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SUBJECT: PHYSICAL SCIENCES

TOPIC: ELECTROCHEMICAL REACTIONS

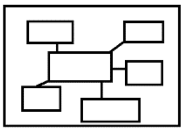
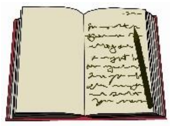
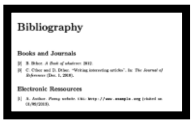
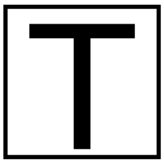
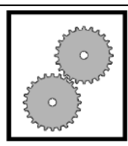
TEACHER AND LEARNER CONTENT MANUAL

Galvanic Cells

Electrolytic Cells

CONTENTS PAGE

ICON DESCRIPTION

 MIND MAP	 EXAMINATION GUIDELINE	 CONTENTS	 ACTIVITIES
 BIBLIOGRAPHY	 TERMINOLOGY	 WORKED EXAMPLES	 STEPS

EXAMINATION GUIDELINES

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

IMPORTANT TERMS AND DEFINITIONS

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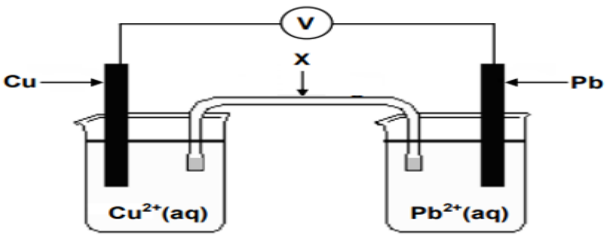
CHEMICAL CHANGE: ELECTROCHEMICAL REACTIONS	
Galvanic cell	A cell in which chemical energy is converted into electrical energy. A galvanic (voltaic) cell has self-sustaining electrode reactions.
Electrolytic cell	A cell in which electrical energy is converted into chemical energy.
Redox reaction	A reaction 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.
Cell notation	A short way to represent a galvanic cell. When writing cell notation, the following convention should be used: o The H_2|H^+ half-cell is treated just like any other half-cell. o Cell terminals (electrodes) are written on the outside of the cell notation. o Active electrodes: reducing agent | oxidised species || oxidising agent | reduced species o 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

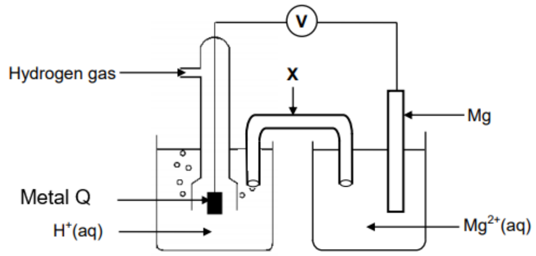
Overall cell reaction	The reaction obtained by combining two half-reactions.
Positive value of the standard emf	The reaction is spontaneous under standard conditions.
Standard conditions for a galvanic cell	Temperature: 25 °C / 298 K Concentration: 1 $mol \cdot dm^{-3}$ Pressure (gases only): 101,3 kPa / 1 atmosphere
Standard hydrogen electrode	The reference electrode used to compile the Table of Standard Reduction Potentials. The hydrogen half-cell was given a standard reduction potential of 0 V. Half-cell notation: Pt | H_2(g) | H^+(aq) Half-reaction: $2H^+ + 2e^- \rightleftharpoons H_2$
Electroplating	The covering of an object with a metal by making it the cathode in an electrolytic cell.

GALVANIC CELLS

Galvanic cells OR Voltaic cells	Define **the galvanic cell** as a cell in which **chemical energy is converted into electrical energy.** **Type of electrochemical cell** in which chemical energy is converted to electrical energy.
	Spontaneous redox reaction takes place. The reaction is **exothermic** i.e. energy is released.

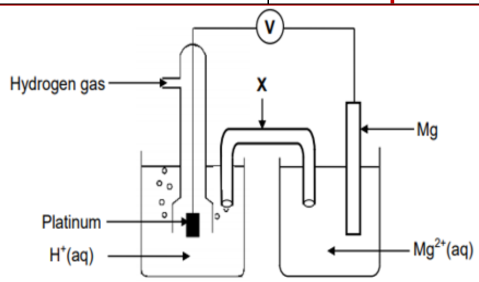
The electrochemical cell illustrated below is set up under standard conditions.





