

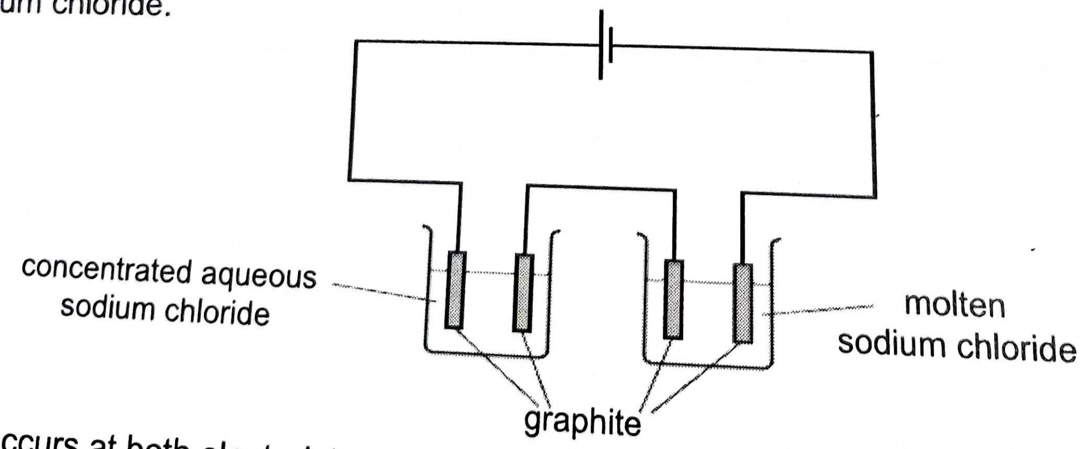
A Solid lead(II) bromide does not contain ions.
B Ions can only move through molten lead(II) bromide. (B)
C Ions are only formed when the lead(II) bromide melts.
D Electrons flow through molten lead(II) bromide.

2 The diagram shows the results of electrolysing **L.** Which **cannot** be the identity of the electrolyte **L**?



A aqueous copper(II) sulfate
B dilute sulfuric acid
C aqueous hydrochloric acid
D dilute aqueous sodium chloride (A)

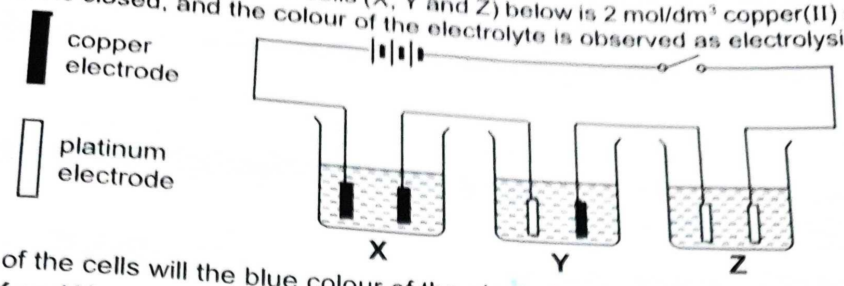
3 The diagram shows the electrolysis of concentrated aqueous sodium chloride and of molten sodium chloride.



What occurs at both electrolytes? (D)

A sodium ions are reduced
B hydrogen ions are reduced
C chlorine is oxidised
D chloride ions are oxidised

4 The electrolyte used in the three cells (X, Y and Z) below is 2 mol/dm^3 copper(II) sulfate solution. The switch is closed, and the colour of the electrolyte is observed as electrolysis progresses.



In which of the cells will the blue colour of the electrolyte fade?

A X and Y
B X and Z
C Y and Z
D X, Y and Z ( )

5 Which is the correct combination of anode, cathode, and electrolyte to electroplate silver coating on a commemorative coin?

	anode	cathode	electrolyte
A	silver	coin ✓	aqueous silver nitrate ✓
B	silver	coin ✓	silver chloride ✗
C	coin	silver	aqueous silver nitrate
D	coin	silver	silver chloride

