DEPARTMENT OF EDUCATION

GRADE 12

CHEMISTRY

MODULE 3

ELECTROCHEMISTRY

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GRADE 12

CHEMISTRY

MODULE 3

ELECTROCHEMISTRY

IN THIS MODULE YOU WILL LEARN ABOUT:

12.3.1: ELECTROLYTIC CELL
12.1.2: HALF-REACTIONS AT THE ELECTRODES
12.1.3: FACTORS AFFECTING THE PERFORMANCE OF ELECTROLYTIC CELL
12.1.4: ELECTROLYSIS USING INERT ELECTRODES
12.1.5: ELECTROLYSIS OF BRINE
12.1.6: MANUFACTURE OF ALUMINIUM FORM BAUXITE
12.3.7: ELECTROLYTIC REFINING OF IMPURE COPPER METAL
12.3.8: ELECTROPLATING
12.3.9: ELETROCHEMICAL CELL (GALVANIC CELL)

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DIANA TEIT AKIS
PRINCIPAL

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TABLE OF CONTENTS

			PAGE
TITLE		....	**1**
ISBN AND ACKNOWLEDGEMENTS		....	**2**
TABLE OF CONTENTS		....	**3**
SECRETARY'S MESSAGE		....	**4**
MODULE 12.3	**ELECTROCHEMISTRY**	....	**5**
Introduction		....	**5**
Learning Outcomes		....	**5**
Terminologies		....	**6**
12.3.1	**Electrolytic Cell**	....	**7**
	Electrolysis of Molten Salts	....	**7**
	Electrolysis of Aqueous Solution	....	**9**
12.3.2	**Half-reactions at the Electrodes**	....	**12**
12.3.3	**Factors Affecting the Performance of Electrolytic Cell**	....	**16**
	Reactivity of Ions	....	**16**
12.3.4	**Electrolysis Using Inert Electrodes**	....	**19**
	How Hoffman Voltameter Works Using Inert Electrodes	....	**20**
12.3.5	**Electrolysis of Brine**	....	**24**
12.3.6	**Manufacture of Aluminium From Bauxite**	....	**25**
12.3.7	**Electrolytic Refining of Impure Copper Metal**	....	**29**
12.3.8	**Electroplating**	....	**30**
12.3.9	**Electrochemical Cell (Galvanic Cell)**	....	**33**
SUMMARY		....	**38**
ANSWERS TO LEARNING ACTIVITIES		....	**40**
REFERENCES AND APPENDICES		....	**44**

SECRETARY'S MESSAGE

Achieving a better future by individual students and their families, communities or the nation as a whole, depends on the kind of curriculum and the way it is delivered.

This course is a part of the new Flexible, Open and Distance Education curriculum. The learning outcomes are student-centred and allows for them to be demonstrated and assessed.

It maintains the rationale, goals, aims and principles of the national curriculum and identifies the knowledge, skills, attitudes and values that students should achieve.

This is a provision by Flexible, Open and Distance Education as an alternative pathway of formal education.

The course promotes Papua New Guinea values and beliefs which are found in our Constitution, Government Policies and Reports. It is developed in line with the National Education Plan (2005 -2014) and addresses an increase in the number of school leavers affected by the lack of access into secondary and higher educational institutions.

Flexible, Open and Distance Education curriculum is guided by the Department of Education's Mission which is fivefold:

- To facilitate and promote the integral development of every individual
- To develop and encourage an education system that satisfies the requirements of Papua New Guinea and its people
- To establish, preserve and improve standards of education throughout Papua New Guinea
- To make the benefits of such education available as widely as possible to all of the people
- To make the education accessible to the poor and physically, mentally and socially handicapped as well as to those who are educationally disadvantaged.

The college is enhanced through this course to provide alternative and comparable pathways for students and adults to complete their education through a one system, two pathways and same outcomes.

It is our vision that Papua New Guineans' harness all appropriate and affordable technologies to pursue this program.

I commend all the teachers, curriculum writers and instructional designers who have contributed towards the development of this course.

UKE KOMBRA, PhD
Secretary for Education

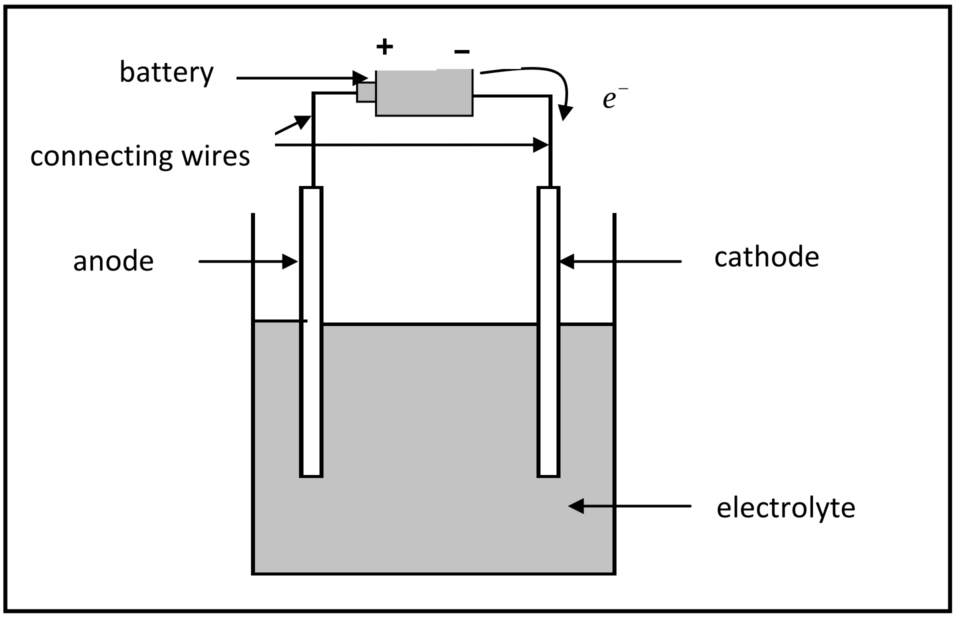
Terminologies

Before you get into the thick of things, let us make sure you know some of the terminologies that are used throughout this module.

Anode	This is the positively charge electrode. It is connected to the positive terminal of the power source.
Anions	These are negatively charge ions. Anions are attracted to the anode during electrolysis.
Cathode	This is the negatively charged electrode. It is connected to the negative terminal of the power source.
Cations	These are the positively charged ions. Cations are attracted to the cathode during electrolysis.
Electrode	A conductor which dips into an electrolyte and allows the current to pass through.
Electrolyte	It is a molten (melted) ionic compound or an aqueous solution which contains ions.
Inert electrodes	Electrodes which do not react with the products or electrolyte during electrolysis. Substances like platinum and graphite (carbon) are normally used as inert electrodes.
Reactive electrodes	Electrodes that react with the electrolyte during electrolysis.

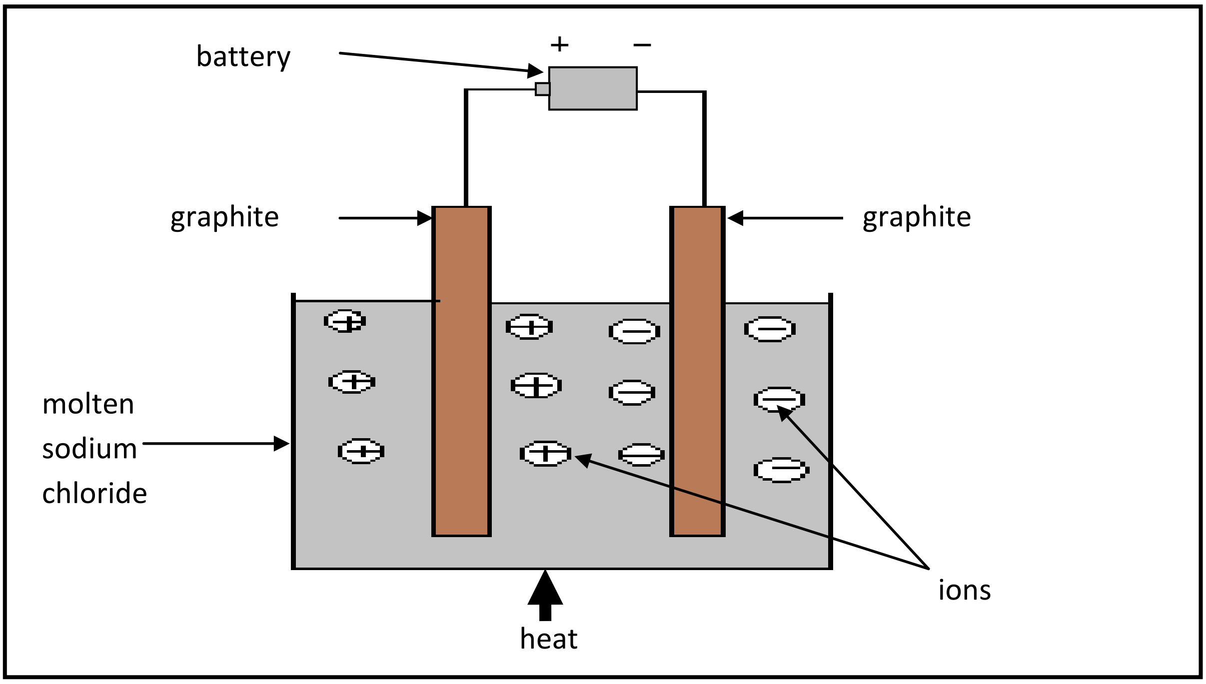
12.3.1 Electrolytic Cell

An electrolytic cell is a vessel (container) in which electrolysis take place. It is made up of a power source, electrolyte and electrodes.



A simple electrolytic cell

Electrolysis of Molten Salts
A molten salt is a compound that is melted, and it is usually an ionic compound.



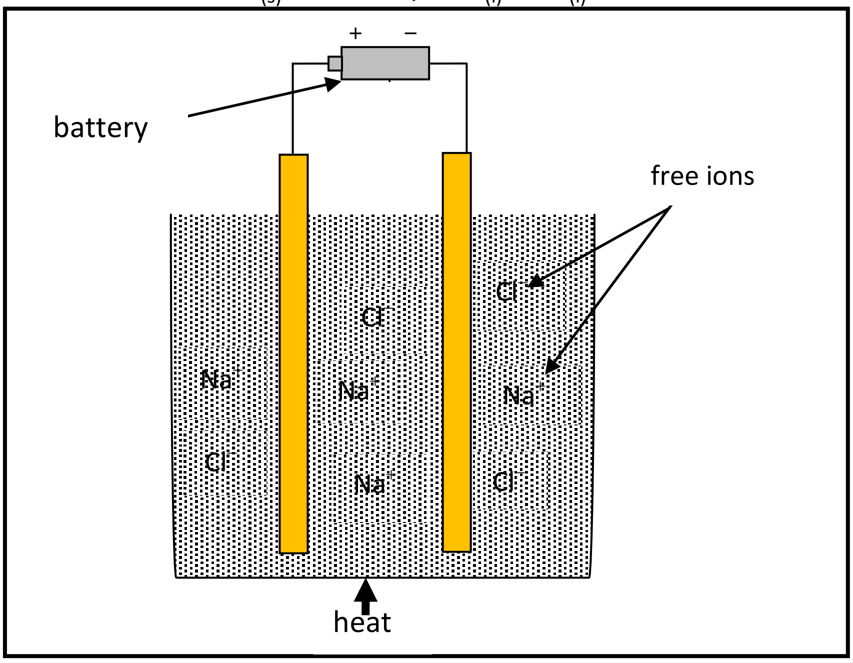
Electrolysis of molten compounds

When an ionic compound is melted, the ionic structure breaks down and the ions which consist of positive and negative charges separate and move apart.