1. State the energy conversion that takes place in this cell.	
Chemical energy is converted to electrical energy.	**Chemical energy is converted to electrical energy.**
2. State the TWO standard conditions that are applicable to the Pb2+|Pb half-cell.	3. Write down THREE standard conditions needed for the hydrogen half-cell to function.
- **Concentration of electrolytes: 1 $mol.dm^{-3}$** - **Temperature: 25 °C / 298 K.**	- **Concentration of electrolytes: 1 $mol. dm^{-3}$** - **Temperature: 25 °C / 298 K** - **Pressure: 1 atm / 101,3 kPa / 1,013 x 105 Pa**
4. Does the above electrochemical cell represent an Electrolytic or Galvanic cell. Give a reason for your answer.	
Galvanic/Voltaic cell, **Converts chemical energy to electrical energy.**	**Galvanic/Voltaic cell,** **Converts chemical energy to electrical energy.**
5. Name the type of electrochemical cell that converts chemical energy to electrical energy.	
Galvanic/voltaic cell.	**Galvanic/voltaic cell.**
6. Write down the name of the item of apparatus labelled. 6.1. X 6.2. V 6.3. Metal Q	
4.1. X – **Salt Bridge** 4.2. V - **Volmeter**	4.1. X – **Salt Bridge** 4.2. V – **Volmeter** 4.3. Metal Q **- Platimun**
7. What is the function of component **X**?	
Completes the circuit. **Provides path for movement of ions. / Ensures electrical neutrality in the cell./Restores charge balance.** √	**Completes the circuit.** **Provides path for movement of ions. / Ensures electrical neutrality in the cell./Restores charge balance.** √
8. What will be the reading on the voltmeter if the salt bridge is removed. Give a reason for this observation.	
There will be no reading on the voltmeter. **Incomplete circuit.**	**There will be no reading on the voltmeter.** **Incomplete circuit.**

9. Write down the voltmeter reading when the net cell reaction in the above electrochemical cell reaches equilibrium.	
0V	**0V**
10. The reading on the voltmeter becomes ZERO after this cell operates for several hours. Give a reason for this reading by referring to the rates of oxidation and reduction half-reactions taking place in the cell/ Give a reason by referring to the cell reaction.	
Reaction reached equilibrium. /(In each half cell) the rate of oxidation is equal to rate of reduction./Rate of the forward reaction is equal to the rate of the reverse reaction. √ **Decrease in V_{cell} as [product ions] increases and [reactant ions] decreases. The rate of forward reaction decreases while rate of reverse reaction increase. V_{cell} = 0 V and no current flows when dynamic equilibrium is reached, i.e., when rate of forward reaction = rate of reverse reaction.**	**Reaction reached equilibrium. /(In each half cell) the rate of oxidation is equal to rate of reduction./Rate of the forward reaction is equal to the rate of the reverse reaction. √** **Decrease in V_{cell} as [product ions] increases and [reactant ions] decreases. The rate of forward reaction decreases while rate of reverse reaction increase. V_{cell} = 0 V and no current flows when dynamic equilibrium is reached, i.e., when rate of forward reaction = rate of reverse reaction.**
11. The voltmeter is now replaced with a bulb. The cell reactions in the above galvanic cell reaches equilibrium. How will this affect the glowing of the above bulb? Give a reason for your answer.	
The bulb will switch off/ it will no longer glow at all. The cell potential will be 0 V.	**The bulb will switch off/ it will no longer glow at all. The cell potential will be 0 V.**



12. Name the standard reference electrode against which all other electrodes potentials are measured.

Standard hydrogen electrode

13. What is the purpose of a standard hydrogen half-cell?

To determine the standard electrode potential of another half-cell.

14. What is the function of platinum in the cell above?

Provides a surface for the transfer of electrons/Allows conduction. Electrode/Conductor of electrons (in hydrogen half-cell)

15. Give TWO reasons, besides being a solid, why platinum is suitable to be used as electrode in the above cell.

Platinum is inert / does not react with the H^+ ions OR acid.
Platinum is a conductor (of electricity).

16. Give a reason why platinum is used as the electrode in half-cell A.

It is a conductor of electricity/a solid to connect wires to./Pt is inert or unreactive. OR H^+ (aq) and hydrogen gas are not solids and cannot be used as an electrode. √

17. How will the initial voltmeter reading change if the size of the platinum electrode is increased. (Write down only INCREASES, DECREASES, or REMAINS THE SAME)

Remains the same.

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

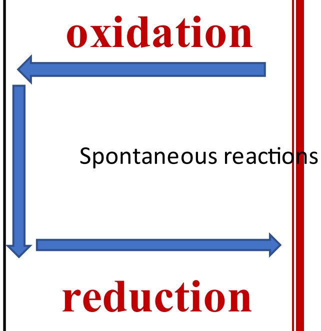
Weak oxidising agents

Strong reducing agents

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
$A\ell^{3+} + 3e^-$	⇌	$A\ell$	– 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
$C\ell_2(g) + 2e^-$	⇌	$2C\ell^-$	+ 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*



Strong oxidising agents

Weak reducing agents

Define **oxidation and reduction in terms of electron (e-) transfer:**

Oxidation is a loss of electrons. **Reduction is a gain of electrons.**

Oxidation half-reaction: $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

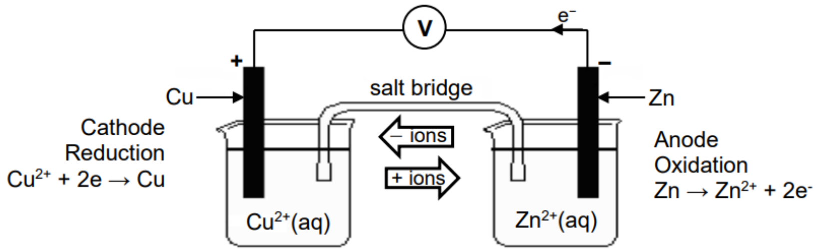
Reduction half-reaction: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Overall redox reaction: $Cu^{2+}(aq) + Zn(s) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Oxidising agent **Reducing agent**

Oxidising agent: A substance that is reduced/gains electrons.

