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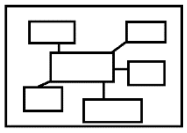
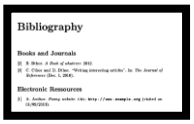
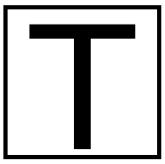
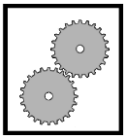
PHYSICAL SCIENCE

GRADE 12

SPRING CLASSES 2026

TEACHER AND LEARNER CONTENT MANUAL

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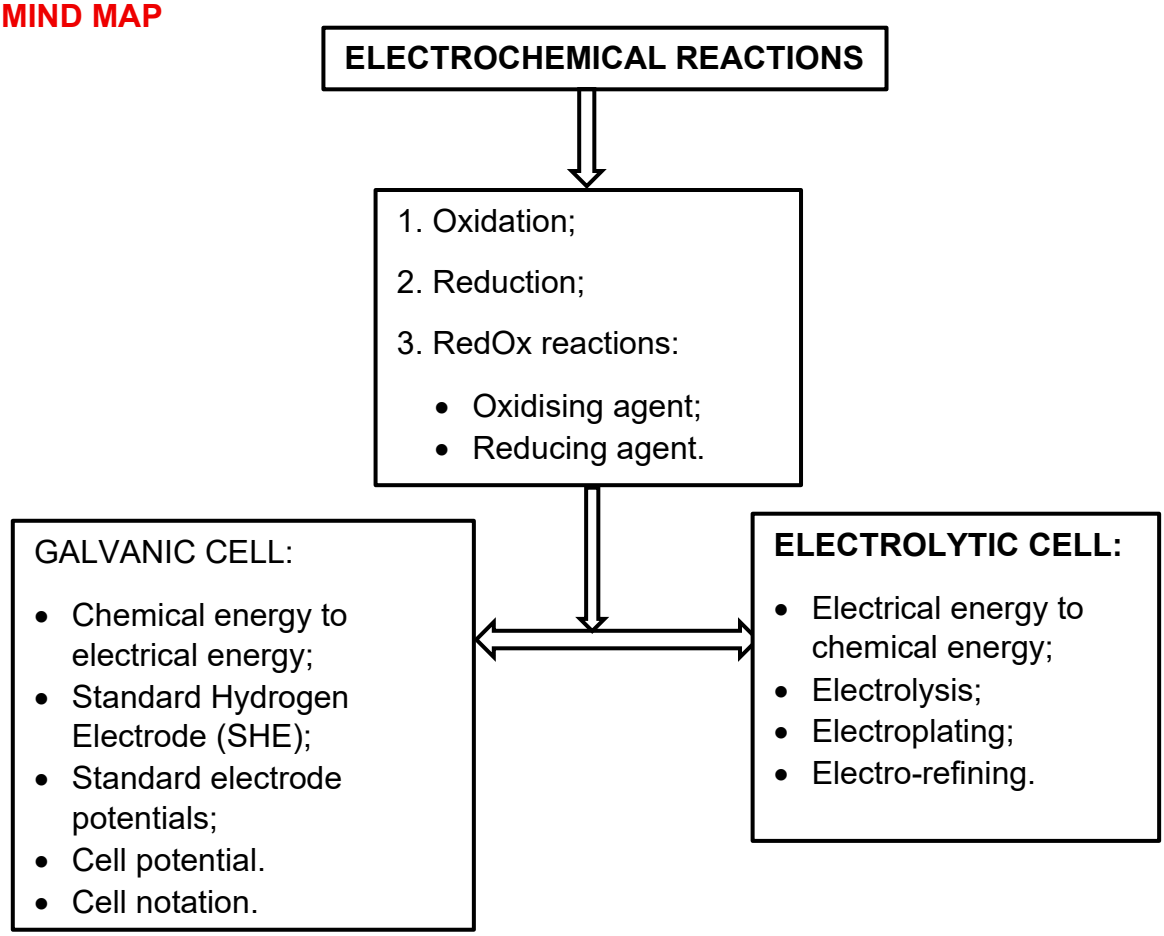
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MIND MAP



EXAMINATION GUIDELINE

Electrochemical reactions

Galvanic cells:

- Define the galvanic cell as a cell in which chemical energy is converted to electrical energy.
- Define oxidation and reduction in terms of electron (e-) transfer:
  Oxidation is a loss of electrons. Reduction is a gain of electrons.
- Define oxidation and reduction in terms of oxidation numbers:
  Oxidation: an increase in oxidation number.
  Reduction: a decrease in oxidation number.
- Define an oxidising agent and a reducing agent in terms of oxidation and reduction:
  Oxidising agent: a substance that is reduced/gains electrons.
  Reducing agent: a substance that is oxidised/loses electrons.
- Define an anode and a cathode in terms of oxidation and reduction:
  Anode: the electrode where oxidation takes place.
  Cathode: the electrode where reduction takes place.
- Define an electrolyte as a substance of which the aqueous solution contains ions OR a substance that dissolves in water to give a solution that conducts electricity.

Relation of current and potential difference to rate and equilibrium

- State that the potential difference of a galvanic cell (V_{cell}) is related to the extent to which the spontaneous cell reaction has reached equilibrium.
- State and use the qualitative relationship between V_{cell} and the concentration of product ions and reactant ions for the spontaneous reaction, namely V_{cell} decreases as the concentration of product ions increases and the concentration of reactant ions decreases until equilibrium is reached at which the $V_{cell} = 0$ (the cell is 'flat'). (Qualitative only. Nernst equation is NOT required.)

Understanding of the processes and redox reactions taking place in galvanic cells

- Describe the movement of ions in the solutions.
- 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.
- Use cell notation or diagrams to represent a galvanic cell.
  When writing cell notation, the following convention should be used:
  - The $H_2|H^+$ half-cell is treated just like any other half-cell.
  - Cell terminals (electrodes) are written on the outside of the cell notation.
  - Active electrodes:
    reducing agent / oxidised species // oxidising agent / reduced species.
  - Inert electrodes (usually Pt or C):
    Pt / reducing agent / oxidised species // oxidising agent / reduced species / Pt
    Example: Pt / $C\ell^-$(aq) /$C\ell_2$(g) // F_2(g) / F^-(aq) / Pt.
- 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.

Standard electrode potentials

- Write down the standard conditions under which standard electrode potentials are determined.
- Describe the standard hydrogen electrode and explain its role as the reference electrode.
- Explain how standard electrode potentials can be determined using the reference electrode and state the convention regarding positive and negative values.

Electrolytic cells

- Define the electrolytic cell as a cell in which electrical energy is converted into chemical energy.
- Electrolysis: The chemical process in which electrical energy is converted to chemical energy OR the use of electrical energy to produce a chemical change.

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 processes:
  - The decomposition of copper(II) chloride.
  - Electroplating, e.g. the electroplating of an iron spoon with silver/nickel.
  - Refining of metals, e.g. copper.
  - The electrolysis of a concentrated solution of sodium chloride.

REVISION OF GRADE 11 WORK:

- **Oxidation** involves a **loss** of electrons;

OR Oxidation is an **increase** in the oxidation number of a substance.
Examples are:
(a) $Na \rightarrow Na^{+} + e^{-}$;
(b) $Mg \rightarrow Mg^{2+} + 2e^{-}$;
(c) $Fe \rightarrow Fe^{3+} + 3e^{-}$;
(d) $2Cl^{-} \rightarrow C\ell_2 + 2e^{-}$.

- **Reduction** involves a **gain** of electrons.

OR Reduction is a **decrease** in the oxidation number of a substance.
Examples are:
(a) $C\ell_2 + 2e^{-} \rightarrow 2Cl^{-}$;
(b) $O_2 + 4e^{-} \rightarrow 2O^{2-}$;
(c) $Na^{+} + e^{-} \rightarrow Na$;
(d) $Mg^{2+} + 2e^{-} \rightarrow Mg$.

An easy way to remember this is:



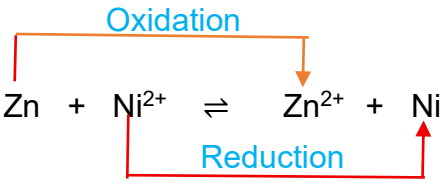
RedOx reactions: are reactions in which oxidation and reduction take place at the same time.
In a redox reaction there is a transfer of electrons from one substance to another.
An **oxidizing agent**: is the substance that gains electrons and is reduced**.**
A **reducing agent**: is the substance that loses electrons and undergoes oxidation**.**
Examples: Identify the oxidation and reduction reactions and the oxidising agent and reducing reagent in each of the following redox reactions.

(a) $2Ag^{+} + Mg \rightleftharpoons 2Ag + Mg^{2+}$

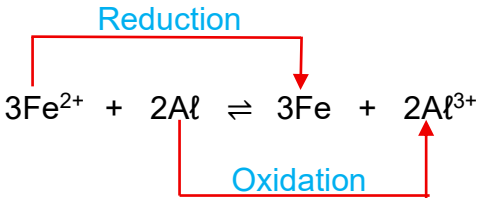
Reduction

Oxidation

- Oxidising agent: Ag^+;
- Reducing agent: Mg.

(b) 

- Oxidising agent: Ni^{2+};
- Reducing agent: Zn.

(c) 

- Oxidising agent: Fe^{2+};
- Reducing agent: Aℓ.

CLASSWORK 1

Identify oxidation and reduction reactions and the oxidizing agent and reducing agent in each of the following balanced redox reactions:

1.1 $Mn + Cu^{2+} \rightleftharpoons Mn^{2+} + Cu$;

1.2 $2Ag^+ + Pb \rightleftharpoons 2Ag + Pb^{2+}$;

1.3 $2Cr + 3Hg^{2+} \rightleftharpoons 2Cr^{3+} + 3Hg$;

1.4 $Sn^{4+} + Ni \rightleftharpoons Sn^{2+} + Ni^{2+}$;

1.5 $2Au^{3+} + 3Co \rightleftharpoons 2Au + 3Co^{2+}$;

1.6 $Cu(s) + 2AgNO_3(aq) \rightleftharpoons Cu(NO_3)_2(aq) + 2Ag(s)$;

1.7 $2NaCℓ(aq) + 2H_2O(ℓ) \rightleftharpoons Cℓ_2(g) + H_2(g) + 2NaOH(aq)$.

Writing half reactions from redox reaction:

Example: write the oxidation and reduction half reactions for each of the following redox reaction:

(a) $2Ag^+ + Mg \rightleftharpoons 2Ag + Mg^{2+}$;

OXIDATION HALF REACTION: $Mg \rightarrow Mg^{2+}$; add electrons to the more positive side to make it the same as the other side;

$Mg \rightarrow Mg^{2+} + 2e^-$

REDUCTION HALF REACTION: $Ag^+ \rightarrow Ag$; add electrons to the more positive side to make it the same as the other side.