(A) ~~(B)~~ ~~(A)~~

<table>
<tr><th>No.</th><th>Answers</th><th>Explanation</th></tr>
<tr><td>1</td><td>B</td><td>Lead(II) bromide, $PbBr_2$ has a giant ionic structure.
In the solid state, the oppositely charged ions, Pb^{2+} and Br^- are held at fixed position by strong electrostatic forces of attraction.
When melted, the strong electrostatic forces of attraction are overcome, oppositely charged ions in the molten state are free to move to conduct electricity</td></tr>
<tr><td>2</td><td>A</td><td>• products at cathode and anode are both gases
• ratio of gas formed at anode: cathode is 1:2
<table>
<tr><th></th><th colspan="2">ions present</th><th colspan="2" rowspan="2">half-equations</th></tr>
<tr><th></th><th>At anode(+)</th><th>At cathode(-)</th></tr>
<tr><td rowspan="2">A</td><td rowspan="2">SO_4^{2-}
$\mathbf{OH^-}$</td><td rowspan="2">$\mathbf{Cu^{2+}}$
H^+</td><td>At anode:</td><td>$\mathbf{4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e}$</td></tr>
<tr><td>At cathode:</td><td>$\mathbf{Cu^{2+}(aq) + 2e \rightarrow Cu(s)}$</td></tr>
<tr><td>B</td><td>Cl^-
$\mathbf{OH^-}$</td><td>$\mathbf{H^+}$</td><td colspan="2" rowspan="3">At anode: $\mathbf{4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e}$
At cathode: $\mathbf{2H^+(aq) + 2e \rightarrow H_2(g)}$</td></tr>
<tr><td>C</td><td>SO_4^{2-}
$\mathbf{OH^-}$</td><td>$\mathbf{H^+}$</td></tr>
<tr><td>D</td><td>Cl^-
$\mathbf{OH^-}$</td><td>Na^+
$\mathbf{H^+}$</td></tr>
</table></td></tr>
<tr><td>3</td><td>D</td><td>At both anodes: Chloride ions loses electrons (oxidised) and are discharged to form chlorine gas. $\mathbf{2Cl^- \rightarrow Cl_2 + 2e}$</td></tr>
<tr><td>4</td><td>C (where the anode at Y and Z are made of inert Pt metal)</td><td>• If anode(+) is made from reactive copper, for each Cu^{2+} discharged at the cathode, one is formed at the anode that goes to the solution. This means blue colour of electrolyte does not fade.
• If anode(+)is made from inert platinum, OH^- discharged at anode to give oxygen gas and water and Cu^{2+} discharged at the cathode. This means that the concentration of Cu^{2+} ions in the electrolyte will decrease ($Cu^{2+} + 2e \rightarrow Cu$), causing the blue solution to fade.</td></tr>
<tr><td>5</td><td>B</td><td>For electroplating:
Anode(reactive metal – plating silver metal) : $\mathbf{Ag(s) \rightarrow Ag^+(aq) + e}$
Cathode (object to be plated e.g. coin): $\mathbf{Ag^+(aq) + e \rightarrow Ag(s)}$
Electrolyte: salt solution containing Ag+ ions: $\mathbf{AgNO_3(aq)}$</td></tr>
</table>

Electrochemistry: Electrolysis & Simple Cell (Part 3)

5 SIMPLE CELLS

A **simple cell / electric cell** is a device which converts chemical energy to electrical energy. It consists of two electrodes placed in an electrolyte.

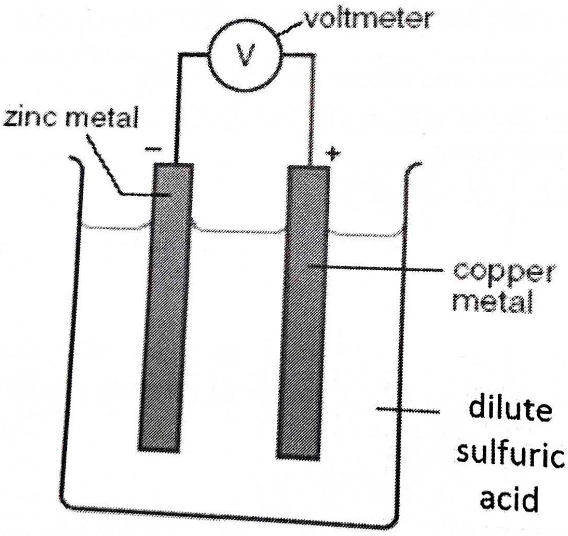
The two electrodes are usually **different metals** and the electrolyte can be an **aqueous salt solution** or a **dilute acid**, etc.

The **negative** electrode is the **more reactive** metal (higher up in the reactivity series), which **gives up electrons more readily**.

The **positive** electrode is the **less reactive** metal (lower down in the reactivity series), where ions in the solution **receive electrons from the more reactive metal**.

The **flow of electrons** is always from the **more reactive metal (negative electrode)** to the **less reactive metal (positive electrode)**.