Reducing agent: A substance that is oxidised/loses electrons.

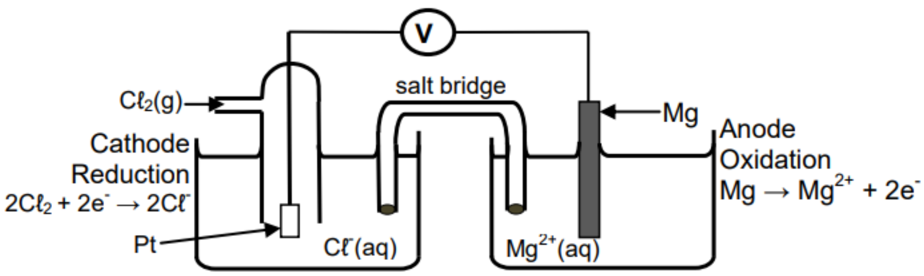


Cell reaction: $Cu^{2+} + Zn \rightarrow Zn^{2+} + Cu$

Cell notation: $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$

Reducing agent | oxidised species || oxidising agent | reduced species

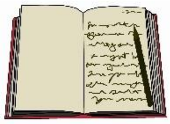
Cell potential: $E^{\theta}_{cell} = E^{\theta}_{cathode} - E^{\theta}_{anode} = 0{,}34 - (-0{,}76) = 1{,}1$ V



Cell reaction: $Mg + C\ell_2 \rightarrow Mg^{2+} + 2C\ell^-$

Cell notation: $Mg(s) \mid Mg^{2+}(aq) \parallel C\ell_2(g) \mid C\ell^-(aq) \mid Pt(s)$

Cell potential: $E°_{cell} = E°_{cathode} - E°_{anode} = (1.36) - (-2.36) = 3.72$ V



Activity 1

1.1. Define the following terms:

1.1.1 Galvanic cell. (2)

1.1.2 Electrolytic cell. (2)

1.1.3. oxidation in terms of electron (e-) transfer. (2)

1.1.4 reduction in terms of electron (e-) transfer. (2)

1.1.5 oxidation in terms of oxidation numbers. (2)

1.1.6 reduction in terms of oxidation numbers. (2)

1.1.7 oxidising agent in terms of oxidation and reduction. (2)

1.1.8 reducing agent in terms of oxidation and reduction. (2)

[16]

ACTIVITY 2: MULTIPLE - CHOICE QUESTIONS

Various options are provided as possible answers to the following questions. Each question has only ONE correct answer. Choose the answer and write only the letter (A–D) next to the question numbers (1.1-1.9) in the ANSWER BOOK, e.g. 1.11 E.

2.1 Which ONE of the following will reduce Sn^{2+} to Sn?

A Zn

B Ag

C Hg

D Pb (2)

2.2 Which ONE of the following metals will reduce $Cd^{2+}(aq)$ to Cd(s), but will NOT reduce $Mn^{2+}(aq)$ to Mn(s)?

A Zn

B Ag

C Ni

D Mg (2)

2.3 Consider the balanced equation for the reaction below:

$$2Cr^{2+}(aq) + Sn^{4+}(aq) \rightarrow 2Cr^{3+}(aq) + Sn^{2+}(aq)$$

The OXIDISING AGENT is:

A $Cr^{2+}(aq)$

B $Cr^{3+}(aq)$

C $Sn^{2+}(aq)$

D $Sn^{4+}(aq)$ (2)

2.4 Which of the following correctly gives the direction, as well as the medium, in which electrons move in a galvanic cell?

	DIRECTION	MEDIUM
A	cathode to anode	salt bridge
B	anode to cathode	external wire
C	cathode to anode	external wire
D	anode to cathode	salt bridge

(2)

2.5 The function of a salt bridge in a galvanic cell is to …

A allow for the movements of protons.

B allow for the movements of electrons.

C provide a site for reduction to occur.

D ensure electrical neutrality of solutions. (2)

2.6 Which ONE of the following statements is TRUE for an oxidising agent?

A It gains electrons.

B It causes another species in the reaction to be reduced.

C Its oxidation number does not change during a chemical reaction.

D Its oxidation number increases during a chemical reaction. (2)

(2)

2.7 The standard zinc half-cell is given below.

$$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$$

Which ONE of the following is TRUE about this half-cell?

A Zn is always a cathode.

B Zn will not be oxidised spontaneously.

C Zn is the anode when the half-cell is connected to the hydrogen half-cell.

D Zn does not lose electrons as easily as hydrogen does. (2)

2.8 Consider the half-reactions and their respective reduction potentials shown below.

Reduction half-reaction	**E^θ (V)**
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+ 1,51
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+ 1,33
$Fe^{3+} + e \rightleftharpoons Fe^{2+}$	+ 0,77
$Cr^{3+} + e^- \rightleftharpoons Cr^{2+}$	- 0,41
$Zn^{2+} + e^- \rightleftharpoons Zn$	- 0,76

Which ONE of the following statements is correct?

A Fe^{3+} can reduce acidified MnO_4^- to Mn^{2+}.

B Zn can oxidise acidified Cr^{2+} to Cr^{3+}.