If an electric current is passed through it (molten compound), the positive ions (cations) will move to the negative electrode (cathode), while the negative ions (anions) will move to the positive electrode (anode).On reaching the cathode, the positive ions will gain electrons and become metal atoms, while the negative ions will lose electrons at the anode and become non–metallic atoms.

Example:
A molten sodium chloride is formed when heat is applied to melt it to form the ions. The ions formed by molten sodium chloride are sodium ions (Na^+) chloride and ions (Cl^-).

$$NaCl_{(s)} \xrightarrow{heat} Na^+_{(l)} + Cl^-_{(l)}$$



Electrolysis of molten sodium chloride

At the anode
During electrolysis, the negative chloride ions (Cl^-) are attracted to the anode. At the anode two chloride ions lose an electron each to form chlorine gas. We say that the chloride ions are discharged.

Half equation at anode:
$2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^-$

When a substance **loses electrons**, it is experiencing **oxidation** reaction.

At the cathode

During electrolysis, the positive sodium ions (Na^+) are attracted to the cathode. At the cathode one sodium ion gain an electron and becomes a sodium atom (Na). We say that the sodium ion is discharged.

Half equation at cathode:

$2Na^+_{(l)} + 2\,e^- \longrightarrow 2Na_{(l)}$

When a substance **gains electrons**, it is experiencing **reduction** reaction.

Overall equation:

$2NaCl_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2\,(g)}$

In the electrolysis of molten compounds, there will be two ions only, the positive ion (cation) and the negative ion (anion).

Reduction gains electrons at the cathode while oxidation loses electrons at the anode.

The table below shows some molten salts undergoing electrolysis.

		Half equations	
Molten Salt	**Ions formed**	**Cathode**	**Anode**
Calcium chloride, $CaCl_2$	Ca^{2+},Cl^-	$Ca^{2+}_{(l)} + 2e^- \longrightarrow Ca_{(l)}$	$2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^-$
Sodium iodide, NaI	Na^+,I^-	$Na^+_{(l)} + e^- \longrightarrow Na_{(l)}$	$2I^-_{(l)} \longrightarrow I_{2(l)} + 2e^-$
Lead(II) bromide, $PbBr_2$	Pb^{2+},Br^-	$Pb^{2+}_{(l)} + 2e^- \longrightarrow Pb_{(l)}$	$2Br^-_{(l)} \longrightarrow Br_{2(g)} + 2e^-$

Electrolysis of aqueous solution

An aqueous salt is formed when a salt (ionic compound) dissolves in water. When an ion compound is dissolved in water, the ions are separated and are free to move around.

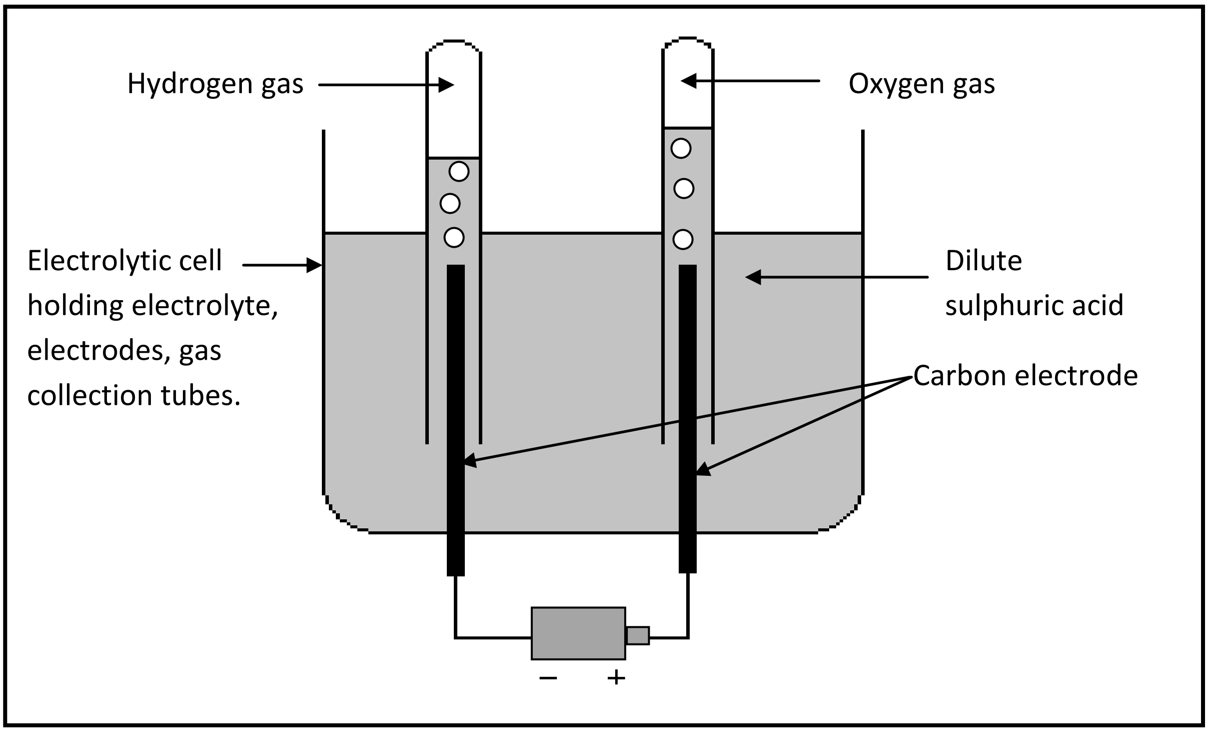
If an electric current is passed into the solution, the positive ions of the compound will move to the cathode, while the negative ions will move to the anode. However, unlike molten compounds, water is present in the solution and it ionizes to produce hydrogen ions (H^+) and hydroxide ions (OH^-). These ions also move to the respective electrodes and compete with the ions of the electrolyte for discharge.

It becomes more difficult to determine the ions which are discharged at each electrode. Therefore, we can use the **law of selective discharge** to determine the ion that is discharged at each electrode.

Example: Electrolysis of dilute sulphuric acid (H_2SO_4).

Pure water does not conduct electricity. However, when some dilute sulphuric acid is added, it will conduct electricity. The electrolysis of dilute sulphuric acid is similar to electrolysis of acidified water.

Dilute sulphuric acid contains the following ions: hydrogen ions (H^+), sulphate ions (SO_4^{2-}) from the acid and hydrogen ions (H^+), hydroxide ions (OH^-) from water.



Electrolysis of dilute sulphuric acid

At the cathode
Hydrogen ions are attracted here and they are discharged by gaining electrons.

This is the half equation which shows this reaction.

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$$

At the anode
Hydroxide ions and sulphate ions will move to the anode. However, it is the hydroxide ions that are discharged and not sulphate ions.

This is the half equation which shows this reaction.

$$4OH^-_{(aq)} \longrightarrow O_{2\,(g)} + 2H_2O_{(l)} + 4e^-$$

The undischarged sulphate and hydrogen ions in the water recombine to form sulphuric acid. Thus, as the water is electrolyzed and the sulphuric acid in the solution becomes more concentrated.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 1 **30 minutes**

Answer the following questions.

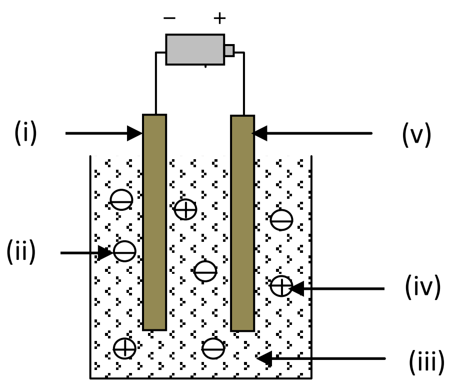
1. A. Define the following:

(i) Electrolysis ____

(ii) Anode ____

(iii) Cathode ____

B. Label the diagram with words connected with electrolysis: anode, cathode, electrolyte, anion and cation.



(i) ________________________

(ii) ________________________

(iii) ________________________

(iv) ________________________

(v) ________________________

2. During electrolysis, to which electrode do the

   a) positive ions of the electrolyte move? ________________________

   b) negative ions of the electrolyte move? ________________________

3. Write gain or lose in the following electrodes.

   a) cathode __________
   b) anode __________

2. What are inert electrodes?
   ____
   ____

5. Copper (II) chloride solution is electrolyzed using carbon (graphite) electrodes.

   a) List all the ions present in the electrolyte.
   ____

   b) Which ions will move to the cathode during electrolysis?
   ____

   c) Which ion is discharged at the cathode? Give a reason for your answer.
   ____

   d) Write the cathode reaction.
   ____

   e) Write the anode reaction.
   ____

Thank you for completing your learning activity 1. Check your work. Answers are at the end of this module.

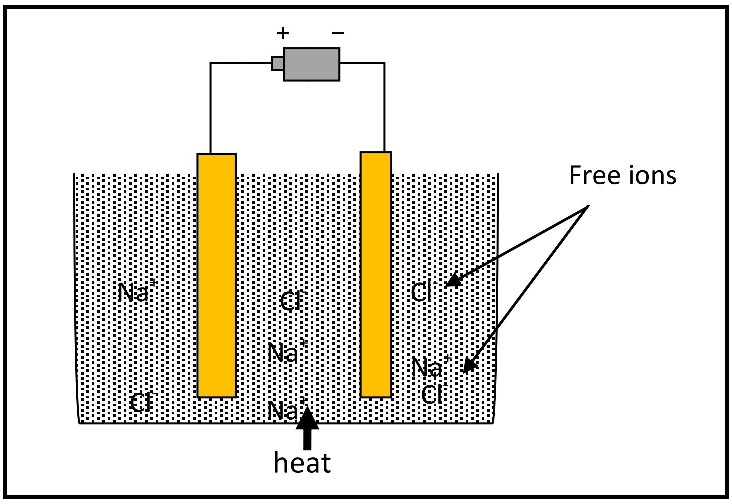
12.3.2 Half-Reactions At The Electrodes

Oxidation and reduction (Redox) are two reactions that always happen together. Oxidation involves reactants losing electrons while in reduction electrons are gained or received by the reactants.

During electrolysis, redox reactions happen at the electrodes. The reactions at each electrode are called **half equations**. The half equations are written so that the same number of electrons are lost and gained in the reaction.

Example 1:

A molten sodium chloride is formed when heat is applied to melt it to form the ions. The ions formed by molten sodium chloride are sodium ions (Na^+) chloride and ions (Cl^-).



$$NaCl_{(s)} \xrightarrow{heat} Na^+_{(l)} + Cl^-_{(l)}$$

Electrolysis of molten sodium chloride

At the cathode

During electrolysis, the positive sodium ions are attracted to the cathode. At the cathode one sodium ions gain electrons and become sodium atom. We say that the sodium ions are discharged. Reduction reaction occurs here at the cathode.

Half equation:

$$Na^+_{(l)} + e^- \longrightarrow Na_{(l)}$$

At the anode

During electrolysis, the negative chloride ions are attracted to the anode .At the anode two chloride ions lose an electron each to form chlorine gas. We say that the chloride ions are discharged. Oxidation reaction has taken place here at the anode.