Example 1:
Electrolyte: dilute sulfuric acid
Electrodes: zinc and copper



At the zinc electrode (negative electrode):

The **zinc atoms** in the electrode **loses electrons** and undergoes **oxidation** to form **zinc ions**.

Half equation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$

The **zinc ions** go into the solution.

The **electrons** then flow along the connecting wire to the copper electrode.

At the copper electrode (the positive electrode):

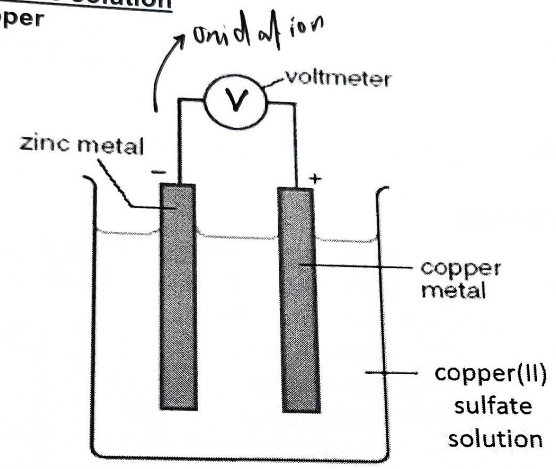
The hydrogen ions **gain electrons** and undergo **reduction** to form **hydrogen gas.**

Half equation: $2H^+(aq) + 2e \rightarrow H_2(g)$

At negative electrode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$
At positive electrode: $2H^+(aq) + 2e \rightarrow H_2(g)$

The overall ionic equation of the reaction: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Example 2:
Electrolyte: **copper(II) sulfate solution**
Electrodes: **zinc and copper**



At the zinc electrode (negative electrode):

The zinc atoms in the electrode **loses electrons** and undergoes **oxidation** to form zinc ions

Half equation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$

The zinc ions go into the solution.

The **electrons** then flow along the connecting wire to the copper electrode.

At the copper electrode (the positive electrode):

The copper(II) ions **gain electrons** and undergo **reduction** to form copper metal

Half reaction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

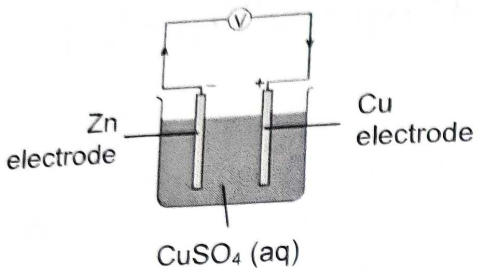
The overall ionic equation of the reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Other Electric Cells:

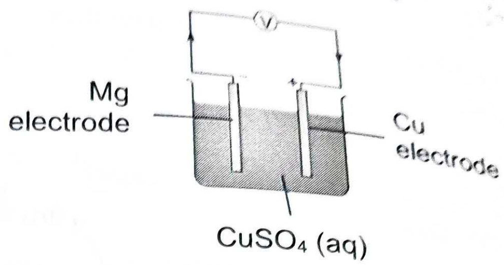
Other electric cells or simple cells can be made the same way.

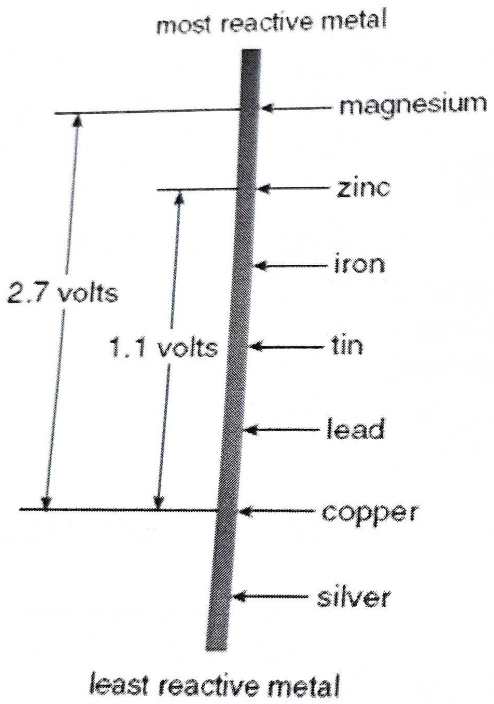
The **magnitude** of the **potential difference (voltage)** depends on the **relative positions** of the t metals in the reactivity series, that is, the **difference in the reactivity** of the two metals.

