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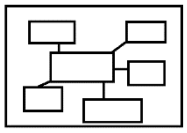
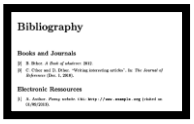
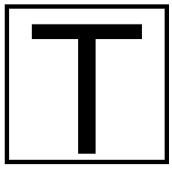
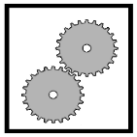
TECHNICAL SCIENCES

GRADE 12

SPRING CLASSES 2026

TEACHER AND LEARNER CONTENT MANUAL

ICON DESCRIPTION

 MIND MAP	 EXAMINATION GUIDELINE	 CONTENTS	 ACTIVITIES
 BIBLIOGRAPHY	 TERMINOLOGY	 WORKED EXAMPLES	 STEPS

CONTENTS **PAGE**

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Examination Guidelines



Chemical Change (Grade 12)
Electrochemical cells
(This section must be read in conjunction with the CAPS, p. 55.)

Electrolytic cells and galvanic cells
- Define the electrolytic cell as an electrochemical cell that converts electrical energy to chemical energy. It is a non-spontaneous.
- Define the galvanic (voltaic) cell as an electrochemical cell that converts chemical energy to electrical energy. It is spontaneous.
- Define an electrolyte as a solution/liquid/dissolved substance that conducts electricity through the movement of ions.

Understanding the processes and redox reactions taking place in electrolytic cells
- Describe the movement of ions in the solution.
- State the direction of electron flow in the external circuit.
- Write equations for the half-reactions taking place at the anode and cathode.
- Write down the overall cell reaction by combining two half-reactions.
- Describe, using half-reactions and the equation for the overall cell reaction as well as the layout of the particular cell using a schematic diagram, the following electrolytic process:
  - The decomposition of copper (II) chloride
  - Electroplating a metal by another metal

Understanding the processes and redox reactions taking place in galvanic cells
- Describe the movement of ions in the solutions and through the salt bridge.
- State the direction of electron flow in the external circuit.
- Write down the half-reactions that occur at the electrodes.
- State the function of the salt bridge.
- Name the components of a galvanic cell.
- State the functions of all components of a galvanic cell.
- Use cell notation or diagrams to represent a galvanic cell.
- When writing cell notation, the following convention should be used:
  - Cell terminals (electrodes) are written on the outside of the cell notation.
  - Active electrodes： reducing agent | oxidised species || oxidising agent | reduced species
- Predict the half-cell in which oxidation will take place when two half-cells are connected.
- Predict the half-cell in which reduction will take place when connected to another half-cell.
- Write down the overall cell reaction by combining two half-reactions.
- Use the Table of Standard Reduction Potentials to calculate the emf of a standard galvanic cell.
- Use a positive value of the standard emf as an indication that the reaction is spontaneous under standard conditions.
- Determine the electrode potential of a galvanic cell, e.g. Cu-Zn cell or any other galvanic cell.

Standard conditions
- Write down the standard conditions under which standard electrode potentials are determined.

Alternate Energies
Discuss the following alternate energies:
- Biodiesel
- Fuel cell
- Photovoltaic cell
- Discuss the use of alternate energies and their environmental impact.

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Terms and Definitions

TERM	DEFINITION
Galvanic cell	A cell in which chemical energy is converted into electrical energy. A galvanic (voltaic) cell has self-sustaining electrode reactions.
Electrolytic cell	A cell in which electrical energy is converted into chemical energy.
Redox reaction	A reaction in which an electron transfer takes place.
Oxidation	A loss of electrons. / An increase in oxidation number.
Reduction	A gain of electrons. / A decrease in oxidation number.
Oxidising agent	A substance that is reduced / gains electrons / whose oxidation number decreases.
Reducing agent	A substance that is oxidised / loses electrons / whose oxidation number increases.
Anode	The electrode where oxidation takes place.
Cathode	The electrode where reduction takes place.
Electrolyte	A solution that conducts electricity through the movement of ions.
Electrolysis	The chemical process in which electrical energy is converted to chemical energy OR the use of electrical energy to produce a chemical change.
Salt bridge	The connection between two half-cells needed to ensure electrical neutrality in the cell. OR: A component used in a galvanic cell to complete the circuit.
Electrodes	An electrical conductor used in a galvanic cell to make contact with a non-metallic part of the circuit, e.g. the electrolyte.

ELECTROLYTIC CELLS AND GALVANIC CELLS

1. KEY DEFINITIONS

Electrolytic cell: An electrochemical cell that converts electrical energy to chemical energy. It is non-spontaneous (requires an external power source).

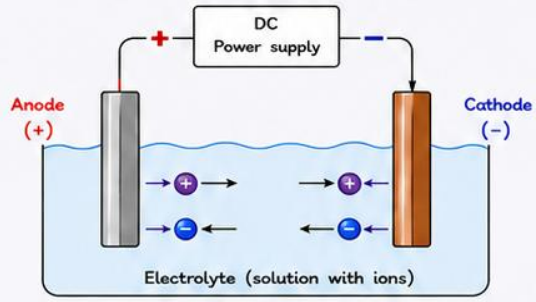
Galvanic (voltaic) cell: An electrochemical cell that converts chemical energy to electrical energy. It is spontaneous (no external power source needed).

Electrolyte: A solution / liquid / dissolved substance that conducts electricity through the movement of ions.