Half equation:

$$2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^-$$

Overall equation

You can also write the overall equation for the electrolysis of molten sodium chloride. This is done by combining the cathode and the anode reactions. You must make sure that the **number of electrons** at the cathode is the **same** as the number of electrons at the anode.

When you look at the electrolysis of sodium chloride, there is only one electron at the cathode and two electrons at the anode. Therefore you need to multiply the cathode reaction by 2 to have the same number of electrons at the cathode and the anode as shown below.

$2(Na^{+}_{(l)} + e^{-} \longrightarrow Na_{(l)})$
$2Na^{+}_{(l)} + 2e^{-} \longrightarrow 2Na_{(l)}$

Now you have the same number of electrons at the cathode and anode. Combine the reactants and products for the cathode and anode reactions.

$2Na^{+}_{(l)} + 2e^{-} + 2Cl^{-}_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2\,(g)} + 2e^{-}$

Subtract the same number of electrons on both sides of the equation. For this case we subtract $2e^{-}$ on both sides of the equation.

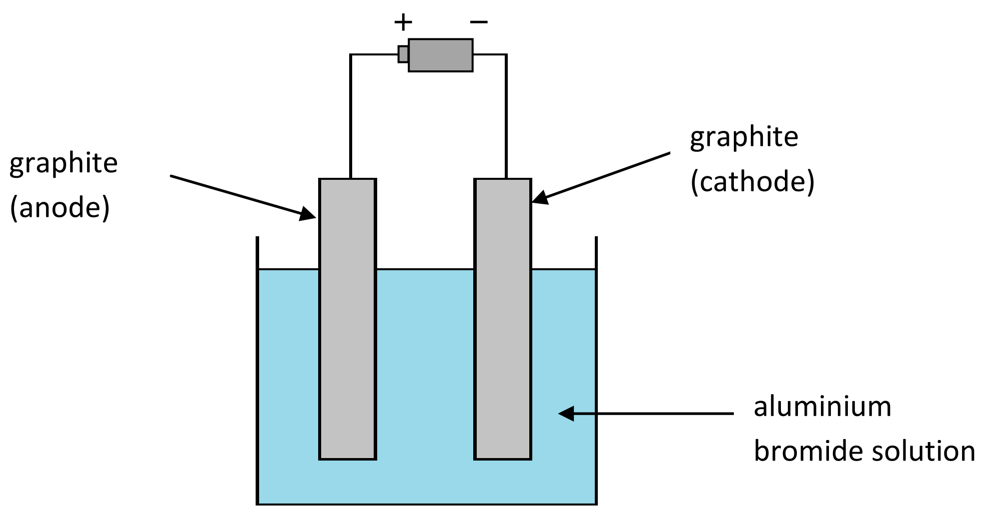
$2Na^{+}_{(l)} + \cancel{2e^{-}} + 2Cl^{-}_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2\,(g)} + \cancel{2e^{-}}$
$2Na^{+}_{(l)} + 2Cl^{-}_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2\,(g)}$

You need one sodium ion and one chloride ion to form one sodium chloride. Since there are two sodium ions and two chloride ions, this gives two sodium chloride.

Therefore, the overall equation becomes:

$2NaCl_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2\,(g)}$

Example 2: Electrolysis of dilute solution of aluminium bromide.



Electrolysis of solution of aluminium bromide.

The ions present in the electrolyte are: Al^{3+}, Br^{-}(from aluminium bromide) and H^{+}, OH^{-} (from water).

At the Cathode

Both Al^{3+} and H^{+} move here.

Hydrogen ion (H^{+}) is discharged instead of aluminium ion ($Al^{3+)}$ because it is an ion of a less reactive element.

Half equation:

$H^{+}_{(aq)} + 2e^{-} \longrightarrow H_{2\,(g)}$

Hydrogen gas is produced at the cathode.

At the Anode

Both Br^{-} and OH^{-} move here. Hydroxide ion (OH^{-}) is discharged instead of bromide ion.

Half equation:

$4OH^{-}_{(aq)} \longrightarrow O_{2\,(g)} + 2H_2O_{(l)} + 4e^{-}$

Oxygen gas and water are produced at the anode.

Overall equation

There are four electrons ($4e^{-}$) for the anode reaction so we must multiply the cathode reaction by two to get four electrons at the cathode as shown below.

$2(2H^{+}_{(aq)} + 2e^{-} \longrightarrow H_{2\,(g)})$

$4H^{+}_{(aq)} + 4e^{-} \longrightarrow 2H_{2\,(g)}$

Combine the cathode and anode reactions together.

$4H^{+}_{(aq)} + 4e^{-} + 4OH^{-}_{(aq} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)} + 2H_2O_{(l)} + 4e^{-}$

Subtract $4e^{-}$ on both sides of the equation.

$4H^{+}_{(aq)} + \cancel{4e^{-}} + 4OH^{-}_{(aq} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)} + 2H_2O_{(l)} + \cancel{4e^{-}}$

$4H^{+}_{(aq)} + 4OH^{-}_{(aq} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)} + 2H_2O_{(l)}$

One H^{+} and one OH^{-} will produce one H_2O.By combining $4H^{+}$ and 4 OH^{-} we will get $4H_2O$.

$4H_2O_{(l)} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)} + 2H_2O_{(l)}$

Subtract $2H_2O$ on both sides of the equation: $4H_2O - 2H_2O = 2H_2O$.

Therefore, overall equation is:

$$2H_2O_{(l)} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)}$$

12.3.3 Factors Affecting The Performance Of Electrolytic Cells

The performance of electrolytic cells is also affected by some factors. These are:
a) reactivity of ions.
b) concentration of aqueous solution.
c) nature of electrodes.

Reactivity of Ions
At the Cathode
If more than one type of ions is present, the discharge of the positive ions depends on the reactivity of the metal in the reactivity series.

- Ions of hydrogen and metals which are less reactive, like silver and copper, are discharged.
- Ions of reactive metals like potassium, sodium, calcium, magnesium and aluminium, are not discharged.

At the Anode
If more than one type of ions is present, the discharge of the negative ions follows this sequence:

1) For iodide, bromide and chloride ions are most easily discharged before oxygen gas (for hydroxide ions).
2) For diluted solutions, oxygen gas (hydroxide ions are easily discharged before iodide, bromide and chloride ions).
3) Nitrate ions and sulphate ions are not discharged.

Law of selective discharge of ions

The Law of Selective Discharge of Ions is related to the Electrochemical Series of Ions.

Electrochemical Series of Anions (Negative ions)

SO_4^{2-}, NO_3^- — Almost never take part (never discharged).

I^-, Br^-, Cl^- — Discharged if solution is concentrated.

OH^- — Nearly always takes part for diluted solution (always discharged).

Electrochemical Series Cations (Positive ions)

Cation	Discharge of Ions	Properties
K^+	$K^+ + e^- \rightarrow K$	Almost never.
Na^+	$Na^+ + e^- \rightarrow Na$	
Ca^{2+}	$Ca^{2+} + 2e^- \rightarrow Ca$	Discharge ability increases.
Mg^{2+}	$Mg^{2+} + 2e^- \rightarrow Mg$	
Al^{3+}	$Al^{3+} + 3e^- \rightarrow Al$	
Zn^{2+}	$Zn^{2+} + 2e^- \rightarrow Zn$	
Fe^{2+}	$Fe^{2+} + 2e^- \rightarrow Fe$	
Sn^{2+}	$Sn^{2+} + 2e^- \rightarrow Sn$	
Pb^{2+}	$Pb^{2+} + 2e^- \rightarrow Pb$	
H^+	$2H^+ + 2e^- \rightarrow H_2$	
Cu^{2+}	$Cu^{2+} + 2e^- \rightarrow Cu$	
Hg^{2+}	$Hg^{2+} + 2e^- \rightarrow Hg$	Most easily discharged.
Ag^+	$Ag^+ + e^- \rightarrow Ag$	

The ions found at the top of the series are more reactive than the ones at the bottom of the series. Moreover, less reactive ions are discharged easily than the reactive ions.

Concentration of aqueous solution

The electrolysis of aqueous solutions is also affected by the concentrations of the ions present in the solution.

If a solution is very **dilute**, the presence of ions from the ionic compound only helps the water to conduct electricity, and the electrolysis is essentially that of the electrolysis of water.

On the other hand, if **a solution is concentrated**, the **anions** which are in **greater concentration** compared to the hydroxide ions will be discharged.

Example 1

Electrolysis of very dilute sodium chloride solution using inert electrodes. The ions present in sodium chloride solution are:

Na^+, Cl^-(from sodium chloride solution) and H^+, OH^- (from water)

At the Cathode

Both Na^+ and H^+ move to the cathode. According to the Law on Selective Discharge of Ions, sodium is higher in the reactivity series, so Na^+ ions are not discharged. Instead H^+ ions from water are discharged to form hydrogen gas:

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$$

Hydrogen gas is produced at the cathode. Hydrogen ions **gain** electrons to become hydrogen gas. Therefore, reduction reaction occurs at the cathode.

At the Anode
Both Cl^- and OH^- move to the anode. According to the Law on Selective Discharge of Ions, OH^- ions are discharged in preference over Cl^- ions to form oxygen gas and water.

$$4OH^-_{(aq)} \longrightarrow O_{2\,(g)} + 2H_2O_{(l)} + 4e^-$$

Here, hydroxide ions **lose** electrons to form oxygen gas and water. Therefore, oxidation reaction occurs at the anode.

The overall reaction is:

$$2H_2O_{(l)} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)}$$

Thus, the electrolysis of very dilute sodium chloride solution is essentially the electrolysis of water: hydrogen gas is given off at the cathode and oxygen gas is given off at the anode.

Example 2
Electrolysis of **concentrated** sodium chloride solution using inert electrodes. The ions present in sodium chloride solution are:

Na^+, Cl^-(from sodium chloride solution) and H^+, OH^- (from water)

At the Cathode
Both Na^+ and H^+ move to the cathode. According to the Law on Selective Discharge of Ions, sodium is higher in the reactivity series, so Na^+ ions are not discharged. Instead H^+ ions from water are discharged:

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$$

Hydrogen gas is produced at the cathode. Hydrogen ions are going through reduction. reaction.

At the Anode
Both Cl^- and OH^- move to the anode. According to the Law on Selective Discharge of Ions, OH^- ions should be discharged. However, due to the **high concentration of Cl^- ions,** they are discharged in preference over the OH^- ions.

$$2Cl^-_{(aq)} \longrightarrow Cl_{2\,(g)} + 2e^-$$

Chlorine gas is produced at the anode. Chloride ions are experiencing oxidation reaction.
The undischarged sodium ions (Na^+) and hydroxide ions (OH^-) recombine to form sodium hydroxide, thus making the solution alkaline.

The electrolysis of concentrated sodium chloride solution is used in the industry to make hydrogen, chlorine and sodium hydroxide.

> It is important to always refer to the **Law on Selective Discharge of Ions** when dealing with electrolysis of aqueous solution. Also consider whether solution is **concentrated** or **not concentrated** (**diluted**).