Voltmeter reading = 1.1 V



Voltmeter reading = 2.7 V





Metal electrodes	Voltage (V)
magnesium/copper	2.7
zinc/copper	1.1
iron/copper	0.8
lead/copper	0.5
copper/copper	0.0

- *What can be deduced about the relative positions of the two metals in the reactivity series and the voltage of the cell?*

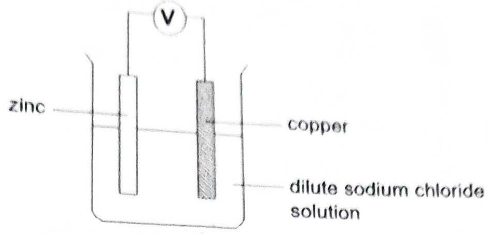
The **further apart the two metals** are in the reactivity series, the greater the voltage produced.

- *What happens when the electrodes are made of the* ***same metal****?*

No current will flow if both electrodes are made of the same metal. The voltmeter will read 0 V.

MCQ Self-Check Exercise for Part 3 notes - Simple Cells:

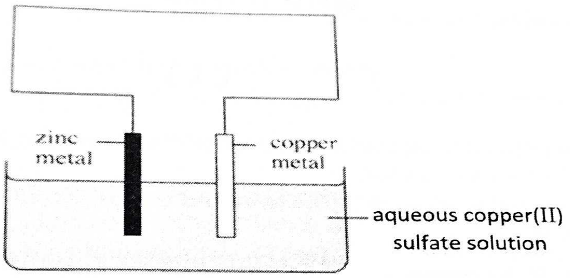
1. Which is most likely to occur in the experiment set up below?



A The copper electrode increase in mass.
B The solution gradually turns blue.
C The solution gradually becomes alkaline.
D Chlorine is evolved at the copper electrode.

(C)

2. Strips of copper and zinc were dipped into aqueous copper(II) sulfate as shown below.

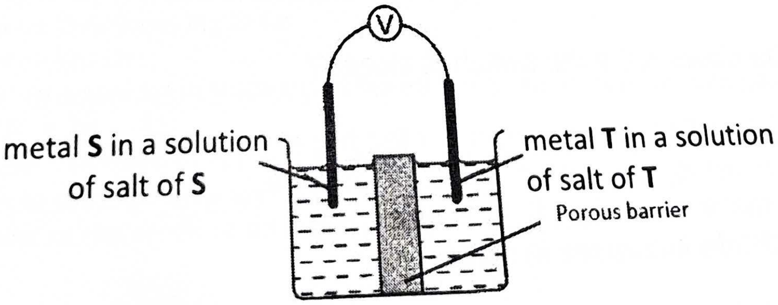


Which of the following will **not** occur?

A Zinc loses electrons.
B Zinc strip decreases in size.
C Pink solid is deposited on the copper strip.
D Electrons flow from copper to zinc in the external circuit.

(D)

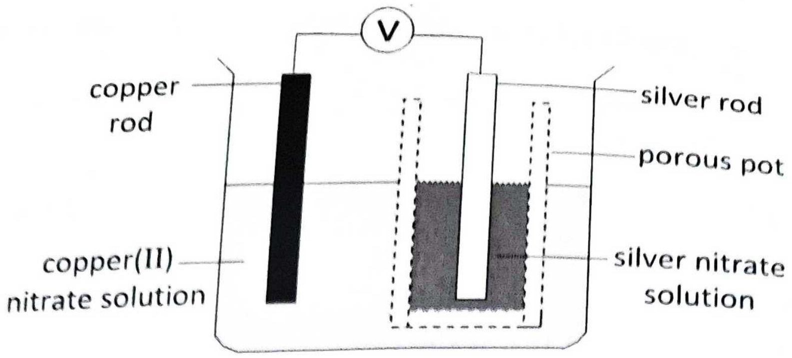
3. Which pair of metals, **S** and **T**, will produce the **lowest** voltage when used as electrodes in a simple cell?



(B)

	S	**T**
A	iron	silver
B	lead	silver
C	magnesium	silver
D	zinc	silver

4. The diagram below shows a simple cell.



$Cu \rightarrow Cu^{2+} + 2e^-$ $Ag^+ + e^- \rightarrow Ag$

What would take place in this simple cell?