2. AT A GLANCE

FEATURE	ELECTROLYTIC CELL	GALVANIC (VOLTAIC) CELL
Energy conversion	Electrical energy → Chemical energy	Chemical energy → Electrical energy
Spontaneity	Non-spontaneous	Spontaneous
External power source	Requires external power source	Produces electrical energy (no external power source)
EMF (Electromotive Force)	Negative ($E < 0$)	Positive ($E > 0$)
Electrodes	Anode is positive (+) Cathode is negative (−)	Anode is negative (−) Cathode is positive (+)
Main use	Electrolysis, electroplating, extraction of metals, charging batteries	Batteries, fuel cells, generators, providing electricity

3. ELECTROLYTIC CELL (Non-spontaneous)



- Cations move to the cathode (negative electrode).
- Anions move to the anode (positive electrode).

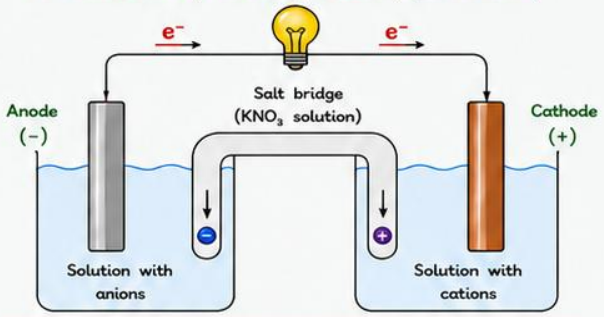
Key points

- External power source forces a non-spontaneous reaction.
- Electrons flow from the positive (+) terminal to the negative (−) terminal.
- Chemical changes occur at the electrodes.

Examples

- Electrolysis of molten NaCl (to produce Na and Cl_2)
- Electroplating of metals
- Charging a rechargeable battery

4. GALVANIC (VOLTAIC) CELL (Spontaneous)



- Anions move through the salt bridge to the anode.
- Cations move through the salt bridge to the cathode.

Key points

- A spontaneous redox reaction occurs.
- Electrons flow from the anode (−) to the cathode (+) through the external circuit.
- The salt bridge keeps the charge balanced and completes the circuit.

Examples

- Zn–Cu Cell (Zn | Zn^{2+} || Cu^{2+} | Cu)
- Dry cells, car batteries, fuel cells

5. SUMMARY COMPARISON

Aspect	Electrolytic cell	Galvanic cell
Energy change	Electrical → Chemical	Chemical → Electrical
Spontaneity	Non-spontaneous	Spontaneous
Power source	Requires external power	No external power (generates electricity)
EMF (E)	Negative ($E < 0$)	Positive ($E > 0$)
Anode	Positive (+)	Negative (−)
Electron flow	From + to − (from anode to cathode)	From anode (−) to cathode (+)
Main use	Electrolysis, electroplating, extraction, charging	Batteries, generators, fuel cells

6. REMEMBER!

- Anode = Oxidation
- Cathode = Reduction
  "An Ox, Red Cat" helps you remember.
- The movement of ions, electrons, and the role of the salt bridge are key to the working of these cells.

Electrolytic cell:
Needs electricity to make chemical change.

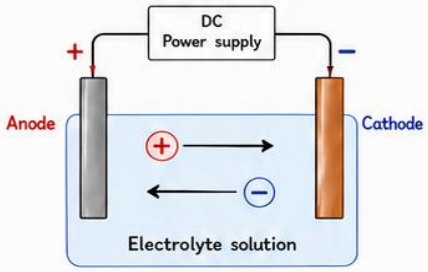
Galvanic cell:
Chemical change makes

ELECTROLYTIC CELLS

An electrolytic cell uses electrical energy to drive a **non–spontaneous** redox reaction.

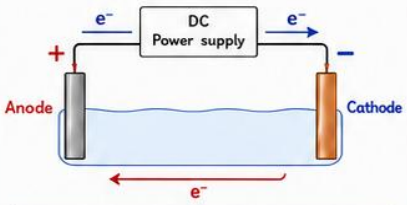
1. MOVEMENT OF IONS IN THE SOLUTION

- Anions (−) move towards the ANODE (+) because it is positively charged.
- Cations (+) move towards the CATHODE (−) because it is negatively charged.



2. DIRECTION OF ELECTRON FLOW

- Electrons flow in the EXTERNAL CIRCUIT from the NEGATIVE (−) terminal of the power supply to the CATHODE (−) and from the ANODE (+) back to the POSITIVE (+) terminal.



Inside the solution, charge is carried by the movement of ions, not electrons.

3. HALF-REACTIONS

CATHODE (−): Reduction

Cations gain electrons.

$M^{n+} + ne^- \rightarrow M_{(s)}$

ANODE (+): Oxidation

Anions lose electrons.

$X^{n-} \rightarrow X + ne^-$

REMEMBER:
An Ox, Red Cat
Anode = Oxidation
Cathode = Reduction

4. OVERALL CELL REACTION

- Write the half-reactions.
- Make sure the number of electrons is the same.
- Add the half-reactions and cancel out electrons.

Example:

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

Reduction: $Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$

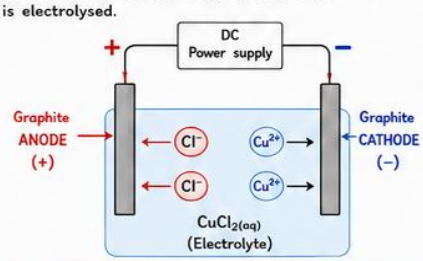
Overall reaction:

$Cu^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow Cu_{(s)} + Cl_{2(g)}$

5. ELECTROLYTIC PROCESSES

a) DECOMPOSITION OF COPPER(II) CHLORIDE ($CuCl_2$)

The molten or aqueous copper(II) chloride is electrolysed.