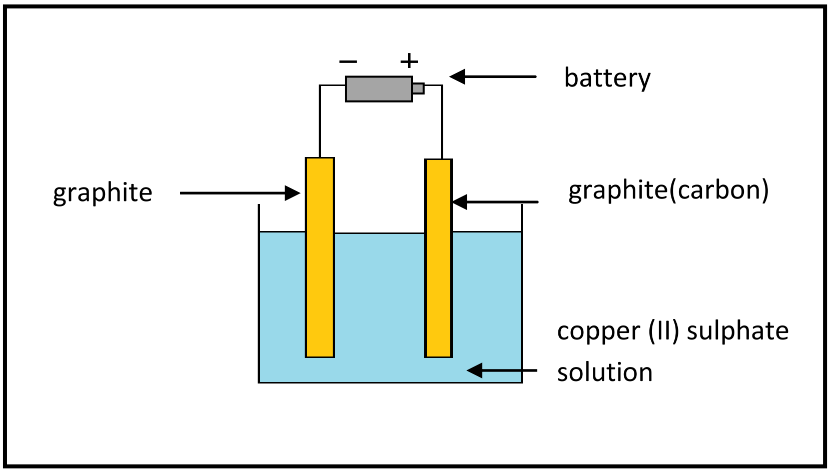
Nature of electrodes

Electrolysis using a reactive electrode will give a different product at the electrode while using inert electrode does not affect the kind of products formed at the electrodes.

12.3.4 Electrolysis Using Inert Electrodes

Inert electrodes are electrodes which do not react with the products or electrolyte during electrolysis. Substances like platinum and graphite (carbon) are normally used as inert electrodes.

Example: Electrolysis of copper (II) sulphate solutions using carbon electrodes



Electrolysis of copper (II) solution using inert electrodes.

The ions present in the copper (II) sulphate solution are: Cu^{2+}, SO_4^{2-}(from copper (II) sulphate) and H^+, OH^- (from water).

At the Cathode

Both Cu^{2+} and H^+ move to the cathode. According to the Law on Selective Discharge of Ions, copper is lower in the reactivity series, so Cu^{2+} ions are discharged in preference over H^+ ions:

$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$

A deposit of copper metal is formed at the cathode.

At the Anode

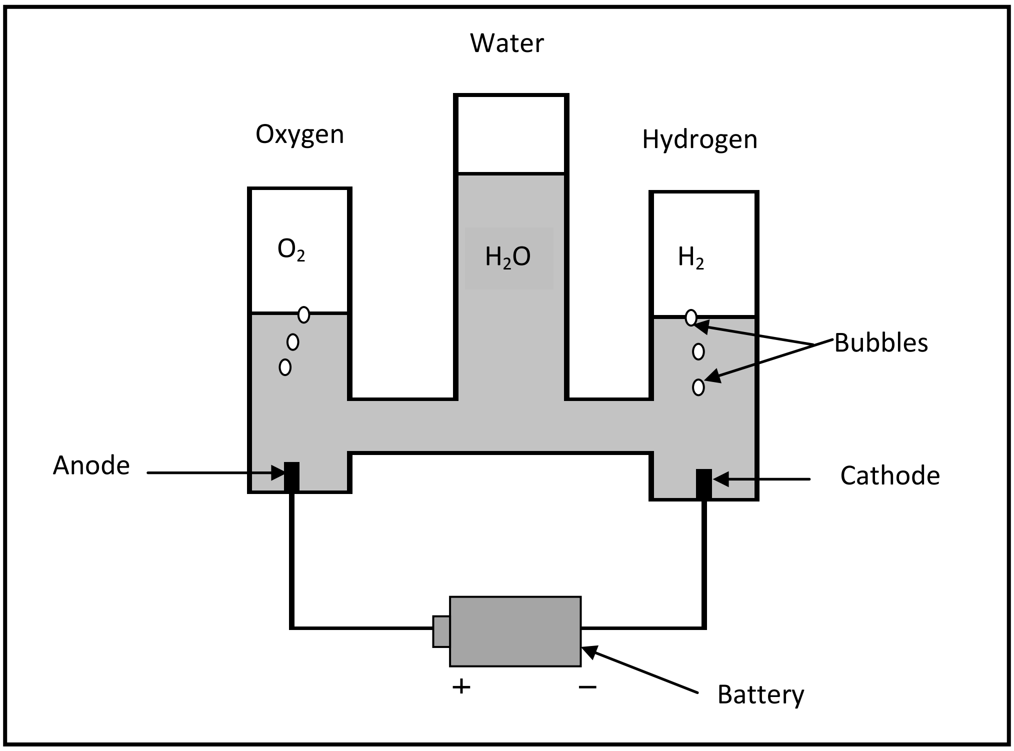
Both SO_4^{2-} and OH^- move to the anode. Based on the Law on Selective Discharge of Ions, SO_4^{2-} ions are not discharged, instead OH^- ions are discharged to form oxygen and water:

$$4OH^-_{(aq)} \longrightarrow O_{2\,(g)} + 2H_2O_{(l)} + 4e^-$$

Oxygen gas and water are produced at the anode.

How a Hoffman Voltameter works using inert electrodes

Acid is added to water to acidify it. The acidified water can be electrolyzed using **Hofmann Voltameter** as shown in the diagram below. Pure water contains only few ions but when mixed with acid it can conduct well.



Hofmann Voltameter

At the Cathode

Hydrogen ions move to the cathode. At the cathode, two hydrogen ions gain one electron each to form hydrogen gas.

Half equation:

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$$

You will see bubbles at the cathode since hydrogen gas is produced.

At the Anode

Both sulphate ion and hydroxide ion will move to the anode. At the anode, four hydroxide ions will lose four electrons to form oxygen gas. According to the law on selective discharge of ions, sulphate ions are not discharged, instead hydroxide ions are discharged to form oxygen and water:

$4OH^{-}_{(aq)} \longrightarrow O_{2\,(g)} + 2H_2O_{\,(l)} + 4e^{-}$

Oxygen gas and water are produced at the anode.

Overall equation

We can also write the overall equation for the electrolysis of acidified water by combining the cathode and anode reaction. There are four electrons for the anode reaction so we must multiply the cathode reaction by two to get four electrons at the cathode as shown below.

$2(2H^{+}_{(aq)} + 2e^{-} \longrightarrow H_{2(g)})$

$4H^{+}_{(aq)} + 4e^{-} \longrightarrow 2H_{2(g)}$

Combine the cathode and anode reactions together.

$4H^{+}_{(aq)} + 4e^{-} + 4OH^{-}_{(aq} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)} + 2H_2O_{\,(l)} + 4e^{-}$

Subtract $4e^{-}$ on both sides of the equation.

$4H^{+}_{(aq)} + \cancel{4e^{-}} + 4OH^{-}_{(aq} \longrightarrow 2H_{2(g)} + O_{2(g)} + 2H_2O_{(l)} + \cancel{4e^{-}}$

$4H^{+}_{(aq)} + 4OH^{-}_{(aq} \longrightarrow 2H_{2\,(g)} + O_{2(g)} + 2H_2O_{\,(l)}$

One H^{+} and one OH^{-} will produce one H_2O. By combining $4H^{+}$ and 4 OH^{-} we will get $4H_2O$.

$4H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2\,(g)} + 2H_2O_{\,(l)}$

Subtract $2H_2O$ on both sides of the equation.

$2H_2O_{(l)} \longrightarrow 2H_{2\,(g)} + O_{2\,(g)}$

You can see that in the electrolysis of sulphuric acid, it is water that undergoes electrolysis and not sulphuric acid. And the inert electrodes remain as they are.

Therefore, the overall equation shows the decomposition of water as shown above. The mole ratio of hydrogen to oxygen is 2:1.Therefore the volume of hydrogen is twice the volume of oxygen.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 2 **30 minutes**

Answer the following questions.

1. Molten magnesium chloride is electrolyzed using the platinum electrodes.

   a) List all the ions present in the electrolyte.

   ____

   b) Which ion will move to the cathode during electrolysis?

   ____

   c) Which reaction happens at the cathode, reduction or oxidation?

   ____

   d) Write the half equations that represent the reactions.

      (i) At cathode

      ____

      (ii) At anode

      ____

   e) Write the overall equation.

   ____

2. Sodium sulphate solution is electrolyzed using graphite electrodes.

   (a) List all the ions in a sodium sulphate solution.

   ____

(b) Write the equation at the

(i) cathode

____

(ii) anode

____

(iii) over all reaction

____

3. Concentrated potassium chloride is electrolyzed using inert electrodes.

a) List all the ions present in the solution.

____

b) Which ion is discharged at the

(i) cathode? Give a reason for your answer.

____

(ii) anode? Give a reason for your answer.

____

c) Write the reaction in

(i) cathode.

____

(ii) anode.

____

(iii) overall.

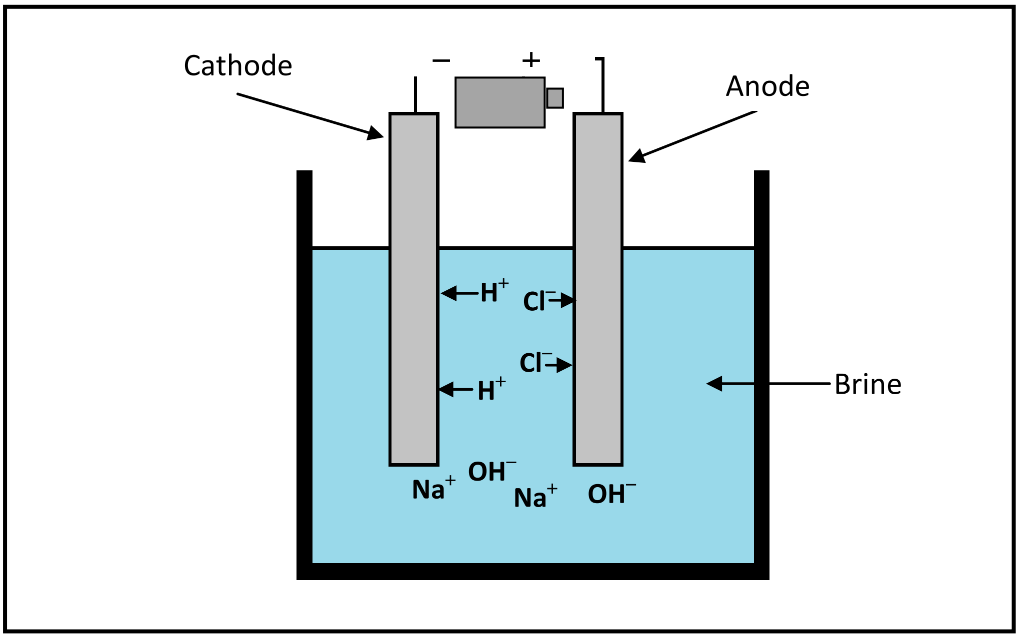
____

Thank you for completing your learning activity 2. Check your work. Answers are at the end of this module.

12.3.5 Electrolysis of Brine

Sodium chloride dissolved in water is called brine. It has a high concentration of sodium chloride. The ions found in this solution are sodium ion (Na^+), chloride ion(Cl^-) hydroxide ion(OH^-) and hydrogen ion(H^+).

The diagram below shows what happens to the ions during the electrolysis of brine using inert electrode.



At the Cathode

Hydrogen ions move to the cathode. At the cathode, two hydrogen ions gain one electron each to form hydrogen gas.

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$$

Hydrogen ions gain electrons (reduction) to form hydrogen atoms. The hydrogen atoms combine to form molecules of hydrogen gas.

At the Anode

Chloride ions will move to the anode. At the anode, two chloride ions will lose one electron each to form chlorine gas.

$$2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$$

Chloride ions lose electrons (oxidation) to form chlorine atoms. The chlorine atoms combine to form chlorine gas.