C Zn^{2+} can reduce acidified $Cr_2O_7^{2-}$ to Cr^{2+}.

D Cr^{2+} can reduce Fe^{3+} to Fe^{2+}. (2)

2.9 Which ONE of the following reactions, when used in a voltaic cell, will give a positive reading on the voltmeter?

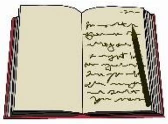
A $Mg^{2+}(aq) + Zn(s) \rightarrow Mg(s) + Zn^{2+}(aq)$

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

C $Co^{2+}(aq) + Sn^{2+}(aq) \rightarrow Co(s) + Sn^{4+}(aq)$

D $3Ni^{2+}(aq) + 2Fe(s) \rightarrow 3Ni(s) + 2Fe^{3+}(aq)$ (2)

[18]



ACTIVITY 3

Consider a GALVANIC cell based on the following half-reactions:

$$Ni^{2+(}aq) + 2e^- \rightleftharpoons Ni(s)$$

$$Cd^{2+} + 2e^- \rightleftharpoons Cd(s)$$

3.1 Define the term *galvanic cell*. (2)

3.2 Use the equations for the half-reactions to write a balanced *net ionic equation* for the overall cell reaction. (3)

3.3 Construct the **cell notation** for this cell. (3)

3.4 State the polarity (NEGATIVE or POSITIVE) of the Cd electrode. (1)

3.5 Calculate the value of the emf of this cell under standard conditions. (4)
[13]

ACTIVITY 4

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

$$\mathbf{X} \mid Zn^{2+}(aq) \parallel Fe^{3+}(aq) , Fe^{2+}(aq) \mid \mathbf{Y}$$

4.1 Write down the NAME or FORMULA of:

4.1.1 Electrode **X** (1)

4.1.2 Electrode **Y** (1)

4.1.3 The oxidising agent (1)

4.2 Write down:

4.2.1 ONE function of electrode **Y** (1)

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

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

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

4.4 How will the initial emf of the cell be affected when the concentration of the iron(III) ions is changed to 0,6 $mol \cdot dm^{-3}$? Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)
[14]

ACTIVITY 5

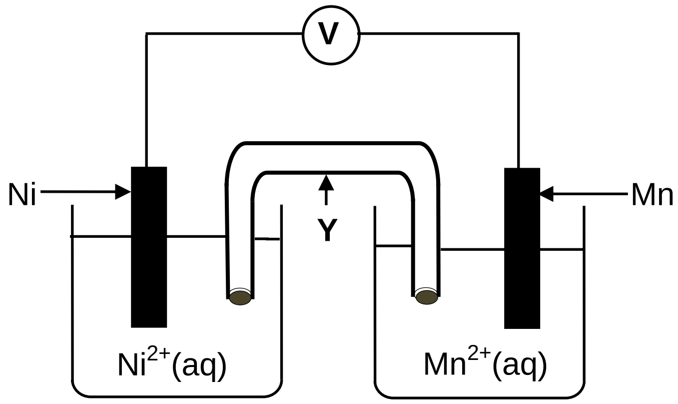
5.1 A piece of zinc (Zn) is placed in a test tube containing an acidified permanganate solution, $MnO_4^-(aq)$. After some time, it is found that a redox reaction has taken place.

Use the Table of Standard Reduction Potentials to answer the following questions:

5.1.1 Write down the NAME or FORMULA of the reducing agent. (1)

5.1.2 Refer to the relative strengths of the OXIDISING AGENTS to explain why a redox reaction has taken place. (3)

5.2 A standard electrochemical cell is set up as shown below.



5.2.1 Write down the function of component **Y**. (1)

5.2.2 In which direction will electrons flow in the external circuit? Choose from 'Ni to Mn' OR 'Mn to Ni'. (2)

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

5.2.4 Write down the balanced equation for the net cell reaction taking place. (3)

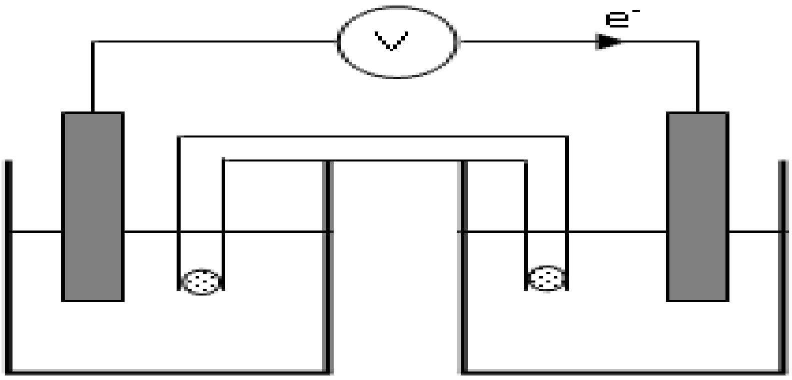
5.2.5 The concentration of $Ni^{2+}(aq)$ is now increased.

Will the reading on the voltmeter INCREASE, DECREASE or REMAIN THE SAME? (1)

[15]

ACTIVITY 6

The galvanic cell represented in the diagram below consists of a Ba electrode dipped into a $Ba(NO_3)_2$ solution, and a Cu electrode dipped into a $Cu(NO_3)_2$ solution. Assume that the cell operates under standard conditions.