ION MOVEMENT

- Cl^- (anions) move to the ANODE (+).
- Cu^{2+} (cations) move to the CATHODE (−).

ANODE (+): Oxidation

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

Chloride ions lose electrons and chlorine gas is produced.

CATHODE (−): Reduction

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

Copper(II) ions gain electrons and copper metal is deposited.

OVERALL CELL REACTION

$Cu^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow Cu_{(s)} + Cl_{2(g)}$

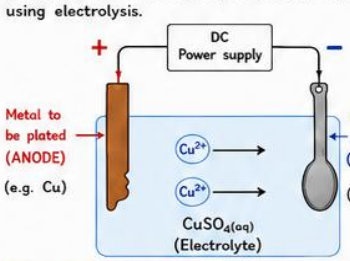
Copper(II) chloride is decomposed into copper metal and chlorine gas.

OBSERVATIONS

- Bubbles of greenish-yellow chlorine gas at the anode (+).
- Reddish-brown copper metal deposited at the cathode (−).
- Colour of the solution may fade as Cu^{2+} ions are removed.

b) ELECTROPLATING A METAL BY ANOTHER METAL

A thin layer of one metal is deposited on another using electrolysis.



ION MOVEMENT

- Cu^{2+} (cations) move to the CATHODE (−).
- (Anions move to the ANODE (+).

ANODE (+): Oxidation

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

Copper atoms lose electrons and enter solution as Cu^{2+} ions.

CATHODE (−): Reduction

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

Copper(II) ions gain electrons and copper is deposited on the object.

OVERALL CELL REACTION

$Cu_{(s)}$ (anode) $\rightarrow Cu_{(s)}$ (cathode)

Copper is transferred from the anode to the cathode.

OBSERVATIONS

- The object becomes coated with a thin layer of copper and gain mass
- The anode loss mass.

6. SUMMARY – ELECTROLYTIC CELLS

- Electrical energy from a DC power supply drives a non-spontaneous reaction.
- Anions (−) move to the ANODE (+); cations (+) move to the CATHODE (−).
- Electrons flow in the external circuit from (−) to the cathode and from the anode back to (+).
- Oxidation occurs at the ANODE (+); reduction occurs at the CATHODE (−).
- Products form at the electrodes depending on the electrolyte used.

WHY DC POWER SOURCE IS USED AND NOT AC POWER SOURCES

1. The current and electrons flow in only one direction.
2. The polarity of the electrodes remains the same.

7. KEY TERMS

Electrolysis:	Decomposition of a compound using electricity.
Electroplating:	Depositing a thin layer of one metal on another using electrolysis.
Oxidation:	Loss of electrons.
Reduction:	Gain of electrons.
Electrodes	Conductors where oxidation or reduction occurs.
Electrolyte:	Substance that contains ions and conducts electricity.

KEY REMEMBER

"An Ox, Red Cat"

Anode = Oxidation Cathode = Reduction

Anions (−) go to ANODE (+)

Cations (+) go to CATHODE (−)

e^- Electrons in external circuit flow from (−) to (+) via the cathode.

Electrolytic cell needs electrical energy.

Worked Example 1

Consider the reaction represented by the following equation:

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

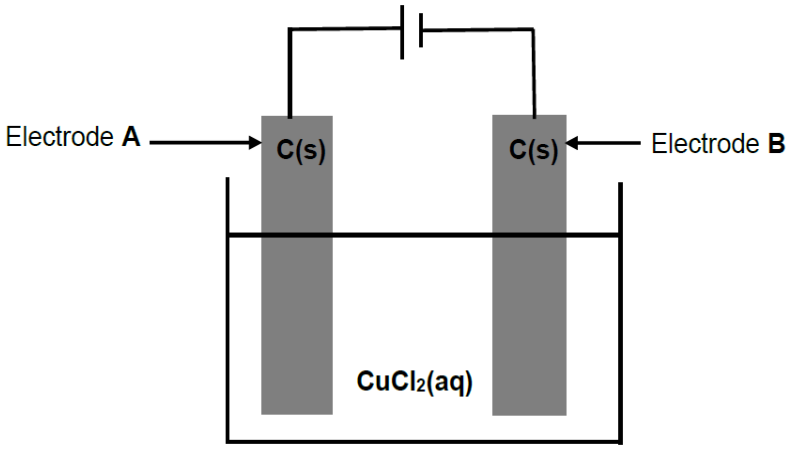
1.1	Identify the substance that is oxidised.	(2)
1.2	Identify the substance that is reduced.	(2)
1.3	Write the oxidation half-reaction.	(3)
1.4	Write the reduction half-reaction.	(3)
1.5	Identify the reducing agent and the oxidising agent.	(4)
		[16]

Solutions

1.1	Cu is oxidized✓✓	(2)
1.2	Ag^{+} ions are reduced✓✓	(2)
1.3	$Cu \rightarrow Cu^{2+} + 2e^{-}$ ✓✓ ✓	(2)
1.4	$2Ag^{+} + 2e^{-} \rightarrow 2Ag$ ✓✓✓	(3)
1.5	Reducing agent: Cu ✓✓ Oxidising agent: Ag^{+}✓✓	(4)
		[16]

Worked Example 2

Study the diagram of an electrochemical cell is given below and then answer the question that follows.