Sodium ions and hydroxide ions in the solution combine and form sodium hydroxide, thus, making the solution become an alkali.

Overall equation:

$2H^{+}_{(aq)} + 2e^{-} \longrightarrow H_{2\,(g)})$

$2Cl^{-}_{(aq)} \longrightarrow Cl_{2\,(g)} + 2e^{-}$

Combine the cathode and anode reactions together.

$2H^{+}_{(aq)} + 2e^{-} + 2Cl^{-}_{(aq} \longrightarrow H_{2\,(g)} + Cl_{2\,(g)} + 2e^{-}$

Since the two equations have same number of electrons, no need to multiply. Subtract $2e^{-}$ on both sides of the equation.

$2H^{+}_{(aq)} + \cancel{2e^{-}} + 2Cl^{-}_{(aq} \longrightarrow H_{2(g)} + Cl_{2(g)} + \cancel{2e^{-}}$

$2H^{+}_{(aq)} + 2Cl^{-}_{(aq} \longrightarrow H_{2\,(g)} + Cl_{2\,(g)}$

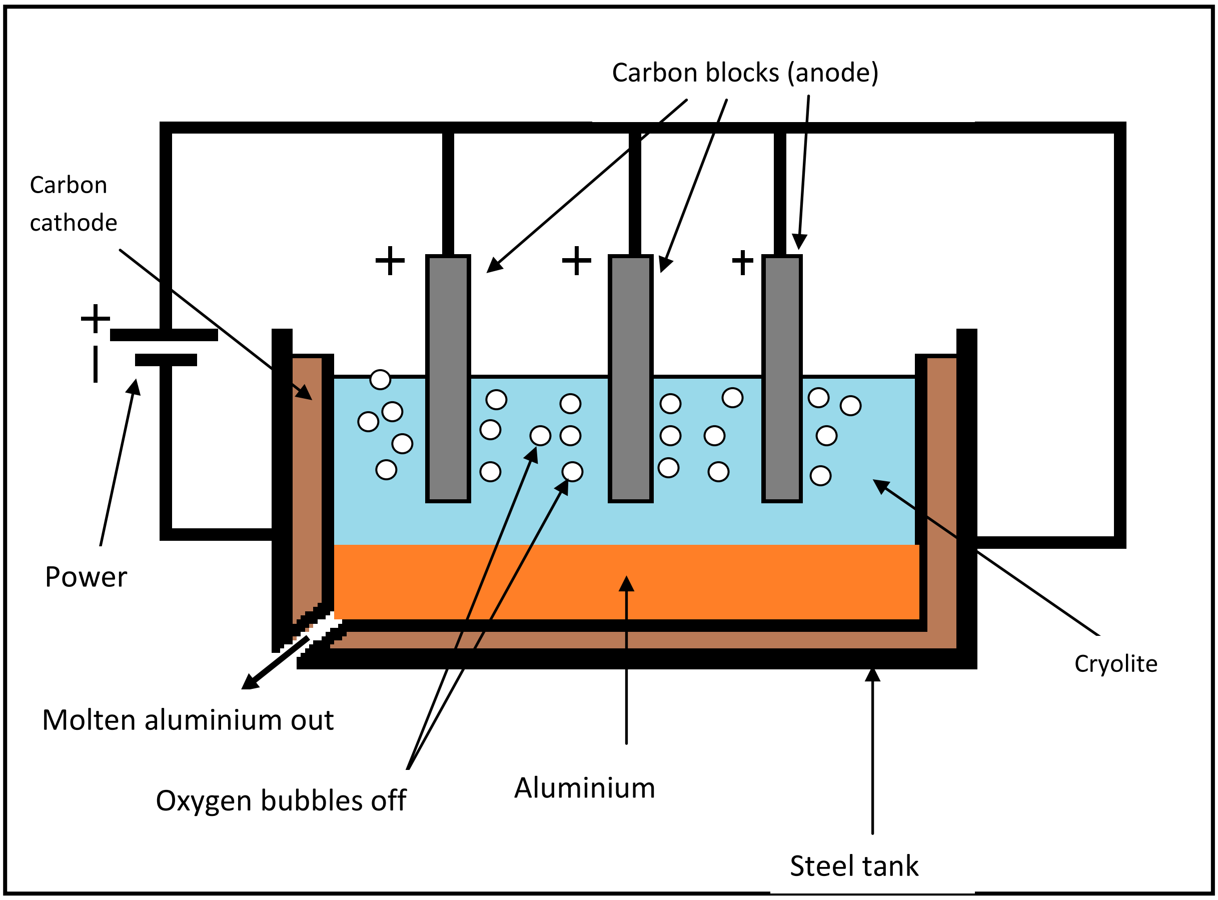
The electrolysis of brine is a large-scale process used to manufacture chlorine from salt. Two other useful chemicals are obtained during the process are sodium hydroxide (NaOH) and Hydrogen (H_2).

12.3.6 Manufacture of Aluminium from Bauxite

Aluminium is the most abundant metal in the Earth's crust. Its main ore is **bauxite**, which is aluminium oxide mixed with impurities such as sand and iron ore. Bauxite is reddish-brown in colour. The ore is taken to a bauxite plant where impurities are removed. The result is white **aluminium oxide** or a**lumina**.

The electrolysis is carried out in a large steel tank. The sides of the tank are lined with carbon which acts as the cathode (–). Big blocks of carbon hang in the middle of the tank and act as the anode (+).

Alumina melts at $2045^{0}C$. It would be impossible to keep the tank that hot. Instead, the alumina is dissolved in molten **cryolite** or sodium aluminium fluoride which has a much lower melting point.



Manufacturing of aluminium by electrolysis process

The reactions at the electrodes

Once the alumina dissolves, its aluminium and oxide ions are free to move. They move to the electrode of opposite charge.

At the cathode: The aluminum ion gains electrons. Reduction reaction occurs here.

$$4Al^{3+}_{(l)} + 12e^{-} \longrightarrow 4Al_{(l)}$$

The pure aluminium drops to the bottom of the cell as molten metal. This is run off at an outlet from the steel tank. Some of the pure aluminium metal will be mixed with other metals to make **alloys**. Some are moulded to harden into blocks.

At the anode:
The oxygen ions lose electrons in the oxidation reaction.

$$6O^{2-}_{(l)} \longrightarrow 3O_{2\,(g)} + 12e^{-}$$

The oxygen gas bubbles off, and reacts with the anode.

$$C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)}$$

So the carbon blocks get used up and need to be replaced.

Overall equation:
The alumina is broken down giving aluminium.

Aluminium oxide $\longrightarrow$ Aluminium + Oxygen

$$2Al_2O_{3\,(l)} \longrightarrow 4Al_{(l)} + 3O_{2\,(g)}$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 3 **30 minutes**

Answer the following questions.

1. Define the following:
   (i) Brine
   ____
   ____

   (ii) Bauxite
   ____
   ____

2. a) Sketch a simple electrolytic cell with parts labeled correctly showing electrolysis of brine.

b) Write the half equation:

(i) cathode

____

(ii) anode

____

c) Write the overall equation

____

3. (a) Sketch and label simple electrolytic cell which shows the extraction of aluminium.

(b) Why does aluminium ions move to the cathode?

____

(c) What happens at the cathode? Give an equation.

____

(d) The anode is replaced regularly. Why?

____

Thank you for completing your learning activity 3. Check your work. Answers are at the end of this module.

12.3.7 Electrolytic Refining of Impure Copper

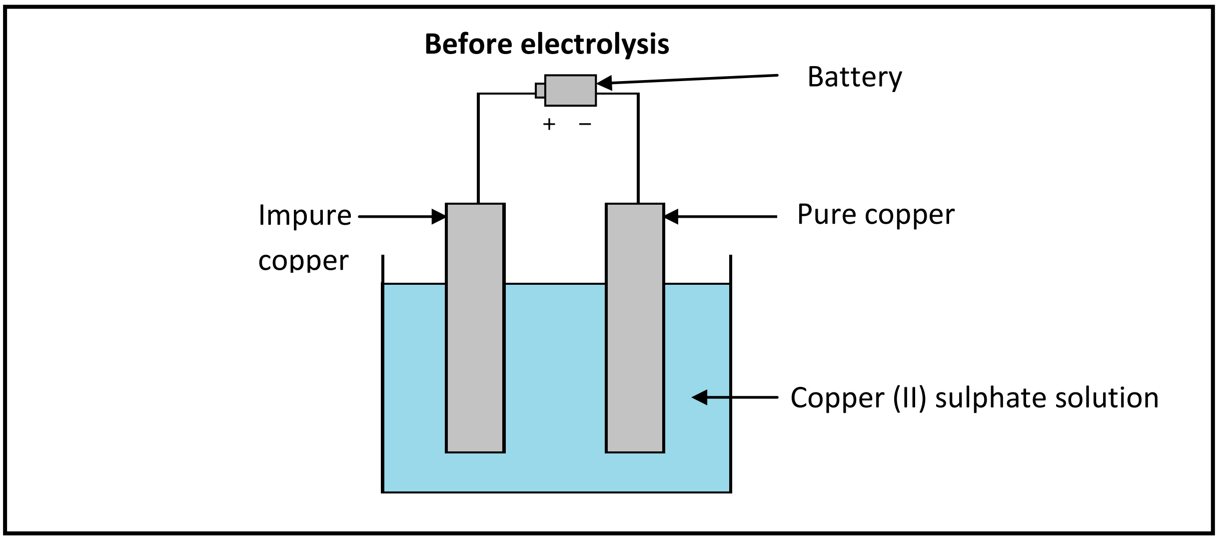
We have looked at electrolysis of an aqueous solution and molten salts using inert electrodes. Electrolysis using a reactive electrode will give a different product.

Reactive electrodes are used for two reasons:
1. Electrolytic refining of a metal.
2. Electroplating of an object.

Electrolytic refining of impure copper metal
Copper is a very good conductor of electricity. However, a small amount of impurities can reduce the conductivity of copper. Copper must be 99.9% pure to be able to conduct well. Impurities in copper can be removed by electrolysis.

Example of electrolytic refining is the purification of copper as show below.
The anode is the impure metal. The cathode is the pure metal and the electrolyte is a soluble salt of the metal that is being purified.

In purification of copper, the anode is the impure copper, the cathode is pure copper and the electrolyte is copper (II) sulphate solution.