Half-cell **A** Half-cell **B**

6.1 State TWO standard conditions under which this cell operates. (2)

6.2 Which half-cell, **A** or **B** is the cathode? Write only **A** or **B**. (1)

6.3 Write down the half-reaction that takes place in half-cell **A**. (2)

6.4 Write down the cell notation for this cell. (3)

6.5 Calculate the emf of this cell. (4)

6.6 How will each of the following changes influence the value of the cell's emf, as calculated in QUESTION 5.5? Write down only INCREASES, DECREASES or REMAINS THE SAME.

6.6.1 Ammonium sulfate is added to the barium nitrate solution. (1)

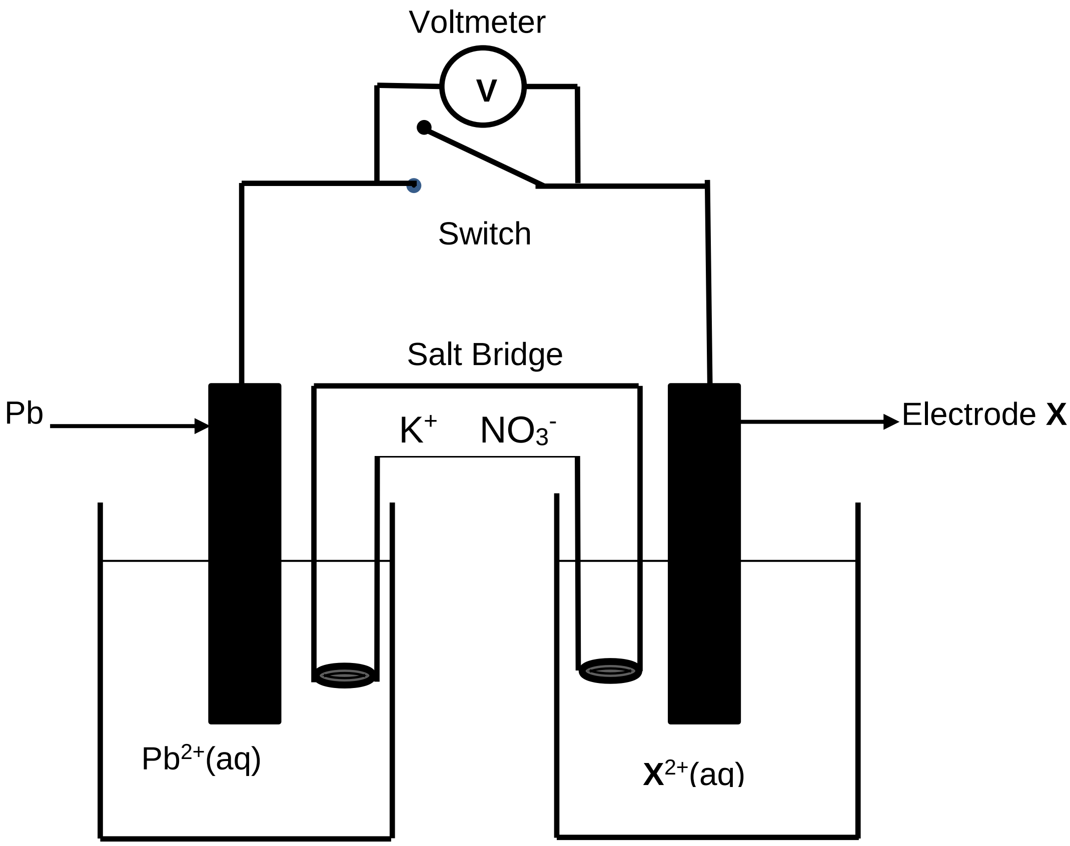
6.6.2 The temperature of the solutions is increased. (1)

[14]

ACTIVITY 7

A standard electrochemical cell is set up using a standard lead half-cell and $X|X^{2+}$ half -cell as shown in the diagram below. A voltmeter connected across the cell, initially registers 0,47 V.

7.1 Define the term *oxidising agent* in terms of ELECTRON TRANSFER. (2)



When the cell is in operation, electrons flow through the Pb electrode towards the **X** electrode in the external circuit.

7.2 Write down the equation for the half reaction that occurs at the cathode. (2)

7.3 Use the STANDARD ELECTRODE POTENTIAL TABLE to identify metal **X**. (5)

7.4 Write down the cell notation of the above cell. (3)

During an experiment, a student set up the electrochemical cell as shown above. After the experiment is over, a student left the switch closed. On the next day, the student opens the switch and takes the voltmeter reading.