$Ag^+ + e^- \rightarrow Ag$.

CLASSWORK 2

Write the oxidation and reduction half reactions for each of the following redox reactions:

2.1 $Mn + Cu^{2+} \rightleftharpoons Mn^{2+} + Cu$;

2.2 $2Ag^{+} + Pb \rightleftharpoons 2Ag + Pb^{2+}$;

2.3 $2Cr + 3Hg^{2+} \rightleftharpoons 2Cr^{3+} + 3Hg$;

2.4 $Sn^{4+} + Ni \rightleftharpoons Sn^{2+} + Ni^{2+}$;

2.5 $2Au^{3+} + 3Co \rightleftharpoons 2Au + 3Co^{2+}$;

2.6 $Cu(s) + 2AgNO_3(aq) \rightleftharpoons Cu(NO_3)_2(aq) + 2Ag(s)$;

2.7 $2NaC\ell(aq) + 2H_2O(\ell) \rightleftharpoons C\ell_2(g) + H_2(g) + 2NaOH(aq)$.

Writing the balanced redox reaction:

Example: write the balanced redox reaction for each of the following pairs of metals:

1. Mn and Ag;

Method 1: Using table 4B

Step 1: Refer to table 4B of the Standard Reduction Potentials

$Mn^{2+} + 2e^{-} \rightleftharpoons Mn$ ..... (Oxidation)

$Ag^{+} + e^{-} \rightleftharpoons Ag$ ..... (Reduction)

Step 2: Re-writing the half-reactions;

- All the half-reactions in the table are written as reduction half-reactions;
- the substance that appears at the top of the other in table 4B undergoes oxidation; This half-reaction is flipped or reversed to make it an oxidation half-reaction.
- and the one that appears below the other undergoes reduction and is written as is from the table;
- Both half-reactions are written with single arrows($\longrightarrow$)

Oxidation half-reaction: $Mn \rightarrow Mn^{2+} + 2e^{-}$

Reduction half-reaction: $Ag^{+} + e^{-} \rightarrow Ag$

Step 3: balance the number of electrons on both sides;

Oxidation half-reaction: $Mn \rightarrow Mn^{2+} + 2e^{-}$

Reduction half-reaction: $Ag^{+} + e^{-} \rightarrow Ag$ x 2

Step 4: writing the balanced overall/net redox reactions.

Oxidation half-reaction: $Mn \rightarrow Mn^{2+} + 2e^{-}$

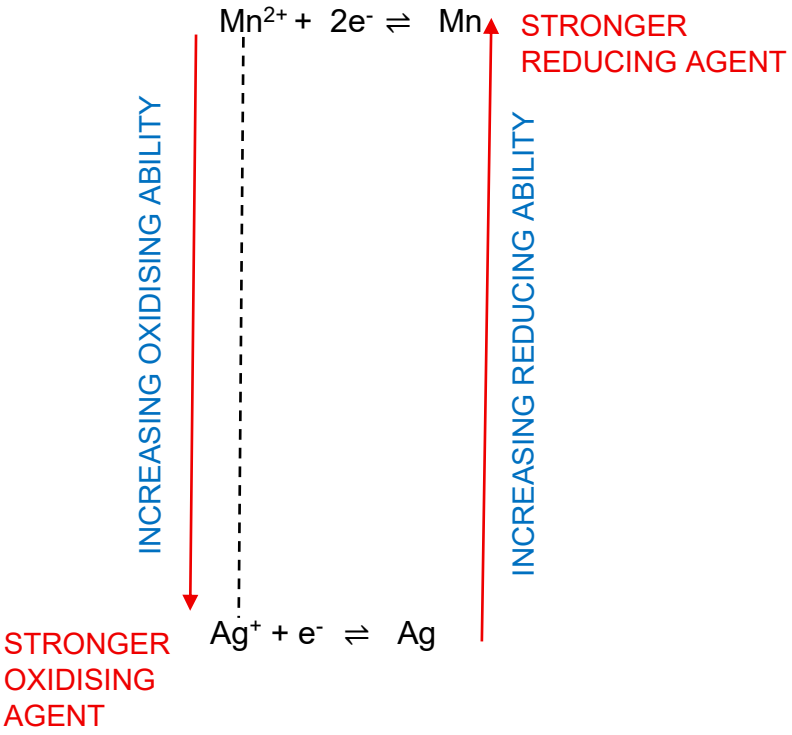
Reduction half-reaction: $2Ag^{+} + 2e^{-} \rightarrow 2Ag$

Balanced redox reaction: $Mn + 2Ag^{+} \rightleftharpoons Mn^{2+} + 2Ag$

Method 2: using the arrows on the sides of the table.

Step 1: using the table to determine;

- The oxidising agent, which undergoes reduction;
- The reducing agent, which undergoes oxidation.



- Ag^+ is a stronger oxidising agent, i.e it undergoes reduction as follows:
  $Ag^+ + e^- \longrightarrow Ag$ ……. Reduction half-reaction
- Mn is a stronger reducing agent, i.e it undergoes oxidation as follows:
  $Mn \longrightarrow Mn^{2+} + 2e^-$ ……. Oxidation half-reaction

Step 2: balance the number of electrons on both sides;

Oxidation half-reaction: $Mn \rightarrow Mn^{2+} + 2e^-$

Reduction half-reaction: $Ag^+ + e^- \rightarrow Ag$ x 2

Step 3: writing the balanced overall/net redox reactions.

Oxidation half-reaction: $Mn \rightarrow Mn^{2+} + 2e^-$

Reduction half-reaction: $2Ag^+ + 2e^- \rightarrow 2Ag$

Balanced redox reaction: $Mn + 2Ag^+ \rightleftharpoons Mn^{2+} + 2Ag$

Method 3: using the standard reduction potentials (E^o)

Step 1: from the table of standard reduction potentials determine the standard reduction potential of each substance.

NOTE: The substance with a more positive standard reduction potential undergoes reduction.

$Mn^{2+} + 2e^- \rightleftharpoons Mn$ $E^\circ_{Mn} = -1{,}18V$; a stronger reducing agent; Oxidation half-reaction.

$Ag^+ + e^- \rightleftharpoons Ag$ $E^\circ_{Ag} = +0{,}80V$; a stronger oxidizing agent; Reduction half-reaction.

Step 2: re-writing the half-reactions:

$Mn \rightarrow Mn^{2+} + 2e^-$ Oxidation half-reaction.

$Ag^+ + e^- \rightarrow Ag$ Reduction half-reaction.

Step 3: balancing the number of electrons in the half-reactions;

$Mn \rightarrow Mn^{2+} + 2e^-$ x 1 Oxidation half reaction.

$Ag^+ + e^- \rightarrow Ag$ x 2 Reduction half-reaction.

Step 4: re-write the half-reactions with the appropriate multiples:

$Mn \rightarrow Mn^{2+} + 2e^-$ Oxidation half-reaction

$2Ag^+ + 2e^- \rightarrow 2Ag$ Reduction half-reaction

Step 5: add the balanced half-reactions, cancelling the electrons on both sides:

$Mn \rightarrow Mn^{2+} + 2e^-$ Oxidation half-reaction.

$2Ag^+ + 2e^- \rightarrow 2Ag$ Reduction half-reaction.

$Mn + 2Ag^+ \longrightarrow Mn^{2+} + 2Ag$ balanced overall/net redox reaction.

CLASSWORK 3

Write a balanced redox reaction for each of the following pairs of substances:

3.1 Aℓ and Pb;

3.2 Cd and Aℓ;

3.3 Zn and Ni;

3.4 Ag and Pt;

GRADE 12 WORK

IMPORTANT TERMS AND DEFINITIONS

1. Electrode: is a rod where oxidation or reduction takes place.

The two types of electrodes are:

(a) Anode: is the electrode where **oxidation** takes place;

(b) Cathode: is the electrode where **reduction** takes place.

Easy way to remember which reaction takes place at which electrode:
AnOx and RedCat OR Vowel to Vowel (A to O) and Consonant to Consonant (C to R)

2. Electrolyte: is a substance of which the aqueous solution contains ions

OR a substance that dissolves in water to give a solution that conducts electricity.

Most electrolytes are solutions of ionic compounds.

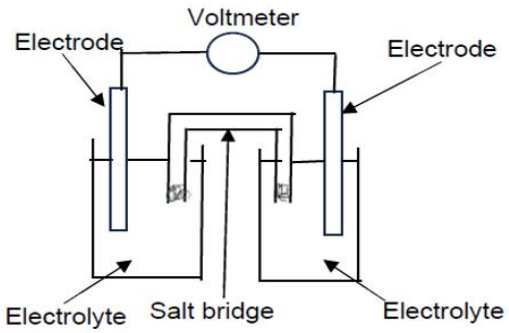
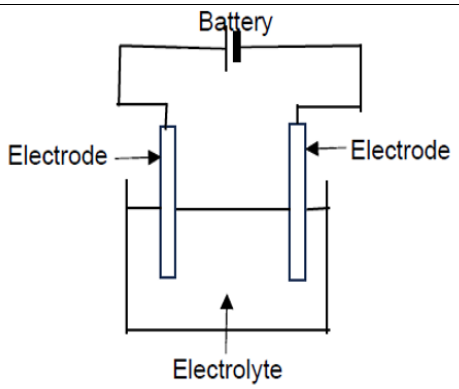
ELECTROCHEMICAL CELLS

Electrochemical cells are cells in which electrochemical (redox) reactions take place and are accompanied by energy changes.

The two types of electrochemical cells are:

1. the galvanic or voltaic cell;
2. the electrolytic cell.

COMPARISON OT THE GALVANIC AND ELECTROLYTIC CELLS

GALVANIC CELL	ELECTROLYTIC CELL
	
1. Energy change: Chemical energy → electrical energy;	1.Energy change: Electrical energy → chemical energy;
2. Reactions are spontaneous, i.e don't need activation energy;	2. Reactions are non-spontaneous, i.e They need activation energy;
3. Reactions are exothermic;	3. Reactions are endothermic;
4. Has a salt bridge;	4. Has no salt bridge;
5. The anode is the negative (-) electrode;	5. The anode is the positive (+) Electrode;
6. The cathode is the positive (+) electrode;	6. The cathode is the negative (-) electrode;
7. The emf of the cell is always positive.	7. The emf of the cell is always negative.

A. THE GALVANIC/VOLTAIC CELL

Consists of the following components:

(a) two electrodes:(i) the anode (-) where oxidation takes place;

(ii) the cathode (+) where reduction takes place.