An electrolytic cell before purification of copper

At the Anode
The anode dissolves during electrolysis to produce copper ions (Cu^{2+}). Copper ions enter the solution. The reaction at the anode is shown below.

$$Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^{-}$$

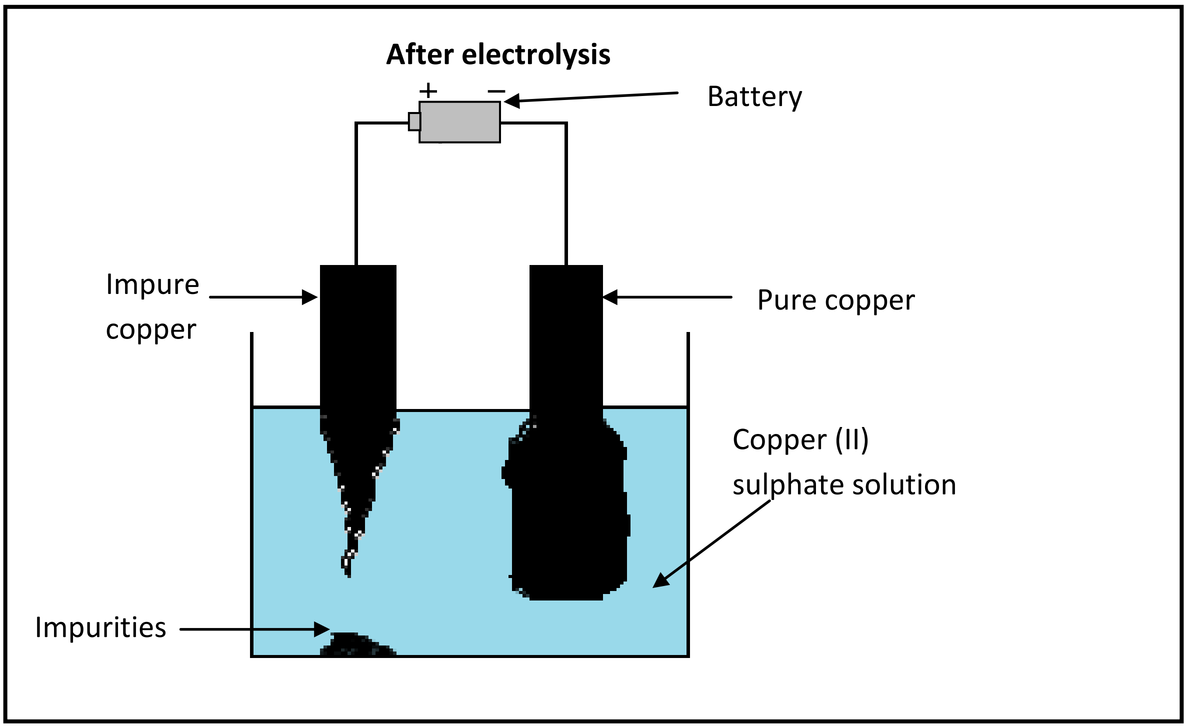
The size of the anode decreases during electrolysis.

At the Cathode

Copper ions in the solution move to the cathode and gain two electrons. Copper ions are converted to copper atoms and are deposited at the cathode. The reaction at the cathode is shown below.

$$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

The size of the cathode increases while the anode decreases as shown below. Copper atoms are converted to copper ions at the anode at the same rate copper ions are converted to copper atoms at the cathode. For this reason the concentration of the electrolyte does not change.



An electrolytic cell after purification of copper

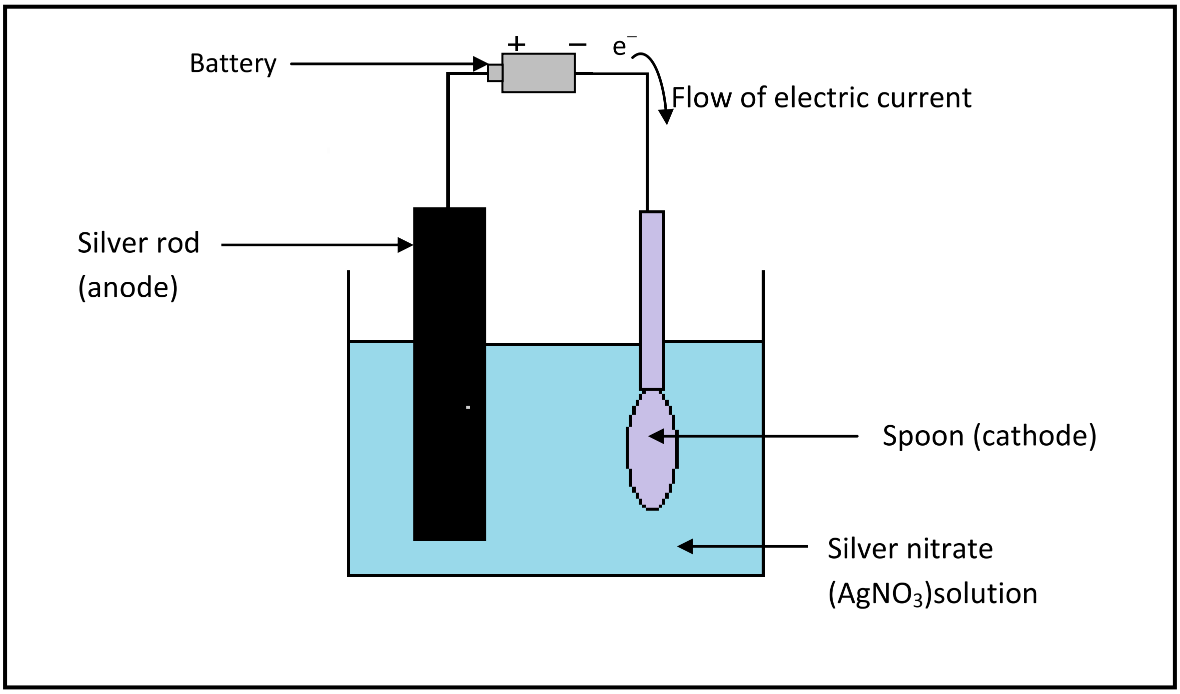
The copper from the anode has been transferred to the cathode. The impurities fall to the bottom of the cell and collected below the anode. The impurities are known as **anode slime** or **anode sludge** and may contain other precious metals such as gold and silver.

12.3.8 Electroplating

Electroplating is a process involving electrolysis to coat metal with other metal. The purpose of electroplating is to prevent the object from **corrosion** and at the same time making it more attractive.

Eating utensil like spoon and fork are often shiny because they are coated with metal called **chromium**.

If you want to electroplate a metal, the metal becomes the cathode. The metal used in electroplating is the anode and the electrolyte is a soluble salt of the metal used in electroplating. The spoon in the diagram below is to be electroplated with silver. **Spoon** becomes the **cathode** and the **silver rod** becomes the **anode**. Silver nitrate solution is used because it is a soluble salt of silver.



Electroplating a spoon

At the anode

The silver rod dissolves to produce silver ions. These ions enter the solution.

The half equation at the anode is shown below:

$Ag_{(s)} \longrightarrow Ag^{+}_{(aq)} + e^{-}$

Silver atom loses one electron to become silver ion. This is oxidation reaction.

At the cathode

Silver ions will move to the cathode and gain an electron to form silver atoms. The silver atoms are deposited on the spoon, thus, coating the spoon.

The half equation is shown below:

$Ag^{+}_{(aq)} + e^{-} \longrightarrow Ag_{(s)}$

Silver atoms ions lose one electron to form silver atom. This is a reduction reaction.
Some metals used in electroplating and commonly plated objects.

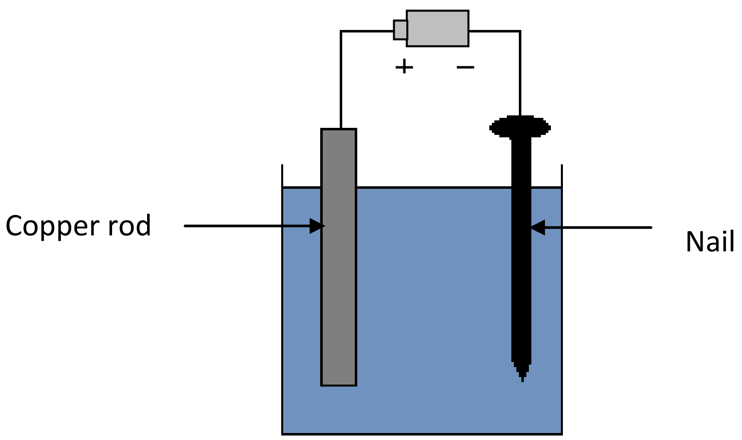
Electroplating Metal	Plated Objects
Chromium	Water taps, motorcar bumpers and bicycle parts
Tin	Tin cans for storing foodstuff
Silver	Sports trophies, plates, cutlery and ornaments
Gold	Watches, cutlery, water taps and ornaments
Rhodium	Silver ware, jewellery, watches and ornaments

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 4 **30 minutes**

Answer the following questions:

1. An iron nail is electroplated with copper using the electrolytic cell shown below.



(a) Write the reaction that will happen at the:

(i) copper

____

(ii) nail

____

(b) Name an electrolyte for the process.

____

(c) Which electrode will decrease in size during electrolysis?
____

(d) What will happen to the concentration of the electrolyte during electrolysis?
____

3. A piece of zinc is purified using electrolysis.

(a) Draw and label a diagram to show the purification of zinc.

(b) Write the reaction at
i) Cathode____
ii) Anode ____

(c) Name a suitable electrolyte that can be used.
____

Thank you for completing your learning activity 4. Check your work. Answers are at the end of this module.

12.3.9 Electrochemical Cell

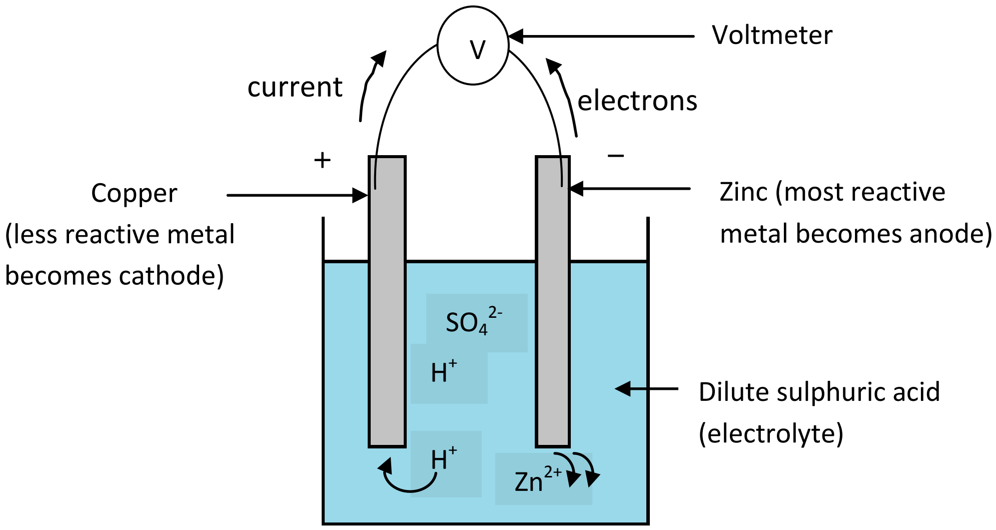
The words **voltaic cell** or **galvanic cell** is often used to refer to the **electrochemical cell**. This idea on electrochemical cell was discovered by an Italian scientist called **Volta**. He found out that there is a difference in the energy of the charges (positive and negative charges) between two different metals when they are placed in a certain electrolyte. This results in electrical current to be produced.

Parts of electrochemical cell

An electrochemical cell is made up of two electrodes from different metals, an electrolyte and a salt bridge. The most reactive metal will lose electrons and so become negatively charged. The less reactive metal will attract positive ions present in the electrolyte and so these positive ions are deposited on the metal and make it positively charged. In doing so, it establish a flow of electrons from most reactive metal to less reactive metal, producing electriccurrent.

Simple electrochemical cell

A simple electrochemical cell is made by dipping a strip of zinc metal and one copper metal strip inside a dilute sulphuric acid solution.



Solutions (electrolyte) containing ions.

As shown in the diagram above, if these two metals are connected by a wire to the solution, electric current will start flowing from copper metal to the zinc metal through the wire. Zinc is reactive, will lose electrons and becomes negative electrode called **anode** while copper is less reactive will become positively charged electrode called **cathode**.