7.5 What will be the possible voltmeter reading? Choose from LESS THAN 0,47 V, EQUAL TO 0,47 V or MORE THAN 0,47 V.

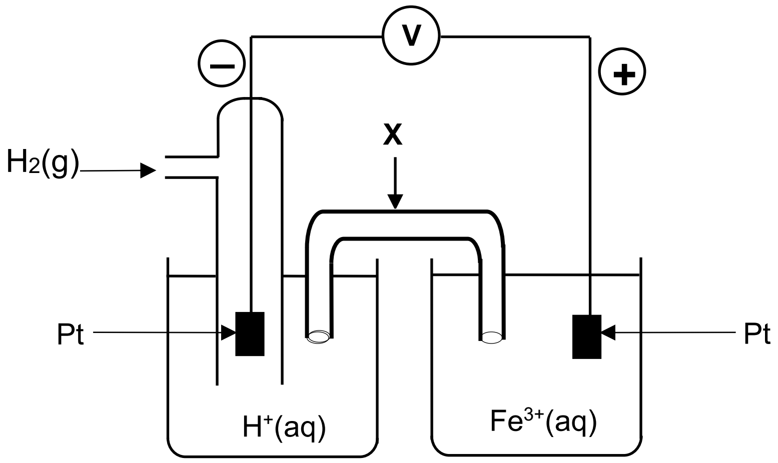
Explain your answer by referring to the concentrations of the electrolytes (3)

[15]

ACTIVITY 8

8.1 The diagram below shows a cell that can be used to measure the STANDARD electrode potential for the following half-reaction:

$$Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}(aq)$$



8.1.1 Define the term *reducing agent* in terms of electron transfer. (2)

8.1.2 Write down the NAME or FORMULA of the strongest reducing agent in the above cell. (1)

8.1.3 Component **X** completes the circuit in the cell. State ONE other function of component **X**. (1)

8.1.4 State ONE condition for this cell to operate under standard conditions. (1)

8.1.5 Give a reason why platinum is suitable to function as electrodes in this cell. (1)

8.1.6 Write down the cell notation for the above cell. (3)

8.1.7 Calculate the initial voltmeter reading of this cell under standard conditions. (4)

8.1.8 The reading on the voltmeter becomes ZERO after this cell operates for several hours. Give a reason for this reading by referring to the rates of the oxidation and reduction half-reactions taking place in the cell. (1)

8.2 Is it advisable to store an iron(III) solution in a container made of tin (Sn)? Choose from YES or NO. Fully explain the answer by referring to the relative strengths of reducing agents. (4)

[18]

ACTIVITY 9

A galvanic cell is set up under standard conditions using half-cells **A** and **B** shown below.

Half-cell **A**: $Cu(s)/Cu^{2+}(aq)$ Half-cell **B**: $H_2O(\ell)$ /$O_2(g)$/ $H^+(aq)$

9.1 Define *oxidation* in terms of electron transfer. (2)

9.2 Write down the:

9.2.1 Initial concentration of the $H^+(aq)$ solution in half-cell **B** (1)

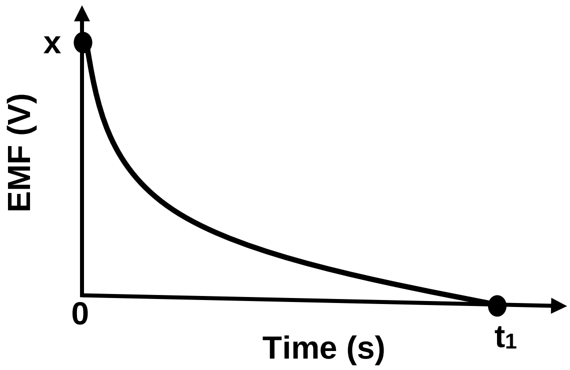
9.2.2 Name of the metal used as the electrode in half-cell **B** (1)

9.2.3 Formula of the reducing agent (1)

9.2.4 Reduction half reaction (2)

9.2.5 Balanced ionic equation for the overall cell reaction (3)

9.3 The graph below shows the EMF of this cell against time.



9.3.1 Calculate the value of **x** on the graph. (4)

9.3.2 Explain the decrease in the EMF of the cell as time proceeds. (2)

9.3.3 What has happened to the reaction in the cell at time $\mathbf{t_1}$? (1)

[17]

ACTIVITY 10

10.1 When a piece of sodium metal (Na) is added to water in a test tube, hydrogen gas is released. When phenolphthalein indicator is added to the test tube, the solution turns pink.

10.1.1 Define the term *reduction* in terms of electron transfer. (2)

10.1.2 Write down the reduction half-reaction. (2)

10.1.3 Write down the balanced equation for the reaction that takes place (3)

10.1.4 Give a reason why the solution turns pink. (1)

When a piece of copper is added to water in a test tube, no reaction is observed.

10.1.5 Refer to the relative strengths of the REDUCING AGENTS to explain why no reaction is observed. (3)

10.2 Consider the cell notation below.

$$Pb(s) \mid Pb^{2+}(aq) \parallel Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt(s)$$