(b) two electrolytes: representing two half-cells;

(c) the salt bridge: containing an ionic solution of potassium chloride (KCl) or potassium nitrate (KNO_3).

The functions of the salt bridge are:

(i) to connect the two half cells or to complete the circuit;

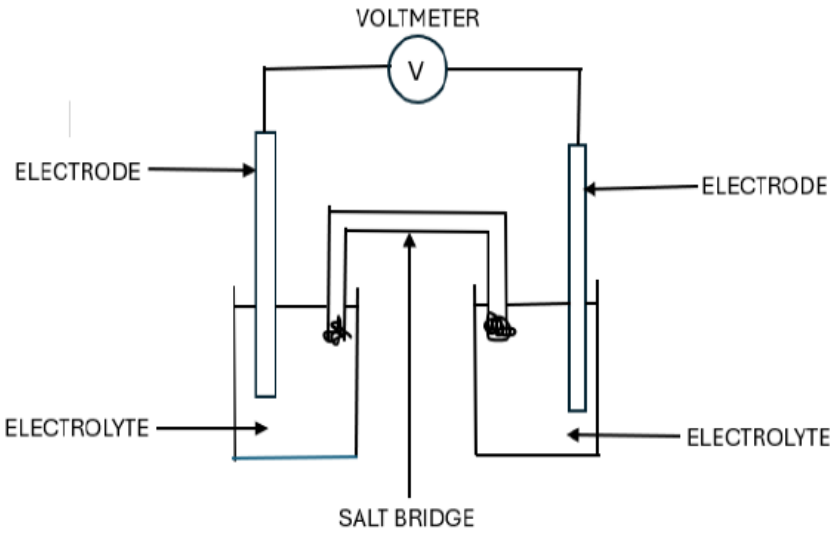
(ii) to allow for the movement of ions;

(iii) to maintain the electrical neutrality of the cell.

(d) a voltmeter: to measure the cell potential/emf.

(e) connecting wires.

The following is a simplified diagram of a galvanic cell.



A galvanic cell functions under standard conditions of:

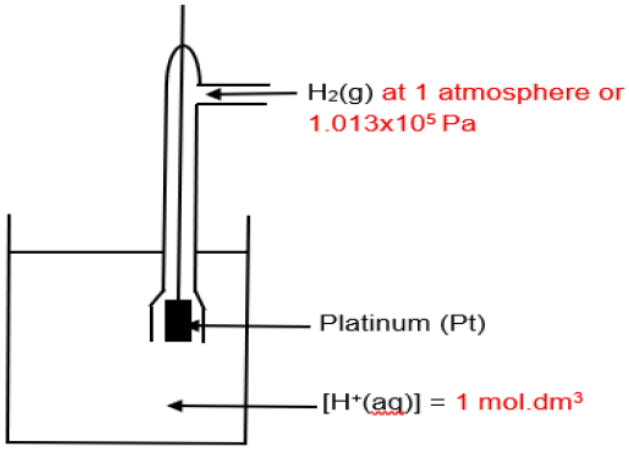
(i) Concentration: 1 $mol.dm^3$;

(ii) Temperature : 25ºC / 298 K;

(iii) Pressure: 1 atmosphere / 1,013 x 10^5 Pa.

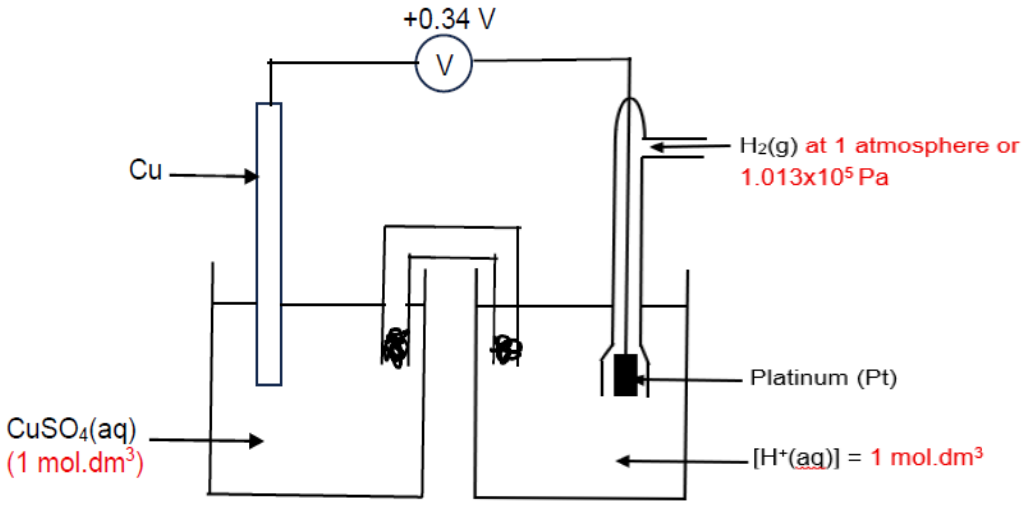
THE STANDARD REDUCTION POTENTIALS AND TABLES

- To measure the standard reduction potentials of all other electrodes, the hydrogen electrode was chosen as the standard electrode (known as the Standard Hydrogen Electrode = SHE);
- It was assigned a standard reduction potential of 0.00V.
  The diagram below shows the Standard Hydrogen Electrode:



Platinum is used as an electrode in the SHE because :
- platinum is a conductor of electricity;
- platinum is inert, i.e it does not take part in the reactions going on in the cell;
- platinum increases the surface area for reaction to take place.

For example : the standard reduction potential of copper (Cu) was measured as below:



The standard reduction potential of copper was found to be +0.34 V.
Standard reduction potentials of all other substances were measured in the same way.
Then tables of standard reduction potentials were drawn (Tables 4A and 4B).

CELL POTENTIAL

Cell potential is calculated using the formula:

$$E^0_{CELL} = E^0_{CATHODE} - E^0_{ANODE}$$

For example, for copper (Cu) and silver (Ag);

- Copper: E^0_{Cu} = +0.34 V;
- Silver: E^0_{Ag} = +0.80 V;
- Ag has a more positive standard electrode potential, it undergoes reduction and it is the cathode;
- Cu undergoes oxidation, it is the anode;
- $E^0_{CELL} = E^0_{CATHODE} - E^0_{ANODE}$
  $E^0_{CELL} = E^0_{Ag} - E^0_{Cu}$
  $E^0_{CELL} = (+0.80) - (+0.34)$
  E^0_{CELL} = 0.46 V

Factors which influence cell potential:

The cell potential of a galvanic cell depends on the following factors:

• Specific reactions that take place at the anode and the cathode;

- The relative strength of oxidising and reducing agents influences the cell potential.
- The stronger the oxidising and reducing agent that are combined, the greater the cell potential will be.
- **Thus, the far apart the substances are on the tables of standard reduction tables, the greater the cell potential.**

• Concentrations of reactants and products;

- Concentration of reactants and products Under standard conditions in a galvanic cell is 1 mol·dm^{-3};
- The galvanic cell described by the cell reaction below has a cell potential of 0,48 V.

$$2A\ell(s) + 3Mn^{2+}(aq) \rightleftharpoons 2A\ell^{3+}(aq) + 3Mn(s)$$

If the concentration of the Mn^{2+} ions increases (greater than1 mol·dm^{-3}):

- the forward reaction will be favoured as can be predicted by using Le Chatelier"s principle;
- consequently the cell potential will be greater than 0,48 V.

If the concentration of the $A\ell^{3+}$ ions increases (greater than 1 mol·dm^{-3});

- the reverse reaction will be favoured as can be predicted by using Le Chatelier"s principle;
- consequently the cell potential will be less than 0,48 V.

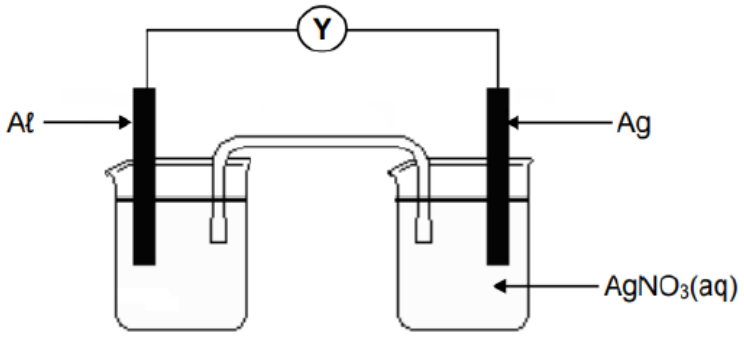
Initially the concentration of the reactants in a galvanic cell is high and the cell potential is at its maximum. As the reaction progresses the concentration of reactants decreases and the concentration of products increases. The result is a decrease in cell potential.

We can therefore say that an increase in the concentration of the reactants will increase the cell potential and that an increase in the concentration of the products will decrease the cell potential. When equilibrium is reached, the cell potential has dropped to 0 V.

ACTIVITIES

ACTIVITY 1

1.1 Learners set up a galvanic cell and measure its emf under standard conditions.



1.1.1 Write down the name of component **Y**. (1)
1.1.2 Is Aℓ the ANODE or the CATHODE? (1)
1.1.3 Write down the oxidation half-reaction. (2)
1.1.4 Write down the reduction half-reaction. (2)
1.1.5 Write down the balanced net / overall cell reaction for this cell. (3)
1.1.6 Write down the cell notation for this cell. (3)
1.1.7 Calculate the emf of this cell. (3)
1.1.8 Do electrons flow from Al to Ag or from Ag to Al? (1)
1.1.9 Do cations (K^+) move from the salt bridge to Al or from the salt bridge to Ag? (1)
1.1.10 Which electrode (Al or Ag) decreases in mass? (1)

1.2 Consider the half-cells, **P**, **Q** and **R**, represented in the table below.

HALF-CELL		
P	**Q**	**R**
$Zn \mid Zn^{2+}(aq)$	$C\ell_2 \mid C\ell^-(aq)$	$Cu \mid Cu^{2+}(aq)$

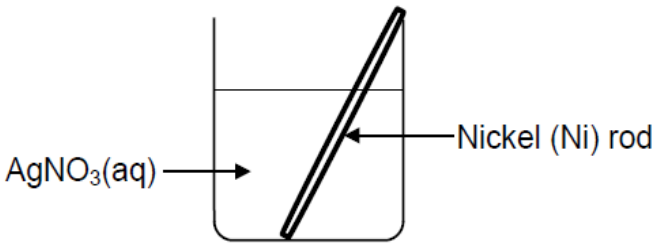
Different combinations of the half-cells above are compared to determine the highest emf produced under standard conditions.