The reactions that will happen at the electrodes are:

At the anode (half equation): $Zn_{(s)} \longrightarrow Zn^{2+} + 2e^-_{(aq)}$

At the cathode (half equation): $Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$

Overall equation: $Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(s)} + Cu_{(s)}$

Electrons are lost at the anode, so **oxidation** reaction is happening. While **reduction** reaction is experienced at the cathode because electrons are gain there. The dry cells and car batteries are practical examples of electrochemical cells.

In galvanic cell (electrochemical cell), oxidation reaction happens at the anode while reduction reaction at the cathode. The chemical reactions that happen in both electrodes produce electric current in a galvanic cell.

Cell Diagram

It is tiring and time consuming to draw diagrams to show electrochemical cells. So scientists have developed a shorthand way of representing electrochemical cells called **cell diagrams**.

Example of cell diagram.

$Zn_{(s)} / Zn^{2+}{}_{(aq)} // Ag^{2+}{}_{(aq)} / Ag_{(s)}$

Anode Salt bridge Cathode

The cell diagram above can be interpreted like this; a piece of zinc metal is dipped into an electrolyte containing zinc ions connected by a salt bridge to another electrolyte which has silver ions which holds silver metal dipped into it.

The Cell Potential

The batteries in the car and in your torch are examples of how chemical reactions create power (current) through flow of electrons.

The cell potential is the measure of how much voltage (electric force) exists between the two half cells of a battery. The equation is given below.

$$\mathbf{E^0_{(cell)} = E^0_{Cathode} - E^0_{Anode}}$$

Example:

Calculate the standard cell potential, E^0 of $Zn_{(s)} / Zn^{2+}{}_{(aq)} // Ag^{2+}{}_{(aq)} / Ag_{(s)}$ given $E^0 = Zn/Zn^{2+} = -0.76V$ and $E^0 = Ag/Ag^{+} = +0.34V$

Solution: $E^0_{(cell)} = E^0_{Cathode} - E^0_{Anode} = +0.34V - -0.76V = +0.34V + 0.76V = 1.10V$

> Normally, to work out the standard cell potential, a table of values of cell potentials (E^0) will be provided with their respective half equations.

- In an electrochemical cell, it is the chemical reactions that take place at the electrodes that produce electric current.
- Galvanic cell and voltaic cell are two words that are also used to mean electrochemical cell.
- In an electrochemical cell, the anode is the negative electrode and cathode is the positive electrode. The anode is always a reactive metal while the cathode is the less reactive metal.

Cell potential (E^0) is the measure of how much voltage exists between the two half cells of a battery. It is calculated by cell potential of the cathode minus cell potential of the anode.

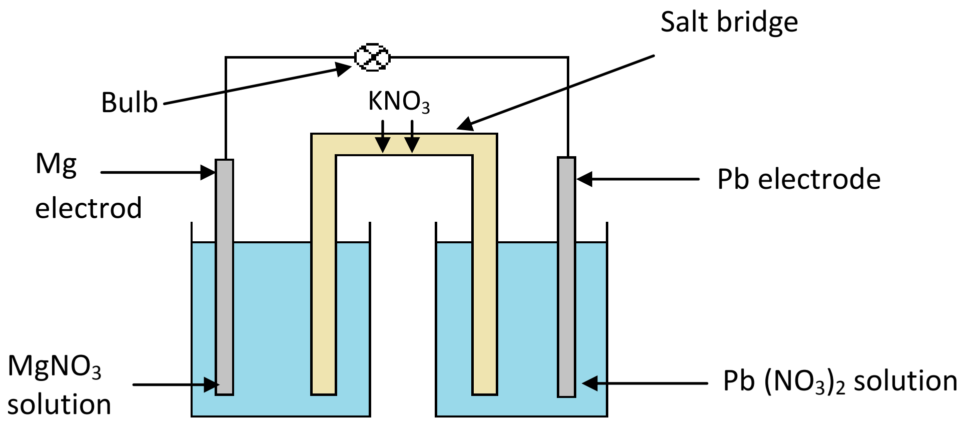
- Cell diagrams are shorthand way of representing electrochemical cells.
- In an electrolytic cell, electrical current causes chemical reactions to occur while in electrochemical cell, it is the chemical reactions that create electric current.
- Dry cells and car batteries are practical examples of electrochemical cells.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 5 **30 minutes**

Answer the following questions:

1. Consider the following electrochemical cell.



a) Write the cell diagram of the above electrochemical cell.

____

b) Which electrode is the anode?

____

c) Which electrode is the cathode?

____

d) Write the equation for the reaction occurring at the surface of the Mg electrode.

____

e) Write out the equation for the reaction occurring at the Pb electrode.

____

2. A galvanic cell consists of a Mg electrode in a $MgNO_3$ solution and a Ag electrode in $AgNO_3$ solution.

   Calculate the standard cell potential (E^0) for this electrochemical cell at 25^0C?

   $Mg^{2+} + 2e^- \longrightarrow Mg$ $E^0 = -2.37V$
   $Ag^+ + e^- \longrightarrow Ag$ $E^0 = 0.80V$

Thank you for completing your learning activity 5. Check your work. Answers are at the end of this module.

REVISE WELL USING THE MAIN POINTS ON THE NEXT PAGE.

SUMMARY

You will now revise this module before doing Assessment 6. Here are the main points to help you revise. Refer to the module topic if you need more information.

- Electrolysis is the splitting or breaking down of molten compounds or aqueous solution using electric current .The current is conducted by the movement of ions in the liquid and chemicals are deposited at the points where the current enters or leaves the liquid.
- Electrode is a piece of metal or carbon placed in an electrolyte through which electric current enters or leaves during electrolysis.
- Anode is the positive electrode. And cathode is the negative electrode.
- Oxidation reaction occurs at the anode while reduction reaction at the cathode.
- Inert electrodes are electrodes which do not react with the products or electrolyte during electrolysis. Platinum and graphite (carbon) are commonly used inert electrodes.
- Electrolytes are liquid which contain ions. The ions are free to move around. When two electrodes which have been connected in a circuit with a battery are put into the liquid electrolyte, the charged ions are attracted to the charged electrodes.
- Electrolysis of molten compounds involves two ions only. One is a positive ion called a cation and one a negative ion called anion.
- Cations will attract to cathode and gain electrons and become neutral atoms. Anions will also attract to anode and lose electrons to become neutral atoms.
- An aqueous solution is formed when an ionic compound dissolves in water. The ions in the solution are free to move around in the solution.
- There are four ions present during the electrolysis of an aqueous solution. Two ions, one cation and anion, are from the compound itself and other two ions, cation and anion are from water.
- Half-equations show the kind of reactions that happen at the electrodes.
- Overall equation is a combination of equations from the electrodes based on a balanced number of electrons.
- The performance of an electrolytic cell is affected by reactivity of elements, concentration of solutions and the nature of electrode used.
- The Law on Selective Discharge of Ions is based on the Reactivity Series. It determines which ions will be attracted to the respective electrodes and become discharged. Ions of elements at the bottom of the Reactivity Series are easily discharged.
- Concentration of the electrolytes also affects the type of ions to be discharged at the electrodes. If aqueous solution is diluted or not concentrated, electrolysis at the anode will be in favour of the anion of water. However, if aqueous solution is concentrated, electrolysis at the anode will be in favour of the anion of the concentrated electrolyte.

- Pure water contains few ions but when mixed with acid, it can conduct electricity. During electrolysis of water using a Hofmann Voltameter, hydrogen gas is formed at the cathode while oxygen gas and water are produced at the anode. Brine is a solution of sodium chloride. The electrolysis of brine which contains concentrated sodium chloride solution leads to the production of chlorine gas. Sodium hydroxide and hydrogen gas are also other products formed during the course of electrolysis of brine.
- Bauxite is the main ore for aluminium. It is aluminium oxide mixed with other impurities like sand and iron ore.
- Impurities are removed from bauxite to form alumina. Alumina is dissolved in molten cryolite. When electric current is passed through it, it causes the release of aluminium ions that move towards the cathode and gain electrons to become an aluminium atom.
- Inert electrodes are used for both the electrolysis of brine and aluminium from its ore.
- Electrolytic refining of a metal and electroplating of an object are two practical applications of electrolysis. Reactive electrodes are used instead of inert electrodes.
- In electrolytic refining of metal, impure metal is to be the anode and the cathode contains the pure metal. The electrolyte is a soluble salt of the metal that is being purified.
- Electroplating is the coating of an object with a metal. The setup of electroplating is as follows:
  - the object to be coated is made the cathode.
  - the metal that will be used to do the coating is made the anode.
  - the electrolyte is a solution containing the ions of the metal used that will do the coating.
- During the course of electrolysis in electrolytic refining and electroplating, the cathode will increase in size while the anode will decrease in its size.

NOW YOU MUST COMPLETE ASSESSMENT 5 AND RETURN IT TO THE PROVINCIAL CENTRE CO-ORDINATOR.

ANSWERS TO LEARNING ACTIVITIES 1-5

Learning Activity 1

1. (a) (i) Electrolysis is the splitting or breaking down of molten compounds or aqueous solution using electric current .

   (ii) Anode is the positive electrode. It is always connected to the positive terminal of the power source.

   (iii) Cathode is the negative electrode. It is always connected to the negative terminal of the power source.

   (b) (i) cathode
   (ii) anion
   (iii) electrolyte
   (iv) cation
   (v) anode

2. (a) cathode
   (b) anode

3. (a) gain electrons
   (b) lose electrons

4. Inert electrodes do not react with electrolyte or other products during electrolysis while reactive electrodes do.

5. (a) Cu^{2+}, Cl^{-},H^{+} and OH^{-} (copper ion, chloride ion, hydrogen ion and hydroxide ion).
   (b) Cu^{2+} and H^{+} (copper ion and hydrogen ion).
   (c) H^{+} (hydrogen ion).Because hydrogen is less reactive than copper, therefore, its ions are easily discharged.
   (d) $2H^{+}{}_{(aq)} + 2e^{-} \longrightarrow H_{2\,(g)}$
   (e) $4OH^{-}{}_{(aq)} \longrightarrow O_{2\,(g)} + 2H_2O_{\,(l)} + 4e^{-}$

Learning Activity 2

1. (a) Magnesium ions (Mg^{2+}), chloride ions (Cl^{-}).
   (b) Magnesium ions (Mg^{2+}).
   (c) Reduction reaction.