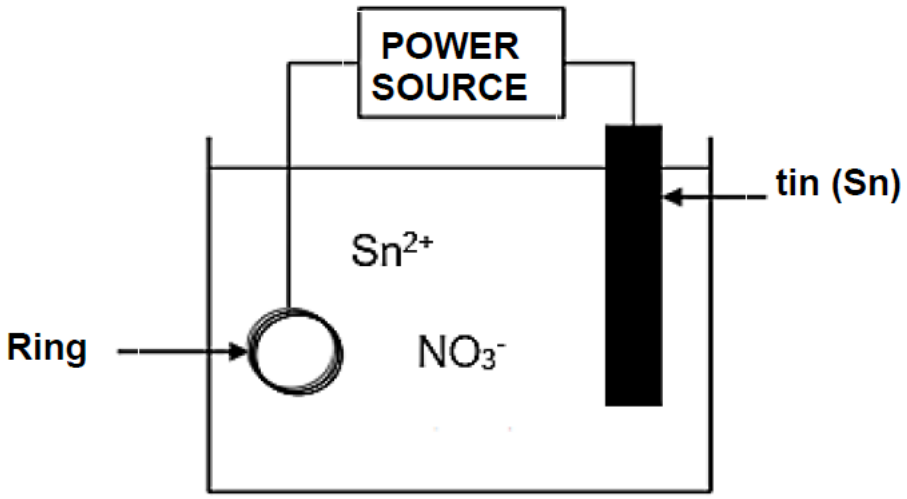
2.1	Identify the type of electrochemical cell shown in the diagram.	(2)
2.2	State the energy conversion that takes place in this cell.	(2)
2.3	Name electrode A and electrode B.	(2)
2.4	Write the half-reaction that takes place at electrode A.	(3)
2.5	Write the half-reaction that takes place at electrode B.	(3)
2.6	State the movement of ions in the electrolyte during electrolysis.	(4)
		[16]

Solutions

2.1	Electrolytic cell ✓✓	(2)
2.2	Electrical energy to Chemical energy ✓✓	(2)
2.3	**Electrode A** is the Anode ✓ **Electrode B** is the Cathode ✓	(2)
2.4	$2Cl^- \rightarrow Cl_2 + 2e^-$ ✓✓✓	(3)
2.5	$Cu^{2+} + 2e^- \rightarrow Cu$ ✓✓✓	(3)
2.6	Cu^{2+} moves to the cathode ✓✓ Cl^- moves to the anode ✓✓	(4)
		[16]

ACTIVITY 1

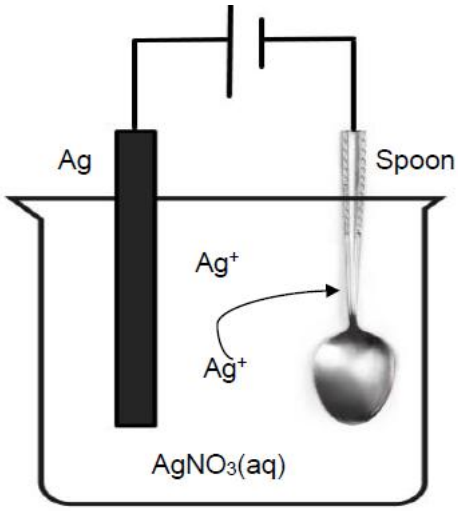
A solution of Tin(II)nitrate, $Sn(NO_3)_2$, is used to plate tin metal onto a ring made of steel. A simplified diagram of such an electroplating cell is shown below.



1.1	Define the term electrolyte.	(2)
1.2	Is the above cell SPONTANEOUS or NON-SPONTANEOUS?	(1)
1.3	Give a reason to support your answer to QUESTION 1.2.	(1)
1.4	Redraw ONLY the power source and indicate the polarity (positive and negative).	(2)
1.5	Identify an anion in $Sn(NO_3)_2$.	(1)
1.6	Write down the half-reaction that takes place at the ring.	(2)
1.7	Give a reason why the concentration of Tin(II)nitrate remains constant.	(1)
		[10]

ACTIVITY 2

The diagram below represents an electrolytic cell that can be used to electroplate an iron spoon with silver.



2.1 Define the term electrolyte. (2)

2.2 Is the reaction taking place in this electrochemical cell SPONTANEOUS or NON-SPONTANEOUS? (1)

2.3 Write down the half-reaction that occurs at the iron spoon (3)

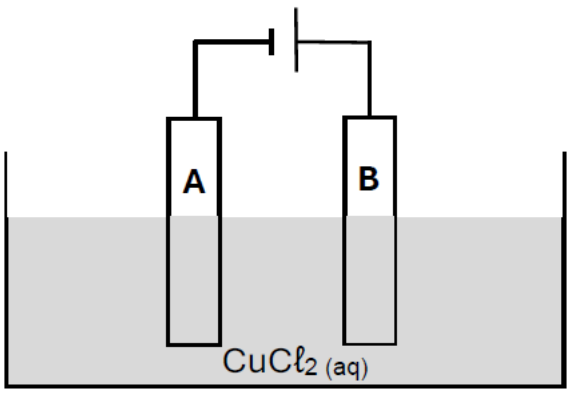
2.4 Write down the NAME or FORMULA of the reducing agent in the cell. Give a reason for your answer. (2)

2.5 Which electrode, the SPOON or SILVER, is the cathode? Give a reason for your answer. (2)

[10]

ACTIVITY 3

The electrochemical cell below makes use of carbon electrodes to decompose copper(II)chloride.