1.2.1 Write down the NAME of a suitable electrode for half-cell **Q**. (1)
1.2.2 State the standard conditions under which the half-cells should operate to ensure a fair comparison. (2)
1.2.3 Write down the NAME or FORMULA of the strongest reducing agent in the half-cells above. (1)
1.2.4 Which combination of half-cells will produce the highest emf? Choose from **PR**, **PQ** or **QR**. (NO calculation is required.) (1)

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ACTIVITY 2

2.1 A nickel (Ni) rod is placed in a beaker containing a silver nitrate solution, $AgNO_3(aq)$ and a reaction takes place.



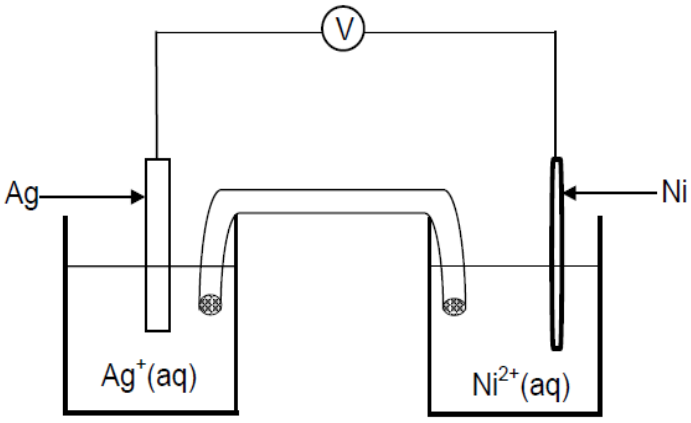
Write down the:

2.1.1 NAME or FORMULA of the electrolyte. (1)

2.1.2 Oxidation half-reaction that takes place. (2)

2.1.3 Balanced equation for the net (overall) redox reaction that takes Place. (3)

2.2 A galvanic cell is now set up using a nickel half-cell and a silver half-cell.



2.2.1 Which electrode (Ni or Ag) must be connected to the negative terminal of the voltmeter? Give a reason for the answer. (2)

2.2.2 Write down the cell notation for the galvanic cell above. (3)

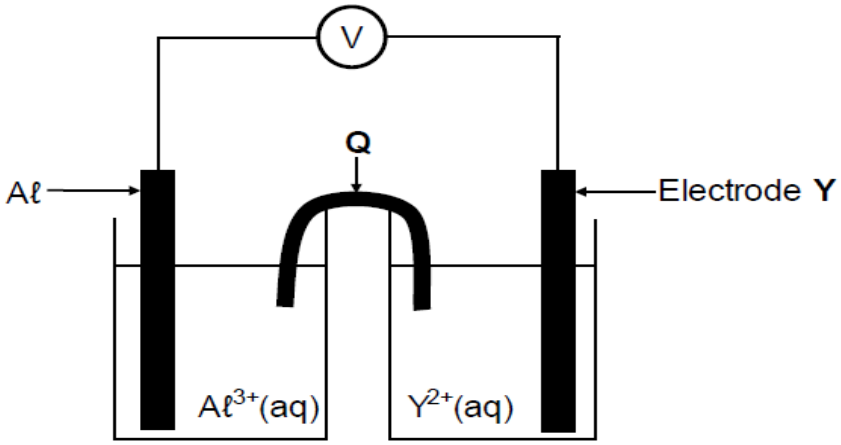
2.2.3 Calculate the initial reading on the voltmeter if the cell functions under standard conditions. (4)

2.2.4 How will the voltmeter reading in QUESTION 2.2.3 be affected if the concentration of the silver ions is increased? Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

[16]

ACTIVITY 3

In the electrochemical cell shown below an aluminium electrode and another metal electrode, **Y**, are used.



3.1 Write down the:

3.1.1 Name of component **Q**. (1)

3.1.2 Type of electrochemical cell represented above. (1)

It is found that the mass of the aluminium electrode increases whilst the cell is functioning.

3.2 How will EACH of the following change while the cell is functioning? Choose from INCREASES, DECREASES or REMAINS THE SAME.

3.2.1 The concentration of $A\ell^{3+}(aq)$. (1)

3.2.2 The concentration of $Y^{2+}(aq)$. (1)

3.3 Write down the:

3.3.1 Half-reaction that takes place at electrode **Y**. (2)

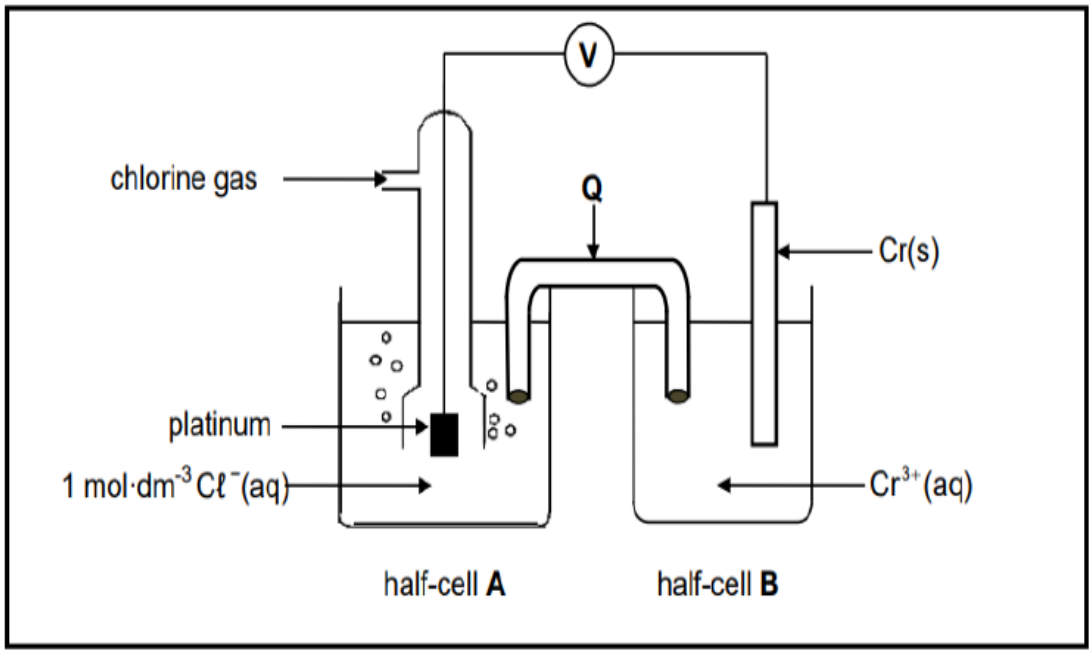
3.3.2 Cell notation of the cell. (3)

3.4 The initial emf of this cell measured under standard conditions is 0,7 V. Identify metal Y by means of a calculation. (5)

[14]

ACTIVITY 4

The electrochemical cell below functions under standard conditions.



4.1 Give a reason why platinum is used as the electrode in half-cell **A**. (1)

4.2 Write down the:

4.2.1 Energy conversion that takes place in this cell. (1)

4.2.2 Half-reaction that takes place at the cathode. (2)

4.2.3 Cell notation for this cell. (3)

4.3 Calculate the initial emf of this cell. (4)

4.4 Silver chloride is an insoluble salt.

What will be the effect on the cell potential when a small amount of silver nitrate solution, $AgNO_3(aq)$, is added to half-cell **A**?

Choose from INCREASES, DECREASES or REMAINS THE SAME. (2)

[13]

ACTIVITY 5

A galvanic cell at standard conditions is represented by the cell notation below. **X** and **Y** are unknown electrodes.

X / $Zn^{2+}(aq)$ // $Fe^{3+}(aq)$, $Fe^{2+}(aq)$ / **Y**

5.1 Write down the NAME or FORMULA of:

5.1.1 Electrode **X**. (1)

5.1.2 Electrode **Y**. (1)

5.1.3 The oxidising agent. (1)

5.2 Write down:

5.2.1 ONE function of electrode **Y**. (1)

5.2.2 The half-reaction that takes place at electrode **Y**. (2)

5.2.3 The net (overall) equation for the cell reaction that takes place in this cell. (3)

5.3 Calculate the initial emf of this cell. (4)

5.4 How will the initial emf of the cell be affected when the concentration of

the iron(III) ions is changed to 0,6 mol·dm^{-3}? Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

[14]

B. THE ELECTROLYTIC CELL

An electrolytic cell is a type of cell in which electrical energy is converted to chemical energy.

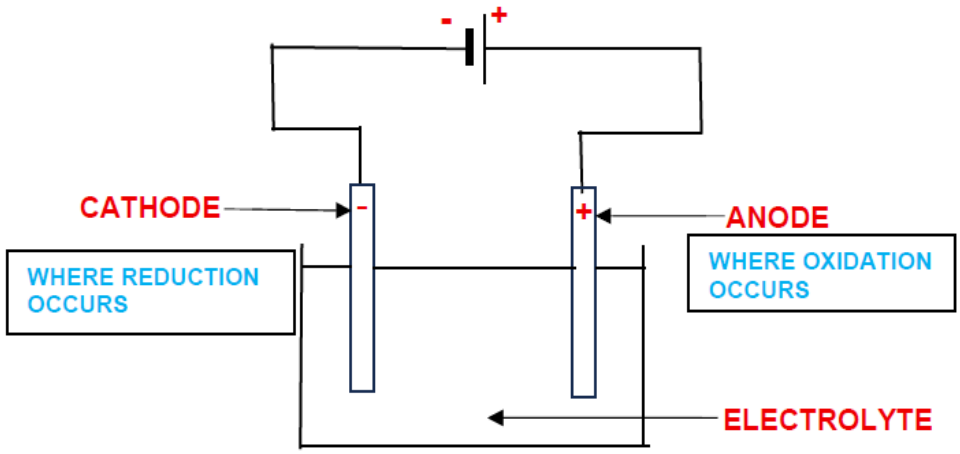
In an electrolytic cell:

- The anode is the electrode connected to the positive terminal of the battery, i.e the positive electrode;
- The cathode is the electrode connected to the negative terminal of the battery, i.e it is the negative electrode;

> **EASY WAY TO REMEMBER:**
>
> **PANIC = Positive Anode; Negative Is Cathode**

- A DC power source (and NOT an AC power source) is used. This is to ensure that:
  (i) the current and electrons flow in only one direction;
  (ii) the polarity of the electrodes remains the same.

The following is a simplified diagram of an electrolytic cell:



The working of an electrolytic cell is based on the process of electrolysis.

Electrolysis: is the process in which electrical energy is converted to chemical energy.

During the process an ionic compound is broken down into its elements.