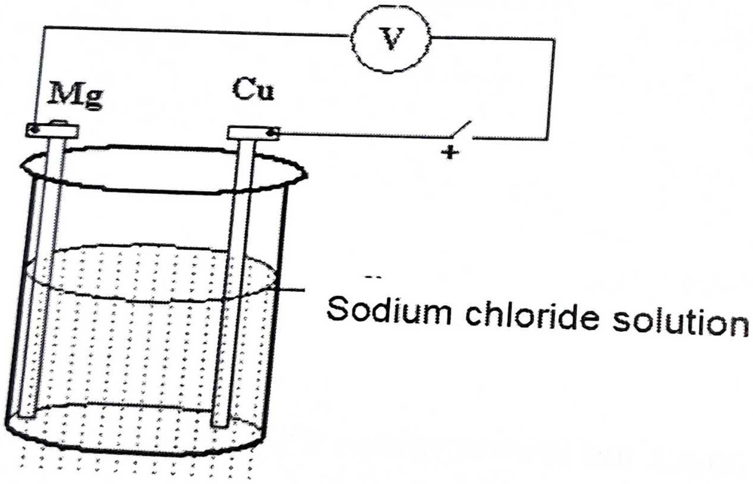
I ~~Electrodes~~ Electrons move from silver to copper.
II The colour intensity of copper(II) nitrate increases. ✓
III Copper acts as the reducing agent. ✓
IV The silver electrode becomes thinner

copper is oxidised (lost electrons)

~~A~~ I and III only
~~B~~ I and IV only
[✓ C] II and III only
~~D~~ III and IV only

(C)

5 A simple cell was set up as shown.



What would be observed if the switch is closed?

A Greenish-yellow gas observed at the Mg electrode.
B Effervescence observed at the Cu electrode.
C Silvery metal coated on the Cu electrode.
D Mg electrode increases in size.

(B)

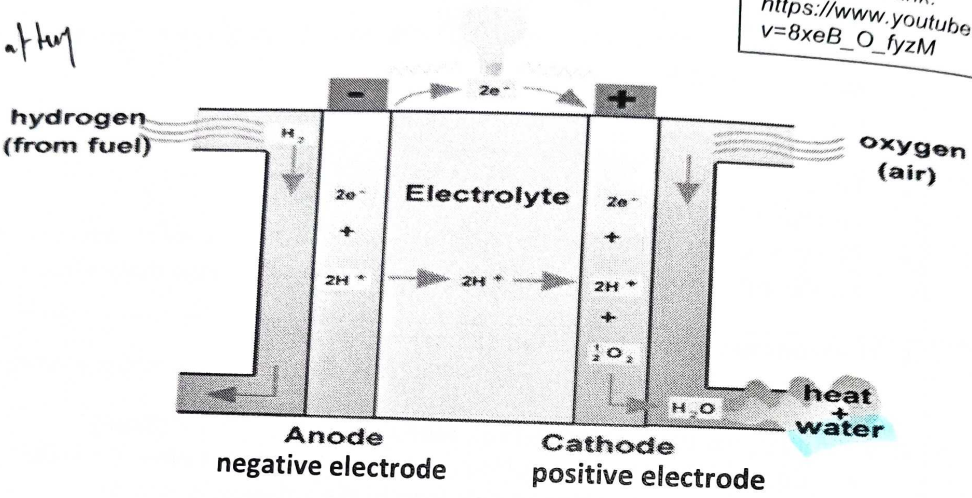
No.	Answers	Explanation
1	C	• At the zinc electrode (negative electrode), zinc loses electrons and undergoes oxidation : Zn → Zn^{2+} + 2e • Electrons move along wire towards copper. • At the copper electrode (positive electrode), hydrogen ions from water in NaOH(aq), gains electrons to form hydrogen gas.: $2H^+(aq)$ + $2e^-$ → $H_2(g)$ • **Concentration of H^+ ions in electrolyte decreases and hence resulting solution becomes more alkaline.**
2	D	• Zinc **loses electrons** (negative electrode; oxidation), **electrons move in wire towards copper**. Zinc **decreases in size.** Zn(s) → $Zn^{2+}(aq)$ + $2e^-$ • At the copper electrode, copper(II) ions in $CuSO_4(aq)$ gains electrons and form **pink/reddish brown copper**. $Cu^{2+}(aq)$ + $2e^-$ → Cu(s)
3	B	• The further apart the two metals are in the reactivity series, the greater the voltage produced. • Position of the metals in the reactivity series (most to least reactive): Mg, Zn, Fe, **Pb**, **Ag =>Position of lead is closer to silver in the reactivity series.**
4	C	**I : Electrons move from Ag to Cu. (False)** • Cu is more reactive than Ag. => Cu loses electrons more readily than Ag. **II: Colour intensity of $Cu(NO_3)_2(aq)$, blue solution increases. (True)** • Cu is oxidised to form Cu^{2+} ions that goes into the solution: Cu → Cu^{2+} + 2e Hence, **concentration of Cu^{2+} ions in the solution increases** **III: Cu is the reducing agent. (True)** • Cu electrode becomes smaller => **Cu loses electrons to Ag^+, reduces Ag^+ to Ag.** **IV:** Ag electrode becomes thinner. **(False)** • Electron flow from Cu to Ag • **Ag^+** (from silver nitrate solution) **gains electrons** from Cu to **form Ag**: Ag^+ + e → Ag => silver electrode **increases** in size.
5	B	• **Mg is more reactive than Cu** • **At Mg electrode:** - **Mg loses electrons more readily than Cu.** - **Mg is oxidised to form Mg^{2+} ions : Mg → Mg^{2+} + 2e** - **Mg electrode becomes smaller / decreases in size.** • **Electron flow from Mg to Cu** • **At Cu electrode:** - **H^+ in water from NaCl(aq) gains electrons to form hydrogen gas:** **$2H^+ + 2e \rightarrow H_2$** - **Effervescence of colourless, odourless gas formed at Cu electrode.** - **Note that Na^+ ions are high up in the electrochemical series and will remain as ions in the solution and will not be discharged.**