3.1 Define the following terms:

3.1.1 Anode (2)

3.1.2 Electrolysis (2)

3.2 Write down:

3.2.1 A balanced half-reaction that occurs at the anode (2)
3.2.2 NAME of a half-reaction that occurs at the cathode (1)

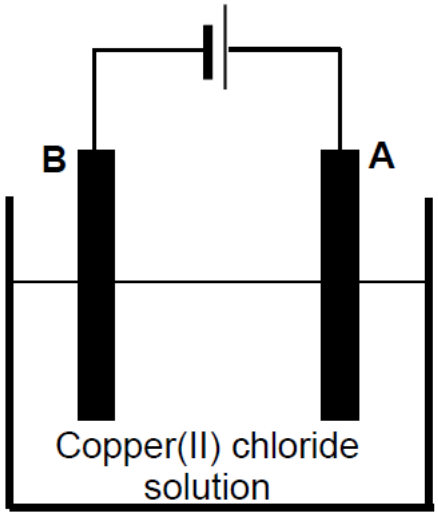
3.3 Write down the NAME or FORMULA of:
3.3.1 An oxidising agent (1)
3.3.2 A reducing agent (1)

3.5 During the reaction in the above cell, bubbles were observed on one of the electrodes. Give the NAME of the chemical substance causing the bubbles. (2)

[10]

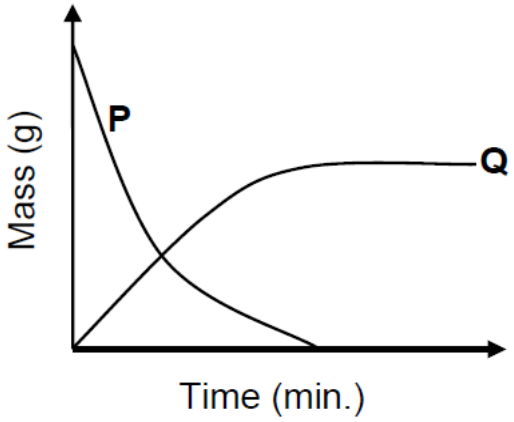
ACTIVITY 4

The electrochemical cell below is set up to demonstrate the purification of copper.



4.1 Write down the type of electrochemical cell illustrated above. (2)

The graphs below show the change in mass of the electrodes whilst the cell is in operation.



4.2 Define a reducing agent in terms of electron transfer. (2)
4.3 Which graph represents the change in mass of electrode A? (1)
4.4 Write down the half-reaction that takes place at electrode A. (2)

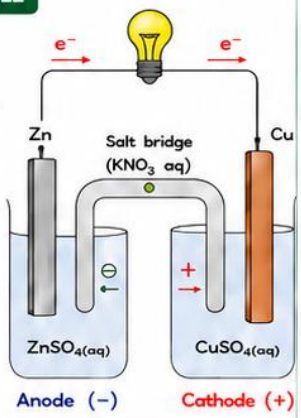
[7]

GALVANIC CELLS (VOLTAIC CELLS)

A galvanic (voltaic) cell converts chemical energy to electrical energy through a spontaneous redox reaction.

1. MOVEMENT IN A GALVANIC CELL

- Ions in the solutions:
  Anions (−) move towards the anode.
  Cations (+) move towards the cathode.
- Ions through the salt bridge:
  Anions (−) move towards the anode.
  Cations (+) move towards the cathode.
  to maintain electrical neutrality.
- Electrons in the external circuit:
  Electrons flow from the anode (−) to the cathode (+).



2. COMPONENTS OF A GALVANIC CELL

COMPONENT	FUNCTION
Anode (−)	Electrode where oxidation takes place. Releases electrons into the external circuit.
Cathode (+)	Electrode where reduction takes place. Electrons from external circuit enter here.
Electrolyte (solutions)	Conducts ions; provides ions for the redox reaction.
Salt bridge	Completes the circuit; allows ions to move and maintains electrical neutrality.
External circuit (wires, bulb, voltmeter)	Provides a path for electron flow; where electrical energy can be used or measured.
Electrodes	Conduct electricity; provide surfaces for redox reactions.

3. HALF-REACTIONS AT THE ELECTRODES

Example: Zn – Cu cell (Zn | Zn^{2+} || Cu^{2+} | Cu)

ANODE (−): Oxidation

Zinc is oxidised.

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

CATHODE (+): Reduction

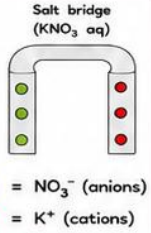
Copper(II) ions are reduced.

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

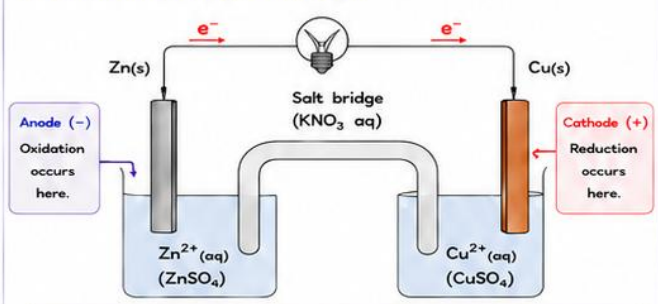
Electrons are released at the anode and flow through the external circuit to the cathode, where they are used.

4. FUNCTION OF THE SALT BRIDGE

- Allows ions to move between the half-cells.
- Maintains electrical neutrality in both solutions.
- Completes the circuit so that electrons can continue to flow.
- Prevents the solutions from mixing.