During the process of electrolysis:

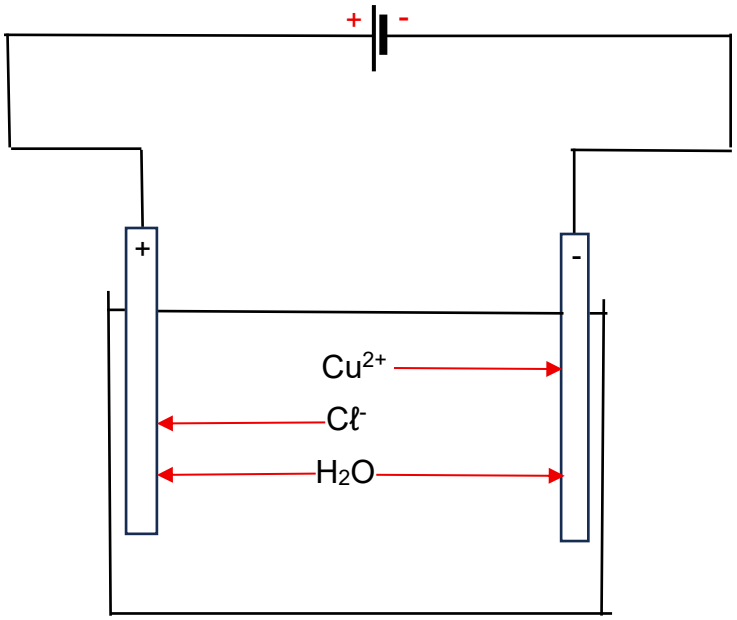
- cations (positive ions) move to the cathode (negative electrode) and they are reduced;
- Multiple cations may move to an electrode, but only one of them will be reduced;
- The cation that is a stronger oxidising agent will be reduced.
- anions (negative ions) move to the anode (positive electrode) and they are oxidised.
- Multiple anions may move to an electrode, but only one of them will be oxidised;

- The anion that is a stronger reducing agent will be oxidised.

  Remember the half-reactions have to be reversed to get oxidation half-reactions and their standard reduction potentials have to be reversed also.
- For solutions, aqueous or concentrated, water (H_2O) moves to both the anode and the cathode.
  - At the cathode, water undergoes reduction as follows:
    $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$ $E^o = -0.83$ V
  - At the anode, water undergoes oxidation as follows:
    $O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$ $E^o = +1.23$ V

    Written as oxidation:
    $2H_2O \longrightarrow O_2 + 4H^+ + 4e^-$ $E^o = -1.23$ V
- **NOTE**: nitrate ions (NO_3^-) and sulphate ions (SO_4^{2-}) are NEVER oxidised at the anode.
  - This is because the sulphur in SO_4^{2-} and the nitrogen in NO_3^- are in their highest oxidation states.

EXAMPLE 1: ELECTROLYSIS OF A COPPER (II) CHLORIDE [$CuC\ell_2(aq)$] SOLUTION USING CARBON (INERT) ELECTRODES



At the anode…….	At the cathode……
• Chloride ions ($C\ell$) and water move to the anode.	• Copper ions (Cu^{2+}) and water move to the cathode.
• Possible half-reactions are: $2C\ell^- \rightarrow C\ell_2 + 2e^-$ $E^o = -1.36V$ $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ $E^o = -1.23V$ $C\ell^-$ is a stronger reducing agent and will be oxidised.	• Possible half-reactions are: $Cu^{2+} + 2e^- \rightarrow Cu$ $E^o = +0.34V$ $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E^o = -0.83V$ Cu^{2+} is a stronger oxidising agent and will be rediced.
• The oxidation half-reaction at the anode is: $2C\ell^- \rightarrow C\ell_2 + 2e^-$	• The reduction half-reaction at the cathode is: $Cu^{2+} + 2e^- \rightarrow Cu$

• It is observed that bubbles of a greenish-yellow gas ($C\ell_2$) collect around the anode.	• It is observed that a reddish-brown layer (Cu) coats the cathode.

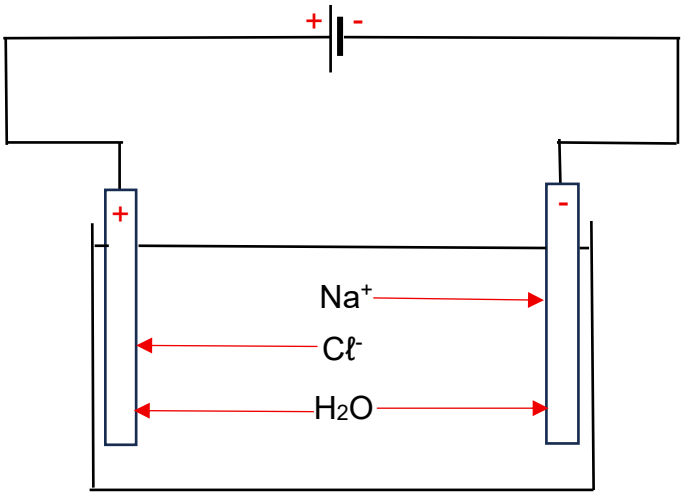
The net cell reaction is:

$$2C\ell^-(aq) \rightarrow C\ell_2(g) + 2e^-$$

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

$$Cu^{2+}(aq) + 2C\ell^-(aq) \rightarrow Cu(s) + C\ell_2(g)$$

EXAMPLE 2: ELECTROLYSIS OF A SODIUM CHLORIDE [NaCℓ(aq)] SOLUTION USING CARBON (INERT) ELECTRODES



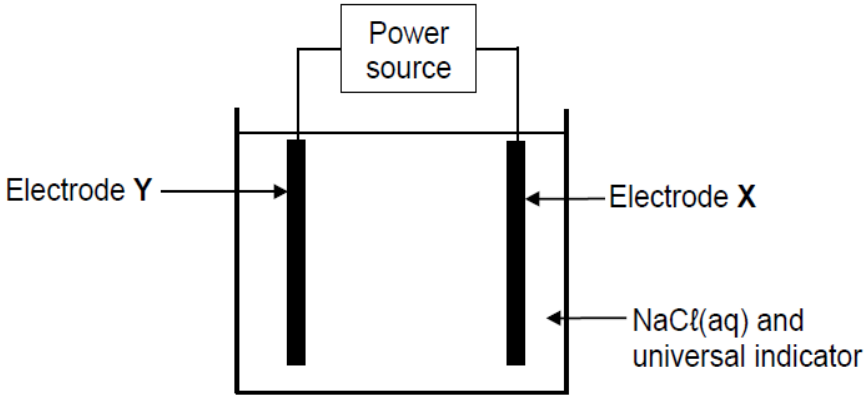
At the anode …….	At the cathode ……..
• Chloride ions ($C\ell$) and water (H_2O) move to the anode.	• Sodium ions (Na^+) and water (H_2O) move to the cathode.
• Possible half-reactions are: $2C\ell^- \rightarrow C\ell_2 + 2e^-$ $E^o = -1.36V$ $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ $E^o = -1.23V$ $C\ell^-$ is a stronger reducing agent and will be oxidised.	• Possible half-reactions are: $Na^+ + e^- \rightarrow Na$ $E^o = -2.71V$ $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E^o = -0.83V$ H_2O is a stronger oxidising agent and will be reduced.
• The oxidation half-reaction at the anode is: $2C\ell^- \rightarrow C\ell_2 + 2e^-$	• The reduction half-reaction at the cathode is: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
• It is observed that bubbles of a greenish-yellow gas ($C\ell_2$) collect around the anode.	• It is observed that the solution around the cathode becomes basic due to presence of OH^-.

ACTIVITIES

ACTIVITY 1

The apparatus below is used to demonstrate the electrolysis of a concentrated sodium chloride solution. Both electrodes are made of carbon. A few drops of universal indicator are added to the electrolyte. The equation for the net cell reaction is:

$$2NaC\ell(aq) + 2H_2O(\ell) \rightarrow C\ell_2(g) + H_2(g) + 2NaOH(aq)$$



Initially the solution has a green colour. Universal indicator becomes red in acidic solutions and purple in alkaline solutions.

1.1 Define the term electrolyte. (2)

When the power source is switched on, the colour of the electrolyte around electrode **Y** changes from green to purple.

1.2 Write down the:

1.2.1 Half-reaction that takes place at electrode **Y**. (2)

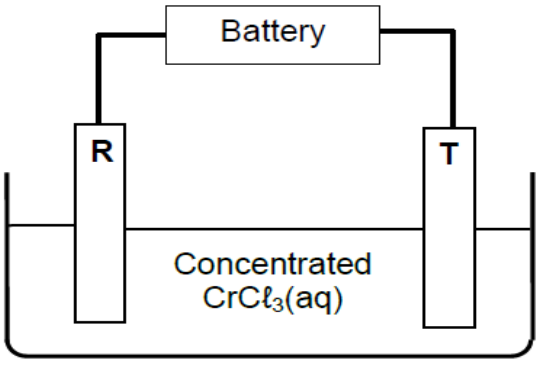
1.2.2 NAME or FORMULA of the gas released at electrode **X**. (1)

1.3 Refer to the Table of Standard Reduction Potentials to explain why hydrogen gas, and not sodium, is formed at the cathode of this cell. (2)

[7]

ACTIVITY 2

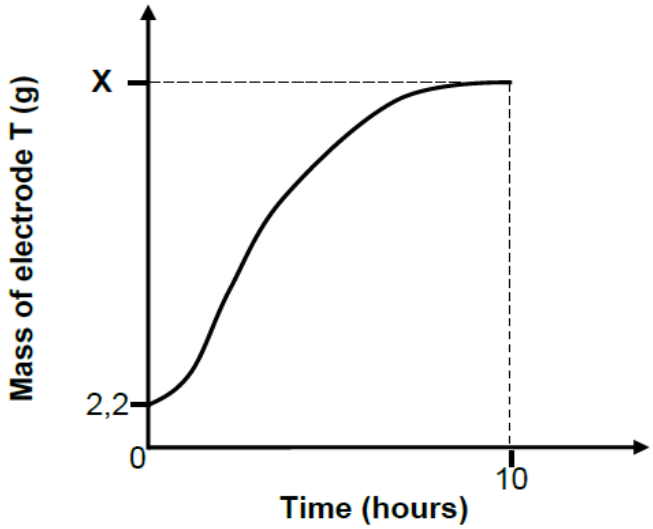
The diagram below represents a simplified cell used for the electrolysis of CONCENTRATED chromium(III) chloride, $CrC\ell3(aq)$. Electrodes **R** and **T** are made of carbon.



The net cell reaction is: $2CrC\ell_3(aq) \rightarrow 2Cr(s) + 3C\ell_2(g)$

2.1 Define the term electrolysis. (2)

2.2 The graph below, NOT drawn to scale, represents the changes in the mass of electrode T during electrolysis.



2.2.1 Write down the half-reaction that takes place at electrode **T**. (2)

A current of 2,5 A passes through the cell for 10 hours.