10.2.1 What does the single line (|) in the cell notation above represent? (1)

10.2.2 State the energy conversion that takes place in this cell. (1)

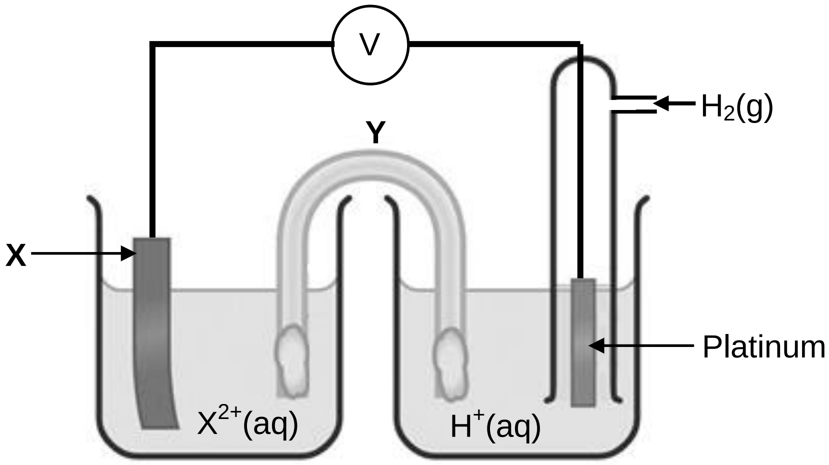
10.2.3. Calculate the initial emf of the cell under standard conditions. (4)

[17]

ACTIVITY 11

Learners want to identify an unknown metal **X** using a standard half-cell, $X|X^{2+}$.

They set up an electrochemical cell under standard conditions using two half-cells, as shown in the diagram below.



The initial emf of this cell is 1,20 V.

11.1 State the standard conditions under which this cell functions. (3)

11.2 State ONE function of component **Y**. (1)

After the cell has operated for some time, it is found that the mass of electrode **X** has increased.

11.3 Identify **X** by means of a suitable calculation. (5)

11.4 Write down the oxidation half-reaction that takes place in this cell. (2)

Half-cell $X|X^{2+}$ is now replaced by an $Au|Au^{3+}$ half-cell.

The initial emf of the cell is now 1,50 V. As the cell operates, the Au electrode increases in mass.

11.5 Arrange the oxidising agents, X^{2+}, Au^{3+} and H^+, in order of increasing strength.

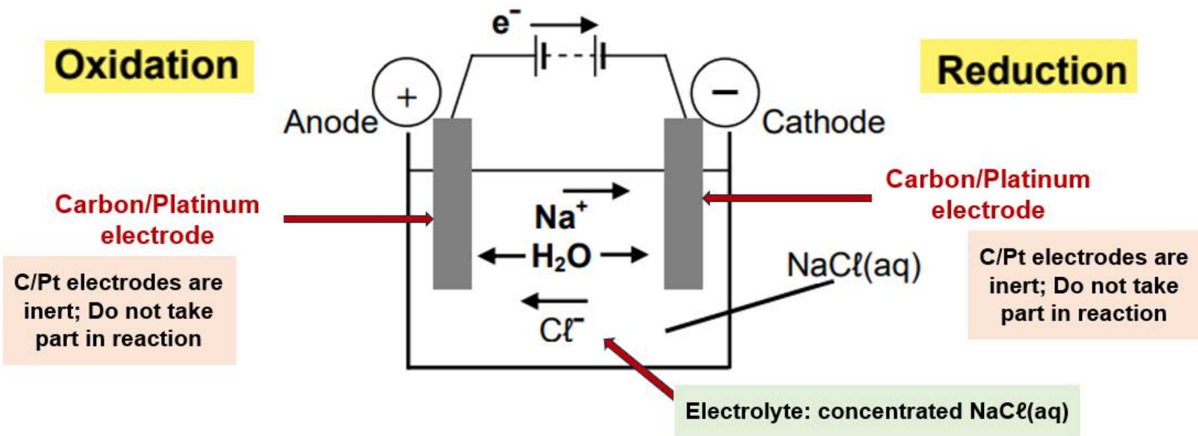
Fully explain the answer. (3)

[14]

ELECTROLYTIC CELLS

Electrolytic cells	Define the **electrolytic cell** as a cell in which **electrical energy is converted into chemical energy**
	Type of electrochemical cell in which electrical energy is converted to chemical energy
	The **process** during which electrical energy is converted into chemical energy is called **'electrolysis'.**
	An electrolytic cell reaction is always a **non-spontaneous**, **endothermic reaction**, which requires a **battery (energy/ power supply).**
	Anode - where **oxidation** takes place – **positive electrode.**
	Cathode - where **reduction** takes place – **negative electrode.**
	The anode and cathode are connected to an external circuit, which is connected to a Direct Current (**DC) power source**.

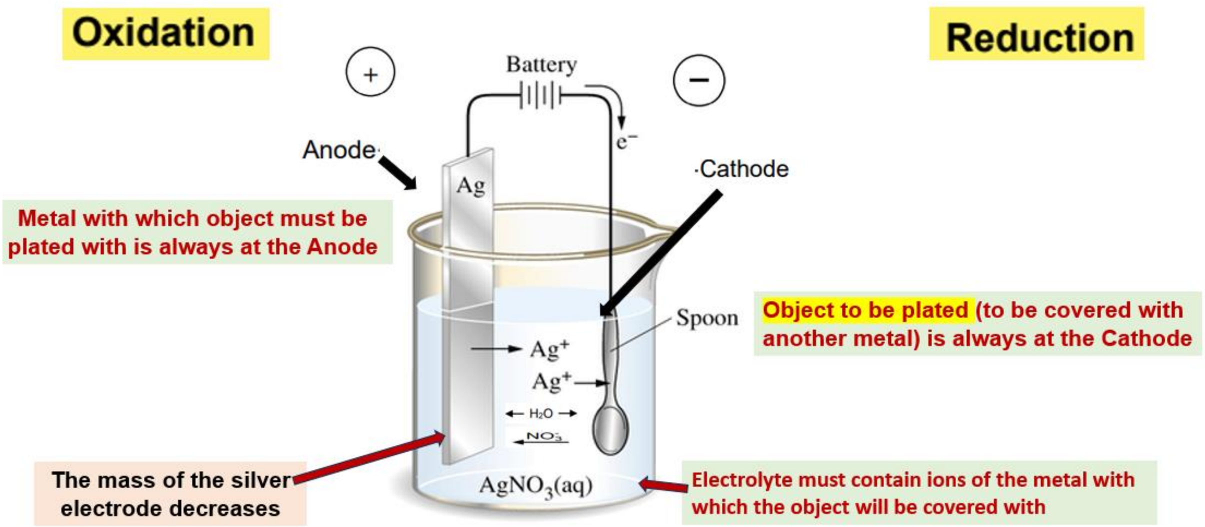
The electrolysis of a concentrated solution of sodium chloride.