5. DIAGRAM OF A Zn – Cu CELL



Electrons flow: Zn (anode) → Cu (cathode)

Ions flow: NO_3^- → anode K^+ → cathode

6. PREDICTING OXIDATION AND REDUCTION

- The half-cell with the more negative standard reduction potential (E°_{red}) will be oxidised (anode).
- The half-cell with the more positive standard reduction potential (E°_{red}) will be reduced (cathode).
  (Use the Table of Standard Reduction Potentials.)

7. OVERALL CELL REACTION

- Write the half-reactions.
- Make sure the number of electrons is the same.
- Add the half-reactions and cancel out electrons.

Half-reactions for Zn – Cu cell

Oxidation (anode): $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

Reduction (cathode): $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

Overall cell reaction:

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

8. CELL NOTATION

- Cell terminals (electrodes) are written on the outside.
- Active electrodes: reducing agent | oxidised species || oxidising agent | reduced species

Example: Zn – Cu cell

$Zn_{(s)}$ | $Zn^{2+}_{(aq)}$ || $Cu^{2+}_{(aq)}$ | $Cu_{(s)}$

Anode (−) Cathode (+)

| = phase boundary (between solid and solution)
|| = salt bridge (between two half-cells)

9. CALCULATING THE STANDARD EMF (E°_{cell}) OF A GALVANIC CELL

1. Write the half-reactions and find the standard reduction potential (E°_{red}) for each from the Table.
2. Identify the cathode (higher E°_{red}) and anode (lower E°_{red}).
3. Use the formulas:

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = E^\circ_{reduction} - E^\circ_{oxidation}$

$E^\circ_{cell} = E^\circ_{oxidizing\ agent} - E^\circ_{reducing\ agent}$

Example: Zn – Cu cell

$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$ $E^\circ_{red} = -0{,}76$ V (anode)

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $E^\circ_{red} = +0{,}34$ V (cathode)

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= (+0{,}34) - (-0{,}76) = +1{,}10$ V

Since E°_{cell} is positive, the reaction is spontaneous under standard conditions.

10. DETERMINING ELECTRODE POTENTIAL (E_{cell})

- Electrode potential is a measure of the tendency of a half-cell to be reduced.
- For a galvanic cell:
  $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
- A more positive value means s.

KEY REMEMBER

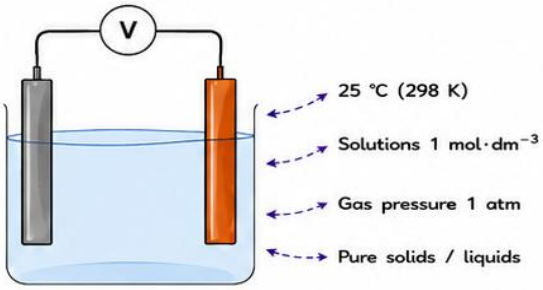
- Galvanic cell: Chemical energy → Electrical energy
- Anode (−): Oxidation occurs. Electrons are released.
- Cathode (+): Reduction occurs. Electrons are gained.
- Salt bridge: Completes circuit and maintains neutrality.
- Positive E°_{cell}: Reaction is spontaneous.

1. STANDARD CONDITIONS

Standard electrode potentials (E°) are determined under standard conditions so that values for different half-cells can be compared.

STANDARD CONDITIONS

- Temperature = 25 °C (298 K)
- All solutions have a concentration of 1 mol·dm^{-3}
- Gases have a pressure of 1 atm (101,3 kPa)
- The cell is measured under these conditions.



2. ALTERNATE ENERGIES

Alternate energies are renewable or cleaner energy sources that reduce our dependence on fossil fuels.

2.1 BIODIESEL

- Biodiesel is a renewable fuel made from vegetable oils, animal fats or recycled cooking oil.
- It is produced by a process called transesterification.

Triglyceride (oil/fat) + 3 Methanol $\xrightarrow{NaOH/KOH}$ 3 Methyl ester (biodiesel) + Glycerol

USES

- Used in diesel engines in place of or blended with conventional diesel (e.g. B20 = 20% biodiesel).

ADVANTAGES

- ✓ Renewable and biodegradable
- ✓ Lower emissions of CO_2, unburned hydrocarbons and sulphur
- ✓ Reduces reliance on fossil fuels



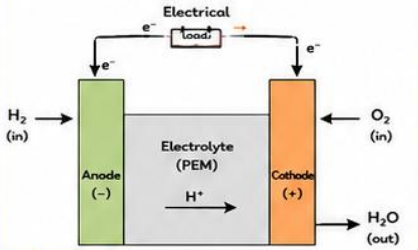
ENVIRONMENTAL IMPACT

- Lower greenhouse gas emissions over its life cycle.
- Non-toxic and less harmful to the environment.
- Large-scale production may compete with food crops and require land and water.

2.2 FUEL CELL

- A fuel cell converts the chemical energy of a fuel directly into electrical energy through a redox reaction.
- Hydrogen fuel cells are the most common.

EXAMPLE: HYDROGEN → OXYGEN FUEL CELL



Half-reactions

Anode (oxidation): $H_2 \rightarrow 2H^+ + 2e^-$

Cathode (reduction): $\frac{1}{2}O_2 + 2H^+ + 2e^- \rightarrow H_2O$

Overall reaction: $2H_2 + O_2 \rightarrow 2H_2O$ + electricity

USES

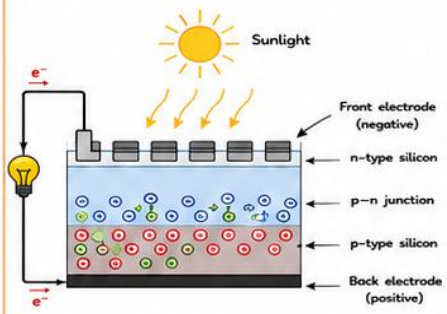
- Used in fuel cell vehicles, backup power systems, and portable power.