Calculate the:

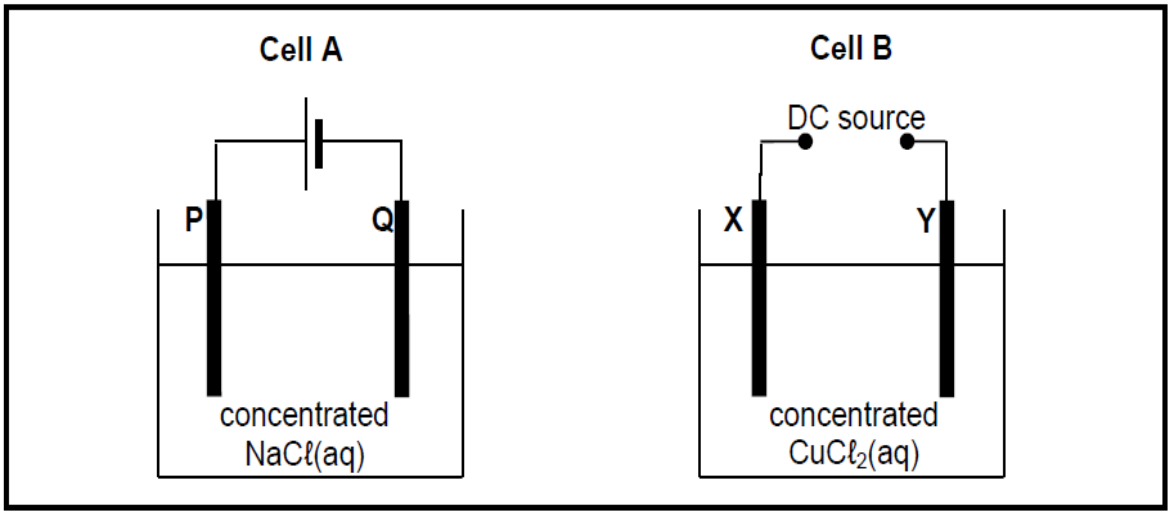
2.2.2 Total charge that flows through the cell during this time. (3)

2.2.3 Value of **X** as shown on the graph. (6)

[13]

ACTIVITY 3

The diagrams below represent two electrochemical cells.
P, **Q**, **X** and **Y** are carbon electrodes.
When cell **B** is functioning, the mass of electrode **X** increases.



3.1 What type of electrochemical cell, GALVANIC or ELECTROLYTIC, is illustrated above? (1)

3.2 Write down the half-reaction that takes place at electrode **Q**. (2)

3.3 The products formed in the two cells are compared.

3.3.1 Name ONE substance that is produced in BOTH cells. (1)

3.3.2 Write down the LETTERS of the TWO electrodes where this product is formed. Choose from **P**, **Q**, **X** and **Y**. (2)

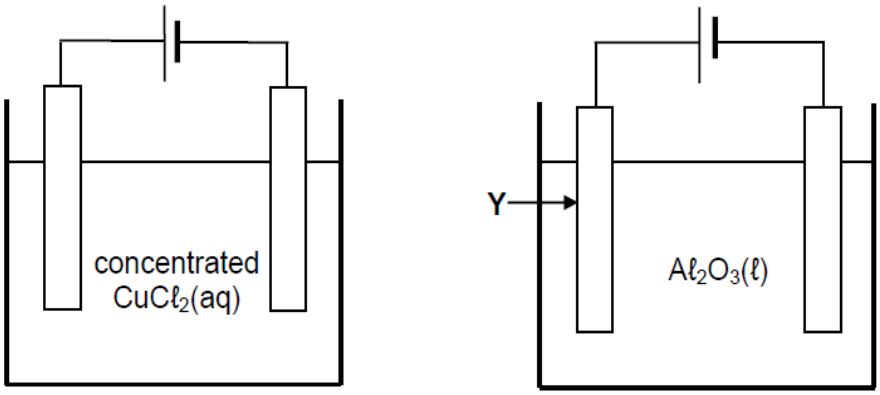
3.4 Is electrode X the CATHODE or the ANODE? Give a reason for the answer. (2)

3.5 Write down the net (overall) cell reaction that takes place in cell **B**. (3)

[11]

ACTIVITY 4

The diagrams below show two electrochemical cells in which carbon electrodes are used. In cell **A**, concentrated copper (II) chloride solution is used and in cell **B**, liquid aluminium oxide is used.



Cell A **Cell B**

4.1 What type of electrochemical cell, ELECTROLYTIC or GALVANIC, is shown above? Give a reason for the answer. (2)

4.2 Write down the:

4.2.1 Half-reaction that takes place at the anode of cell **A**. (2)

4.2.2 Half-reaction that takes place at the cathode of cell **B**. (2)

4.2.3 NAME or FORMULA of the product formed at the cathode of cell **A**. (1)

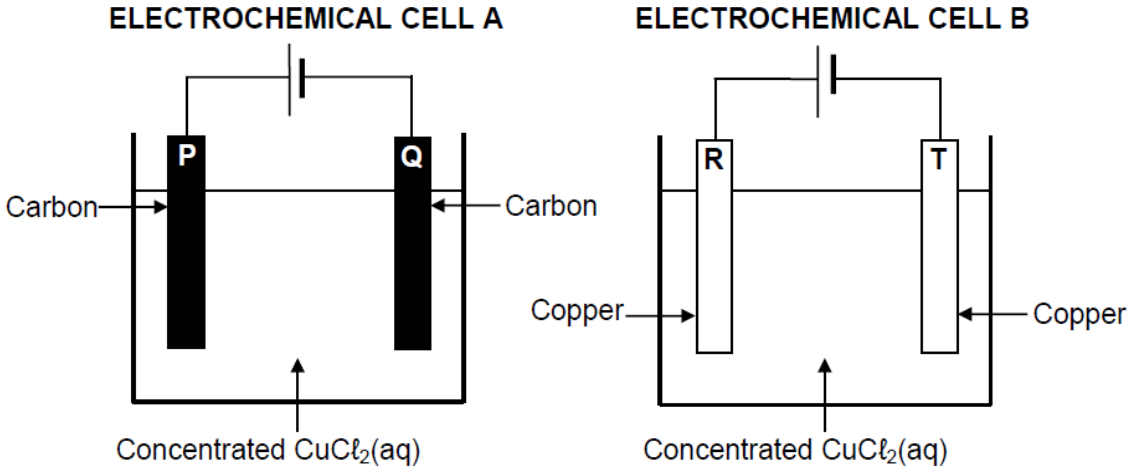
4.3 Give a reason why the mass of electrode **Y** decreases after a while. (1)

[8]

ACTIVITY 5

The simplified diagrams below represent two electrochemical cells, **A** and **B**.

A concentrated copper(II) chloride solution is used as electrolyte in both cells.



5.1 Are A and B ELECTROLYTIC or GALVANIC cells? (1)

5.2 Which of the electrodes (**P**, **Q**, **R** or **T**) will show a mass increase?

Write down a half-reaction to motivate the answer. (4)

5.3 Write down the NAME or FORMULA of the product formed at:

5.3.1 Electrode **P**. (1)

5.3.2 Electrode **R**. (1)

5.4 Fully explain the answer to QUESTION 5.3.2 by referring to the relative strengths of the reducing agents involved. (3)

[10]

PRACTICAL APPLICATIONS OF ELECTROLYSIS

The following are the practical applications of electrolysis:

1. Decomposition of copper (II) chloride;
2. Electroplating;
3. Refining/purification of copper;
4. Electrolysis of a concentrated solution of sodium chloride.

1. ELECTROPLATING

Electroplating is the process in which electrolysis is used to coat one metal over another metal or object.

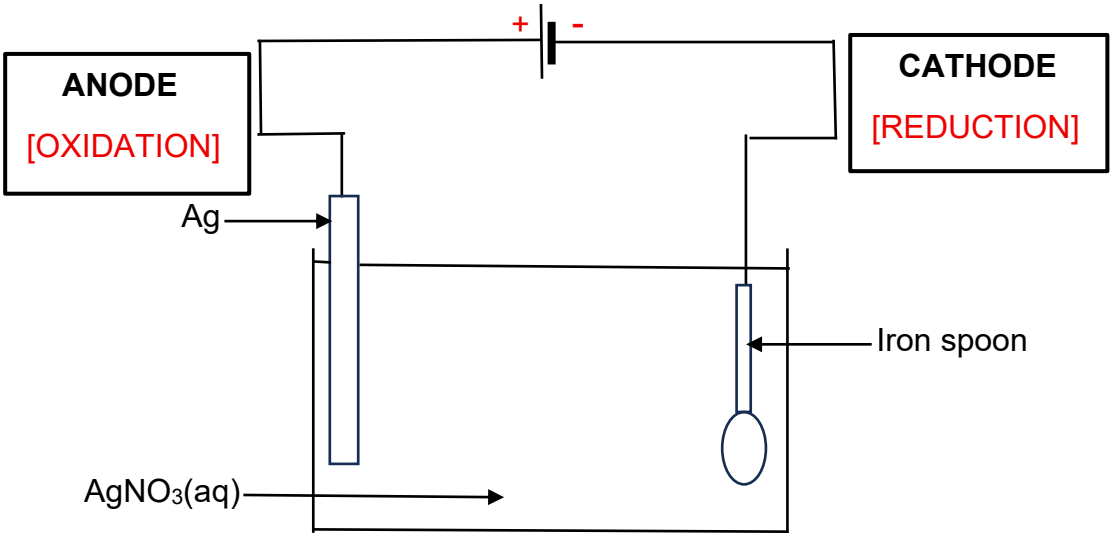
This is in order to:

1. protect the metal underneath;
2. give the metal a shiny and beautiful colour.

During electroplating:

- the metal / object to be electroplated is used as the cathode;
- the coating metal is used as the anode and its solution used as the electrolyte.

The following simplified diagram shows how cutlery, such as spoons or forks, are electroplated with silver.



At the anode……	At the cathode ……
• Nitrate ions (NO_3^-) and water (H_2O) move to the anode. • The anode already contains silver atoms (Ag).	• Silver ions (Ag^+) and water move to the cathode (iron spoon).
• Possible half-reactions are: $Ag \rightarrow Ag^+ + e^-$ $E^o = -0.80V$ $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ $E^o = -1.23V$ NO_3^- are NEVER oxidised because the nitrogen in NO_3^- is in its highest oxidation state; Ag is a stronger reducing agent than H_2O; Ag will be oxidised to Ag^+;	• Possible half-reactions are: $Ag^+ + e^- \rightarrow Ag$ $E^o = +0.80V$ $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E^o = -0.83V$ Ag^+ is a stronger oxidising agent than H_2O; Ag^+ will be reduced to Ag;
• Oxidation half-reaction; $Ag \rightarrow Ag^+ + e^-$	• Reduction half-reaction: $Ag^+ + e^- \rightarrow Ag$
• It is observed that the anode is corroded and decreases in mass.	• It is observed that the iron spoon is coated with a layer of silver metal.