(d) (i) $Mg^{2+}_{(l)} + 2e^- \longrightarrow Mg_{(l)}$
(ii) $2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^-$

(e) $Mg^{2+}_{(l)} + 2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + Mg_{(l)}$
Alternative answer: $2Mg\,Cl_{2\,(l)} \longrightarrow Cl_{2\,(g)} + 2Mg_{(l)}$

1. (a) Na^+,SO_4^{2-},H^+ and OH^- (sodium ion, sulphate ion, hydrogen ion and hydroxide ion).
(i) $2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$
(ii) $4OH^-_{(aq} \longrightarrow O_{2\,(g)} + 2H_2O_{(l)} + 4e^-$
(iii) $2H_2O_{(l)} \longrightarrow H_{2\,(g)} + O_{2\,(g)}$

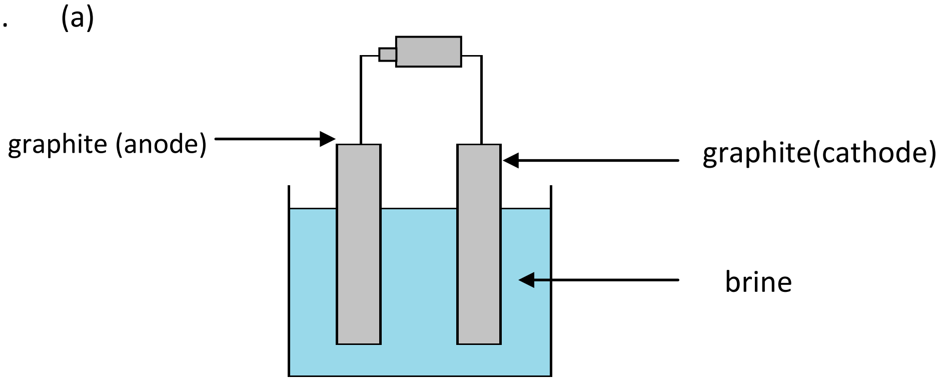
3. (a) potassium ion, chloride ion, hydrogen ion and hydroxide ion (K^+,Cl^-,H^+ and OH^-).
(b) (i) Hydrogen ion (H^+).Because hydrogen is less reactive than potassium. Therefore, its ion is easily discharged.
(ii) Chloride ion (Cl^-).Because the solution is concentrated.

(c) (i) $2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$
(ii) $2Cl^-_{(aq)} \longrightarrow Cl_{2\,(g)} + 2e^-$
(iii) $2H^+_{(l)} + 2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + H_{2\,(g)}$
Alternative answer: $2HCl_{2\,(aq)} \longrightarrow Cl_{2\,(g)} + H_{2(g)}$

Learning Activity 3

1. (i) Brine is a concentrated solution of sodium chloride.
(ii) Bauxite is the aluminium ore.

2. (a)

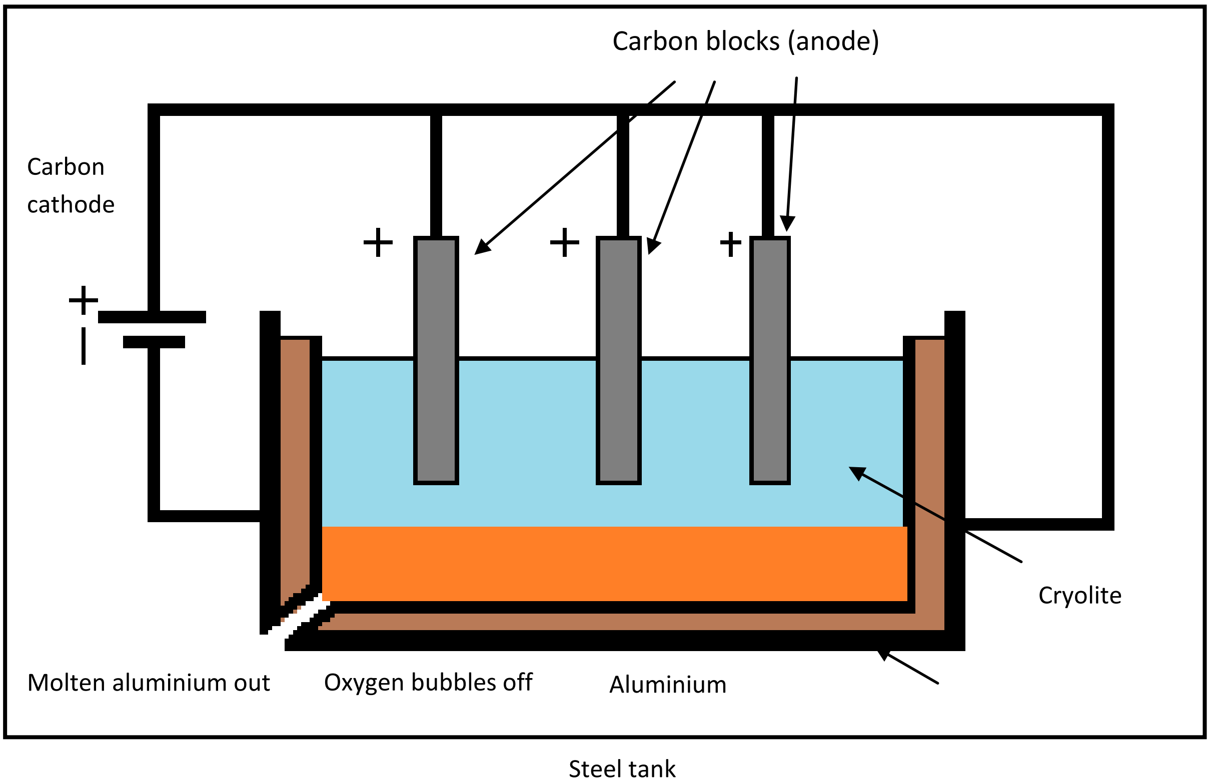


(i) $2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$

(ii) $2Cl^-_{(aq)} \longrightarrow Cl_{2\,(g)} + 2e^-$

(iii) $2HCl_{(aq)} \longrightarrow H_{2\,(g)} + Cl_{2\,(g)}$

3. (a)



(b) To gain electrons.

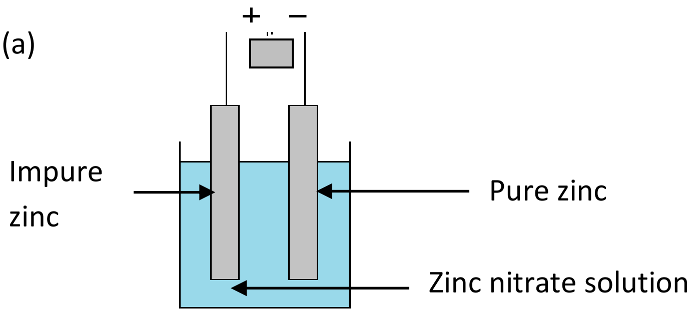
(c) Al^{3+}(aluminium ion) receives 3 electrons and become aluminium atom (Al).
Half equation: $Al^{3+}_{(aq)} + 3e^{-} \longrightarrow Al_{(s)}$

(d) Anode is used up because it is continuously reacting with oxygen to form carbon dioxide.

Learning Activity 4

1. (a) i) $Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^{-}$
   ii) $Cu^{2+}_{(aq)} + 2e^{-} \longrightarrow Cu_{(s)}$
   (b) Copper sulphate or copper nitrate.
   (c) Anode (copper rod).
   (d) Concentration of electrolyte will remain the same.

2. (a)



(b) i) $Zn^{2+}{}_{(aq)} + 2e^{-} \longrightarrow Zn_{(s)}$

ii) $Zn_{(s)} \longrightarrow Zn^{2+}{}_{(aq)} + 2e^{-}$

(c) Zinc nitrate

Learning Activity 5

1. (a) $Mg_{(s)} / Mg^{2+}{}_{(aq)} // Pb^{2+}{}_{(aq)} / Pb_{(s)}$
   (b) Mg electrode (magnesium electrode).
   (c) Pb electrode (lead electrode).
   (d) $Mg_{(s)} \longrightarrow Mg^{2+}{}_{(aq)} + 2e^{-}$
   (e) $Pb^{2+}{}_{(aq)} + 2e^{-} \longrightarrow Pb_{(s)}$
2. $E^0{}_{(cell)} = E^0{}_{Cathode} - E^0{}_{Anode} = +0.80V - -2.37V = 3.17V$

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3	CENTRAL	C/- FODE HQ	3419228	72228110	72229050
4	DARU	P. O. Box 68, Daru	6459033	72228146	72229047
5	GOROKA	P. O. Box 990, Goroka	5322085/5322321	72228116	72229054
6	HELA	P. O. Box 63, Tari	73197115	72228141	72229083
7	JIWAKA	c/- FODE Hagen		72228143	72229085
8	KAVIENG	P. O. Box 284, Kavieng	9842183	72228136	72229069
9	KEREMA	P. O. Box 86, Kerema	6481303	72228124	72229049
10	KIMBE	P. O. Box 328, Kimbe	9835110	72228150	72229065
11	KUNDIAWA	P. O. Box 95, Kundiawa	5351612	72228144	72229056
12	LAE	P. O. Box 4969, Lae	4725508/4721162	72228132	72229064
13	MADANG	P. O. Box 2071, Madang	4222418	72228126	72229063
14	MANUS	P. O. Box 41, Lorengau	9709251	72228128	72229080
15	MENDI	P. O. Box 237, Mendi	5491264/72895095	72228142	72229053
16	MT HAGEN	P. O. Box 418, Mt. Hagen	5421194/5423332	72228148	72229057
17	NCD	C/- FODE HQ	3230299 ext 26	72228134	72229081
18	POPONDETTA	P. O. Box 71, Popondetta	6297160/6297678	72228138	72229052
19	RABAUL	P. O. Box 83, Kokopo	9400314	72228118	72229067
20	VANIMO	P. O. Box 38, Vanimo	4571175/4571438	72228140	72229060
21	WABAG	P. O. Box 259, Wabag	5471114	72228120	72229082
22	WEWAK	P. O. Box 583, Wewak	4562231/4561114	72228122	72229062

FODE SUBJECTS AND COURSE PROGRAMMES

GRADE LEVELS	SUBJECTS/COURSES
Grades 7 and 8	1. English
	2. Mathematics
	3. Personal Development
	4. Social Science
	5. Science
	6. Making a Living
Grades 9 and 10	1. English
	2. Mathematics
	3. Personal Development
	4. Science
	5. Social Science
	6. Business Studies
	7. Design and Technology- Computing
Grades 11 and 12	1. English – Applied English/Language& Literature
	2. Mathematics – General / Advance
	3. Science – Biology/Chemistry/Physics
	4. Social Science – History/Geography/Economics
	5. Personal Development
	6. Business Studies
	7. Information & Communication Technology

REMEMBER:

- For Grades 7 and 8, you are required to do all six (6) subjects.
- For Grades 9 and 10, you must complete five (5) subjects and one (1) optional to be certified. Business Studies and Design & Technology – Computing are optional.
- For Grades 11 and 12, you are required to complete seven (7) out of thirteen (13) subjects to be certified.

Your Provincial Coordinator or Supervisor will give you more information regarding each subject and course.

Notes: You must seek advice from your Provincial Coordinator regarding the recommended courses in each stream. Options should be discussed carefully before choosing the stream when enrolling into Grade 11. FODE will certify for the successful completion of seven subjects in Grade 12.

GRADES 11 & 12 COURSE PROGRAMMES			
No	**Science**	**Humanities**	**Business**
1	Applied English	Language & Literature	Language & Literature/Applied English
2	General / Advance Mathematics	General / Advance Mathematics	General / Advance Mathematics
3	Personal Development	Personal Development	Personal Development
4	Biology	Biology/Physics/Chemistry	Biology/Physics/Chemistry
5	Chemistry/ Physics	Geography	Economics/Geography/History
6	Geography/History/Economics	History / Economics	Business Studies
7	ICT	ICT	ICT

CERTIFICATE IN MATRICULATION STUDIES		
No	**Compulsory Courses**	**Optional Courses**
1	English 1	**Science Stream:** Biology, Chemistry and Physics
2	English 2	**Social Science Stream:** Geography, Intro to Economics and Asia and the Modern World
3	Mathematics 1	
4	Mathematics 2	
5	History of Science & Technology	

REMEMBER:

You must successfully complete 8 courses: 5 compulsory and 3 optional.