ENVIRONMENTAL IMPACT

- Produces only water as a by-product.
- Very low greenhouse gas emissions.
- Hydrogen production must be sustainable (e.g. from electrolysis using renewables).

2.3 PHOTOVOLTAIC CELL

- A photovoltaic (PV) cell converts light energy directly into electrical energy using the photovoltaic effect.
- Made from semiconductor materials (usually silicon).



USES

- Used in solar panels for homes, businesses, satellites, calculators, street lights and solar farms.

ENVIRONMENTAL IMPACT

- Produces clean electricity with no air pollution during operation.
- Reduces greenhouse gas emissions.
- Manufacturing requires energy and materials; panels must be recycled at end of life.

3. USE OF ALTERNATE ENERGIES AND THEIR ENVIRONMENTAL IMPACT

USE OF ALTERNATE ENERGIES

- Provide sustainable, long-term energy resources.
- Reduce dependence on non-renewable fossil fuels (coal, oil, natural gas).
- Help provide energy security and reduce fuel costs.
- Support rural and remote areas with local energy solutions.
- Promote technological innovation and create jobs in green industries.

ENVIRONMENTAL IMPACT – COMPARISON

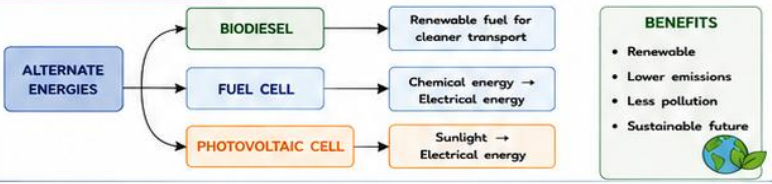
Energy Source	Renewable?	Emissions (GHG)	Air Pollution	Waste / By-products	Overall Impact
Biodiesel	Yes	Lower	Lower (low SO_x, CO, HC)	Biodegradable	Positive (if sustainably produced)
Fuel cell (H_2)	Yes*	Very low (water only)	None	Water	Very positive (clean energy)
Photovoltaic (solar)	Yes	None during operation	None	Solid waste from old panels	Positive (recyclable, low impact)

* Dependent on how hydrogen is produced. Green hydrogen (from renewable electricity) has very low emissions.

KEY TAKEAWAYS

- ☑ Standard conditions ensure fair comparison of electrode potentials.
- ☑ Alternate energies such as biodiesel, fuel cells and photovoltaic cells provide cleaner, sustainable solutions.
- ☑ These technologies reduce environmental impact and help create a greener future.

SUMMARY DIAGRAM



Worked Example 1

The Grade 12 learners conducted an experiment to determine the emf of an Al–Cu galvanic cell under standard conditions.

1.1	State the standard conditions under which the emf of the cell is determined.	(3)
1.2	Identify the anode in the Al–Cu galvanic cell.	(1)
1.3	Identify the cathode in the Al–Cu galvanic cell.	(1)
1.4	Write the reduction half-reaction occurring at the copper electrode.	(3)
1.5	Write the oxidation half-reaction occurring at the aluminium electrode.	(3)
1.6	State the direction of electron flow in the external circuit.	(2)
		[16]

Solutions

1.1	• Temperature = **25 °C (298 K)** ✓ • Concentration of solutions = **1 mol·dm^{-3}** ✓ • Pressure of gases = **101,3 kPa (1 atm)** ✓	(3)
1.2	Aluminium (Al) electrode. ✓	(1)
1.3	Copper (Cu) electrode. ✓	(1)
1.4	Cu^{2+} (aq) + 2e$^-$ →Cu(s) ✓✓✓	(3)
1.5	Al(s)→Al^{3+} (aq) +3e$^-$ ✓✓✓	(3)
1.6	Electrons flow from the aluminum electrode to the copper electrode. ✓✓	(2)
		[16]

Worked Example 2

Study the information below and answer the questions that follow.

A copper electrode is connected to an unknown metal electrode; **X**. Metal **X** has a **higher reducing ability** than copper. The cell reaction is **spontaneous** and **exothermic**.

The cell notation is:

$$\boxed{X(s) \ | \ X^{2+}(aq) \ || \ Cu^{2+}(aq) \ | \ Cu(s)}$$

The initial voltmeter reading is **1.10 V**.