NOTE: The rate of oxidation at the anode is equal to the rate of reduction at the cathode, and so, the concentration of the electrolyte remains the same during this process.

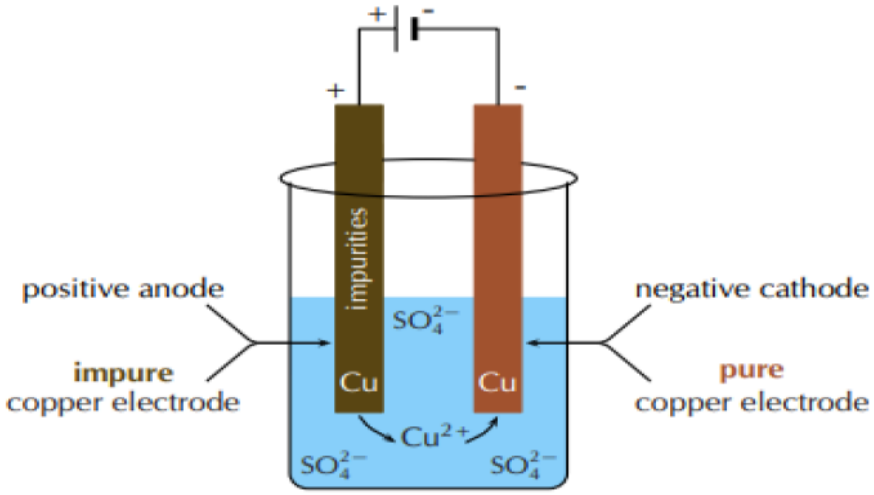
2.ELECTRO-REFINING/ PURIFICATION OF COPPER

When copper is mined, it is an ore, i.e an impure form of copper. It contains impurities such as silver (Ag), gold (Au) and platinum (Pt).

In the electro-refining of copper:

- The impure copper ore is used as the anode;

- Pure copper is used as the cathode;
- An acidified copper (II) sulphate solution is used as the electrolyte;

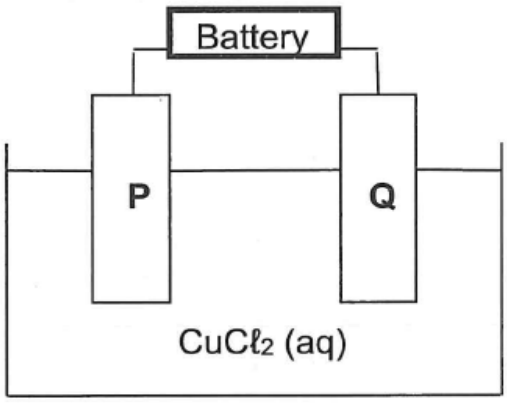


At the anode ……	At the cathode ……
• Sulphate ions (SO_4^{2-}) and water (H_2O) move to the anode. • The anode already contains copper atoms (Cu). • Sulphate ions (SO_4^{2-}) are NEVER oxidised.	• Copper ions (Cu^{2+}) and water move to the cathode.
• Possible half-reactions are: $Cu \rightarrow Cu^{2+} + 2e^-$ $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ Copper (Cu) is a stronger reducing agent than water (H_2O). Copper (Cu) will easily undergo oxidation than water. The other metals impurities are **not** oxidized. They fall to the bottom of the cell and collect as **sludge**.Since they are precious and valuable metals, they are purified and sold.	• Possible half-reactions are: $Cu^{2+} + 2e^- \rightarrow Cu$ $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ Copper ion (Cu^{2+}) is a stronger oxidising agent than water (H_2O). Copper ions will easily undergo reduction than water (H_2O).
• Oxidation half-raection: $Cu \rightarrow Cu^{2+} + 2e^-$	• Reduction half-reaction: $Cu^{2+} + 2e^- \rightarrow Cu$
• It is observed that the anode/impure copper electrode is corroded and decreases in mass.	• It is observed that the cathode/pure copper electrode increases in mass.

ACTIVITIES ON APPLICATIONS OF ELECTROLYSIS

ACTIVITY 6

In the electrolytic cell below, two carbon rods are used as electrodes and a concentrated copper (II) solution ($CuC\ell_2$) is used as electrolyte.



6.1 Define the term electrolyte**.** (2)

When the cell is in operation, a gas is released at electrode **P** while electrode **Q** is covered with a brown layer.

6.2 Write down a half-reaction to explain the observation made at:

6.2.1 Electrode **P**. (2)

6.2.2 Electrode **Q**. (2)

6.3 Which electrode, **P** or **Q**, is the cathode? Give a reason for the answer. (2)

6.4 The carbon rods in the above cell are now replaced with COPPER RODS and the cell is allowed to operate for some time.

The following observations are made at electrode **P**:

- No gas is released;
- Its surface appears rough and corroded.

6.4.1 Refer to the RELATIVE STRENGTHS OF REDUCING AGENTS to explain these observations. (3)

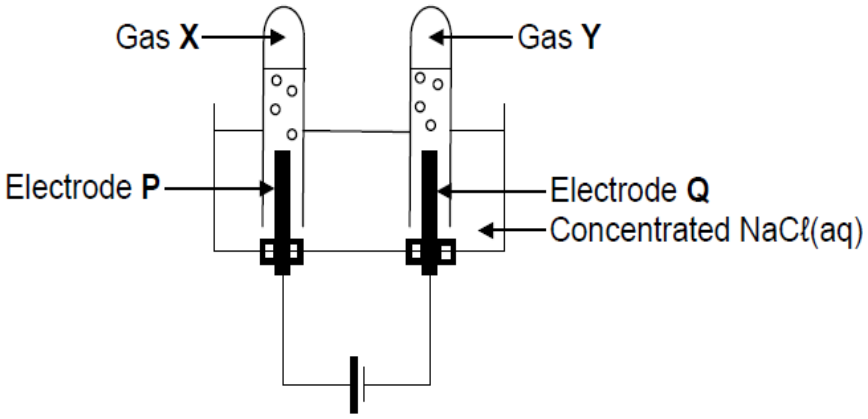
6.4.2 This cell can be used for the purification of copper.

Which electrode (**P** or **Q**) will be replaced with the impure copper during this purification process? (1)

[12]

ACTIVITY 7

In the electrochemical cell below, carbon electrodes are used during the electrolysis of a concentrated sodium chloride solution.



The balanced equation for the net (overall) cell reaction is:

$$2H_2O(\ell) + 2C\ell^-(aq) \rightarrow C\ell_2(g) + H_2(g) + 2OH^-(aq)$$

7.1 Is the reaction EXOTHERMIC or ENDOTHERMIC? (1)

7.2 Is electrode **P** the ANODE or the CATHODE? Give a reason for the answer. (2)

7.3 Write down the:

7.3.1 NAME or FORMULA of gas **X**. (1)

7.3.2 NAME or FORMULA of gas **Y**. (1)

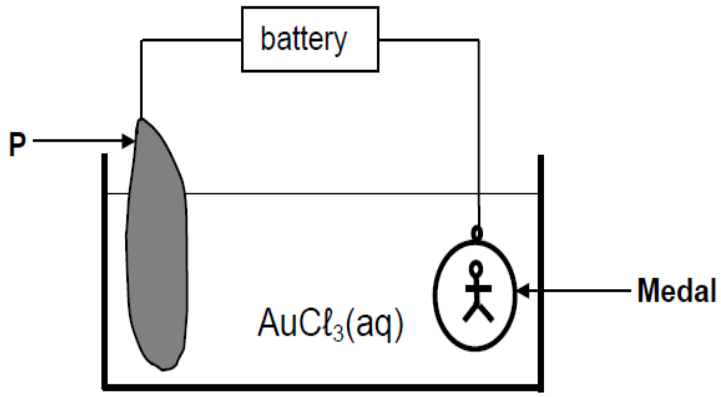
7.3.3 Reduction half-reaction. (2)

7.4 Is the solution in the cell ACIDIC or ALKALINE (BASIC) after completion of the reaction? Give a reason for the answer. (2)

[9]

ACTIVITY 8

The simplified diagram below represents a cell used to electroplate an iron medal with a thin layer of gold.



8.1 Is this an ELECTROLYTIC or a GALVANIC cell? (1)

8.2 Which electrode, **P** or the Medal, is the anode? (1)

8.3 Write down the:

8.3.1 Half-reaction that takes place at electrode **P**. (2)

8.3.2 Oxidation number of gold (Au) in the electrolyte. (1)

8.3.3 Energy change that takes place in this cell. (1)

8.3.4 Visible change that occurs on electrode **P** after the cell functions for a while. (1)

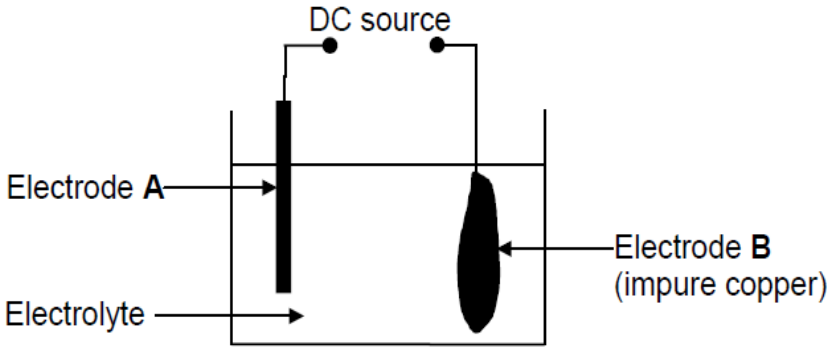
8.4 Besides improving appearance, state ONE other reason why the medal is electroplated. (1)

8.5 State ONE of the two possible changes that should be made to the cell above to electroplate the medal with silver instead of gold. (1)

[9]

ACTIVITY 9

The simplified diagram below represents an electrochemical cell used for the purification of copper.