6 Fuel Cells

- A **fuel cell** is a chemical cell in which **reactants** **(usually a fuel and oxygen)** are **continuously supplied** to produce electricity directly.
- In a fuel cell, one electrode is supplied with a **fuel** (e.g. hydrogen) and the other electrode is supplied with **oxygen**. The two electrodes are immersed in an electrolyte.
- The chemicals in the fuel cells are continuously replaced as they are used up.
- The most common example of a fuel cell is the **hydrogen-oxygen fuel cell** which is used as a source of electrical power in space vehicles.
- The diagram below shows the **hydrogen-oxygen fuel cell.**

Refer to this link:
https://www.youtube.com/watch?v=8xeB_O_fyzM

battery



- The cell consists of three compartments that are separated from one another by porous **carbon electrodes with platinum** acting as a catalyst. The porous carbon electrodes allow the hydrogen gas and oxygen gas to come into contact with the electrolyte (a solution that ions can move through).
- **Hydrogen** gas is fed into one compartment and **oxygen** gas into the other:

At the **negative electrode (anode):** ***(oxidation occurs)***	At the **positive electrode (cathode):** ***(reduction occurs)***
Hydrogen gas **undergoes oxidation** to form **hydrogen ions.** Electrons released make the electrode negative. **Electrons** then **move across the wire** giving the electric current. The **hydrogen ions** formed which **move through the electrolyte** towards the positive electrode.	Oxygen gas **undergoes reduction** by **reacting** with the **hydrogen ions** and **electrons** that were formed at the negative electrode. Electrons are then removed at the electrode, making the electrode positive. Heat and **water** produced at the positive electrode **leaves the fuel cell at this outlet.**
$H_2(g) \rightarrow 2H^+(aq) + 2e$	$O_2(g) + 4H^+(aq) + 4e \rightarrow 2H_2O(l)$

- Combining the two half equations above shows that hydrogen and oxygen is converted to water :

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

This is equivalent to the combustion of hydrogen.

- Fuel cells **produce electricity indefinitely** provided there is continuous supply of fuel. This is unlike ordinary chemical cell (battery) which will be used up.
- Fuel cells have found uses in **space shuttles**, satellites and orbiting space stations. In case you are wondering, why specifically in space? **Water produced** can be used as drinking water for the astronauts in space.
- Comparing the **pros and cons of using hydrogen as a fuel**.

Advantages	**Disadvantages**
Combustion of hydrogen produces water, a non-polluting product. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$	There is no cheap sources of hydrogen - Obtaining hydrogen from electrolysis is an expensive process - Hydrogen is currently obtained from non-renewable energy sources such as cracking of petroleum
Hydrogen is considered a renewable source of energy. Hydrogen can be obtained from the electrolysis of water. $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$	Hydrogen is extremely flammable and explosive. Special precautions must be taken in the storage and transport of the gas.
Hydrogen is an efficient source of energy. When hydrogen burns, it releases more than twice the amount of energy compared to many fuels	Hydrogen is also very light and being a gas require large storage volumes compared to other fuels.