2.1	Identify the type of electrochemical cell represented above.	(2)
2.2	State the energy conversion that takes place in this cell.	(2)
2.3	Which electrode is the anode?	(1)
2.4	State the process that occurs at the anode.	(1)
2.5	Write the oxidation half-reaction taking place at the anode.	(3)
2.6	State the direction of electron flow in the external circuit.	(2)
		[11]

Solutions

2.1	Galvanic (Voltaic) cell. ✓✓	(2)
2.2	Chemical energy is converted into electrical energy. ✓✓	(2)
2.3	**Electrode X** is the anode. ✓	(1)
2.4	Oxidation ✓	(1)
2.5	$X(s) \rightarrow X^{2+}(aq) + 2e^-$ ✓✓✓	(3)
2.6	Electrons flow from the **X electrode** to the **Cu electrode**. ✓✓	(2)
		[11]

ACTIVITY 5

5.1 The cell notation below represents an electrochemical cell:

Cu/Cu^{2+} (1 mol.dm^{-3}) // Ag^+ (1 mol.dm^{-3}) / Ag 298 K/25 ºC

5.1.1 What energy conversion is taking place in the above cell? (2)

5.1.2 Write down TWO indicators from the cell notation that prove that the cell is operating under standard conditions. (2)

5.2 Write down a balanced half-reaction that occurs at the:

5.2.1 Cathode (2)

5.2.2 Anode (2)

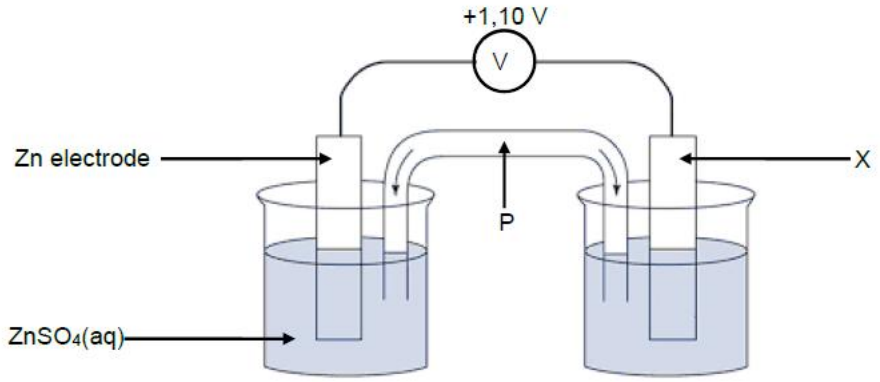
5.3 Use calculations to conclude whether the reaction is SPONTANEOUS or NON-SPONTANEOUS. (5)

5.5 Give a reason for the answer to QUESTION 5.3. (1)

[14]

ACTIVITY 6

Study the following diagram of a Galvanic cell.



6.1 Give two reasons why the above cell can be classified as a galvanic cell. (2)

6.2 Use a calculation to identify electrode X (4)

6.3 Name component P and give one function of this component in this cell. (2)

6.4 Is the Zinc in this cell the oxidizing agent or the reducing agent? (1)

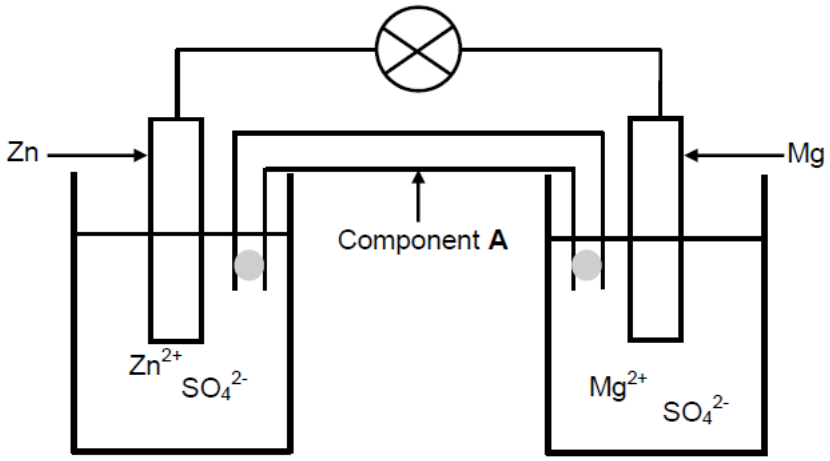
6.5 Write down the half reaction that occur at the cathode of this cell. (2)

6.6 Write down the cell notation of the electrochemical cell above. (3)

[14]

ACTIVITY 7

Learners performed an experiment to determine the electrode potential of an electrochemical cell under standard conditions. They assembled the apparatus, as shown in the diagram below.



7.1 State the energy conversion taking place in this cell. (2)

7.2 Write down a balanced net ionic reaction of the cell. (2)

7.3 In which direction will the $SO4^{2-}$ ions migrate through the salt bridge? Write down only FROM Zn TO Mg or FROM Mg TO Zn. (1)

7.4 Is the cell reaction spontaneous or non-spontaneous? (1)

7.5 Calculate the emf of the cell. (1)

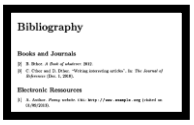
7.6 Component A is removed.

7.6.1 Write down the NAME of component A. (1)

7.6.2 Will the light bulb glow? Write down YES or NO. (1)

7.6.3 Explain the answer to QUESTION 7.2.2. (2)

[14]



BIBLIOGRAPHY

1	Department of Basic Education 2013-2025. The Curriculum Assessment and PolicyStatement National and Provincial question papers.
2	Department of Basic Education 2021. The Curriculum Assessment and Policy Statementexamination guideline. Pretoria: Government Printing Works.
3	National Senior Certificate, 2016 – 2024 diagnostic reports. Pretoria: Government PrintingWorks.
4	Science Clinic Grade 12 DBE Smart Prep
5	Technical Sciences June 2023 Eastern Cape
6	Technical Sciences Preparatory Exam 2024 Limpopo
7	Technical Sciences Preparatory Exam 2024 North West
8	Technical Sciences Preparatory Exam 2024 Gauteng
9	Technical Sciences IEB Final Exam 2019
10	Technical Sciences Preparatory Exam 2024 Eastern Cape
11	JENN Physical Sciences Content manual 2025