9.1 Define the term electrolysis. (2)

9.2 Give a reason why a direct-current (DC) source is used in this experiment. (1)

9.3 Write down the half-reaction which takes place at electrode **A**. (2)

9.4 Due to small amounts of zinc impurities in the impure copper, the electrolyte becomes contaminated with Zn^{2+} ions.
Refer to the attached Table of Standard Reduction Potentials to explain why the Zn^{2+} ions will not influence the purity of the copper obtained during this process. (3)

9.5 After the purification of the impure copper was completed, it was found that $2,85 \times 10^{-2}$ moles of copper were formed.
The initial mass of electrode **B** was 2,0 g. Calculate the percentage of copper that was initially present in electrode **B**. (4)

[12]

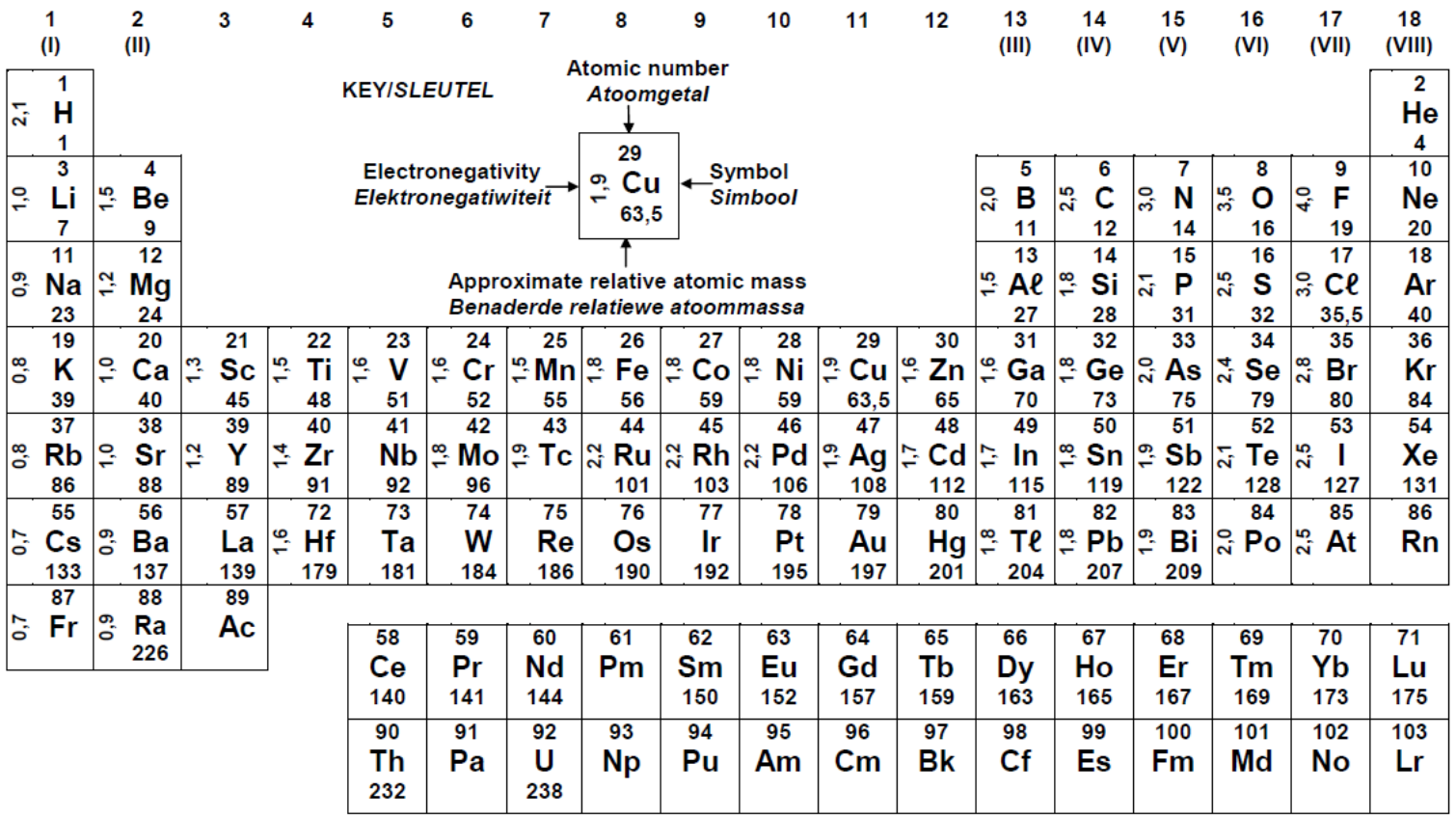
TABLE 1: PHYSICAL CONSTANTS/*TABEL 1: FISIESE KONSTANTES*

NAME/*NAAM*	SYMBOL/*SIMBOOL*	VALUE/*WAARDE*
Standard pressure *Standaarddruk*	p^θ	$1{,}013 \times 10^5$ Pa
Molar gas volume at STP *Molêre gasvolume by STD*	V_m	$22{,}4\ dm^3 \cdot mol^{-1}$
Standard temperature *Standaardtemperatuur*	T^θ	273 K
Charge on electron *Lading op elektron*	e	$-1{,}6 \times 10^{-19}$ C
Avogadro's constant *Avogadro-konstante*	N_A	$6{,}02 \times 10^{23}\ mol^{-1}$

TABLE 2: FORMULAE/*TABEL 2: FORMULES*

$n = \frac{m}{M}$	$n = \frac{N}{N_A}$
$c = \frac{n}{V}$ or/*of* $c = \frac{m}{MV}$	$n = \frac{V}{V_m}$
$\frac{c_a v_a}{c_b v_b} = \frac{n_a}{n_b}$	$pH = -\log[H_3O^+]$
$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}$ at/*by* 298 K	
$E^\theta_{cell} = E^\theta_{cathode} - E^\theta_{anode}$ / $E^\theta_{sel} = E^\theta_{katode} - E^\theta_{anode}$ or/*of* $E^\theta_{cell} = E^\theta_{reduction} - E^\theta_{oxidation}$ / $E^\theta_{sel} = E^\theta_{reduksie} - E^\theta_{oksidasie}$ or/*of* $E^\theta_{cell} = E^\theta_{oxidisingagent} - E^\theta_{reducingagent}$ / $E^\theta_{sel} = E^\theta_{oksideermiddel} - E^\theta_{reduseermiddel}$	

TABLE 3: THE PERIODIC TABLE OF ELEMENTS
TABEL 3: DIE PERIODIEKE TABEL VAN ELEMENTE



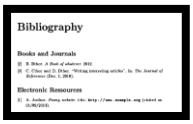
NSC

TABLE 4B: STANDARD REDUCTION POTENTIALS
TABEL 4B: STANDAARD-REDUKSIEPOTENSIALE

Increasing strength of oxidising agents/*Toenemende sterkte van oksideermiddels* ↓

Half-reactions/*Halfreaksies*			E^{θ} (V)
$Li^+ + e^-$	⇌	Li	− 3,05
$K^+ + e^-$	⇌	K	− 2,93
$Cs^+ + e^-$	⇌	Cs	− 2,92
$Ba^{2+} + 2e^-$	⇌	Ba	− 2,90
$Sr^{2+} + 2e^-$	⇌	Sr	− 2,89
$Ca^{2+} + 2e^-$	⇌	Ca	− 2,87
$Na^+ + e^-$	⇌	Na	− 2,71
$Mg^{2+} + 2e^-$	⇌	Mg	− 2,36
$Al^{3+} + 3e^-$	⇌	Al	− 1,66
$Mn^{2+} + 2e^-$	⇌	Mn	− 1,18
$Cr^{2+} + 2e^-$	⇌	Cr	− 0,91
$2H_2O + 2e^-$	⇌	$H_2(g) + 2OH^-$	− 0,83
$Zn^{2+} + 2e^-$	⇌	Zn	− 0,76
$Cr^{3+} + 3e^-$	⇌	Cr	− 0,74
$Fe^{2+} + 2e^-$	⇌	Fe	− 0,44
$Cr^{3+} + e^-$	⇌	Cr^{2+}	− 0,41
$Cd^{2+} + 2e^-$	⇌	Cd	− 0,40
$Co^{2+} + 2e^-$	⇌	Co	− 0,28
$Ni^{2+} + 2e^-$	⇌	Ni	− 0,27
$Sn^{2+} + 2e^-$	⇌	Sn	− 0,14
$Pb^{2+} + 2e^-$	⇌	Pb	− 0,13
$Fe^{3+} + 3e^-$	⇌	Fe	− 0,06
$2H^+ + 2e^-$	⇌	**$H_2(g)$**	**0,00**
$S + 2H^+ + 2e^-$	⇌	$H_2S(g)$	+ 0,14
$Sn^{4+} + 2e^-$	⇌	Sn^{2+}	+ 0,15
$Cu^{2+} + e^-$	⇌	Cu^+	+ 0,16
$SO_4^{2-} + 4H^+ + 2e^-$	⇌	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + 2e^-$	⇌	Cu	+ 0,34
$2H_2O + O_2 + 4e^-$	⇌	$4OH^-$	+ 0,40
$SO_2 + 4H^+ + 4e^-$	⇌	$S + 2H_2O$	+ 0,45
$Cu^+ + e^-$	⇌	Cu	+ 0,52
$I_2 + 2e^-$	⇌	$2I^-$	+ 0,54
$O_2(g) + 2H^+ + 2e^-$	⇌	H_2O_2	+ 0,68
$Fe^{3+} + e^-$	⇌	Fe^{2+}	+ 0,77
$NO_3^- + 2H^+ + e^-$	⇌	$NO_2(g) + H_2O$	+ 0,80
$Ag^+ + e^-$	⇌	Ag	+ 0,80
$Hg^{2+} + 2e^-$	⇌	$Hg(\ell)$	+ 0,85
$NO_3^- + 4H^+ + 3e^-$	⇌	$NO(g) + 2H_2O$	+ 0,96
$Br_2(\ell) + 2e^-$	⇌	$2Br^-$	+ 1,07
$Pt^{2+} + 2e^-$	⇌	Pt	+ 1,20
$MnO_2 + 4H^+ + 2e^-$	⇌	$Mn^{2+} + 2H_2O$	+ 1,23
$O_2(g) + 4H^+ + 4e^-$	⇌	$2H_2O$	+ 1,23
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	⇌	$2Cr^{3+} + 7H_2O$	+ 1,33
$Cl_2(g) + 2e^-$	⇌	$2Cl^-$	+ 1,36
$MnO_4^- + 8H^+ + 5e^-$	⇌	$Mn^{2+} + 4H_2O$	+ 1,51
$H_2O_2 + 2H^+ + 2e^-$	⇌	$2H_2O$	+1,77
$Co^{3+} + e^-$	⇌	Co^{2+}	+ 1,81
$F_2(g) + 2e^-$	⇌	$2F^-$	+ 2,87

Increasing strength of reducing agents/*Toenemende sterkte van reduseermiddels* ↑



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