<u>**H_2O is a stronger oxidising agent than Na^+ ions and will be reduced at the cathode.**</u>

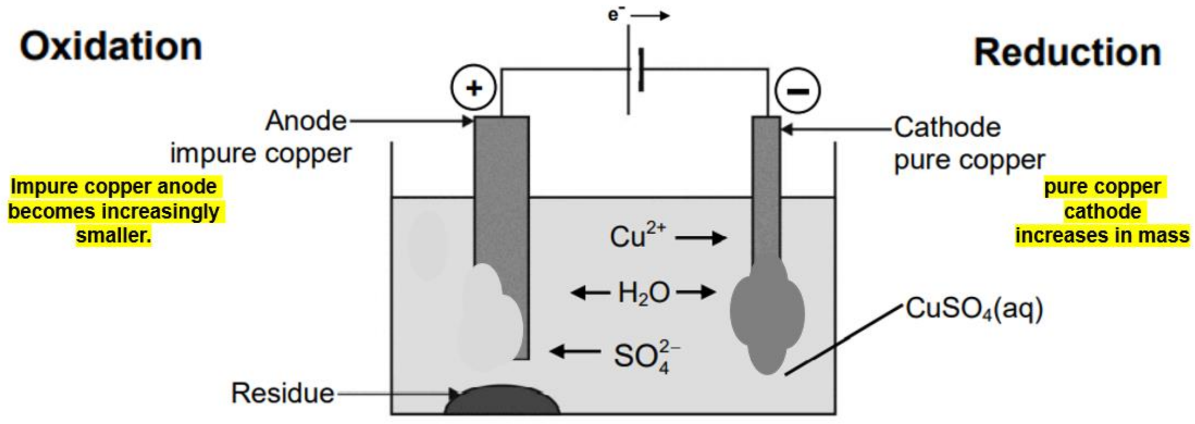
Oxidation half-reaction	$2C\ell^- \rightarrow C\ell_2 + 2e^-$
Reduction half-reaction	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
Overall cell reaction	$2H_2O(\ell) + 2C\ell^-(aq) \rightarrow H_2(g) + 2OH^-(aq) + C\ell_2(g)$

The electroplating of an iron spoon with silver/nickel



Oxidation half-reaction	$Ag \rightarrow Ag^{+} + e^{-}$
Reduction half-reaction	$Ag^{+} + e^{-} \rightarrow Ag$
Overall cell reaction	$Ag^{+}(aq) + Ag(s) \rightarrow Ag(s) + Ag^{+}(aq)$

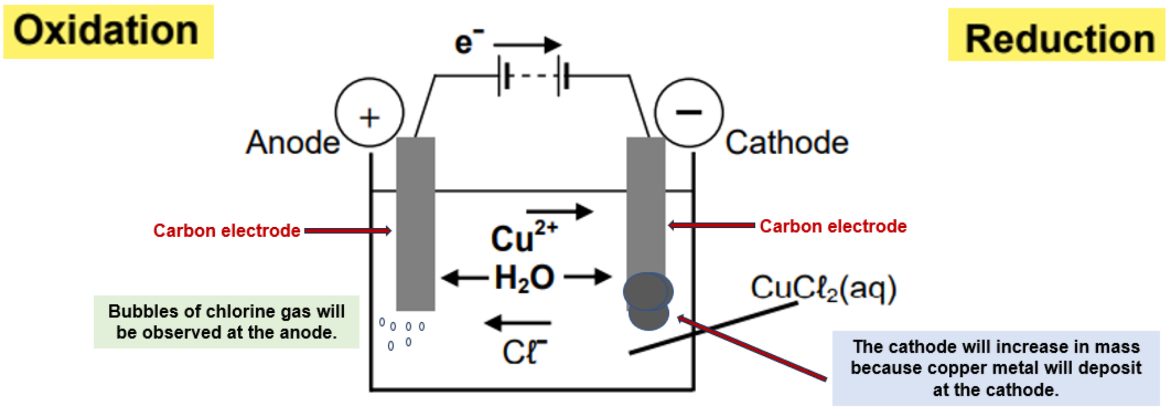
Refining of metals, e.g. copper (Purification of copper)



- The positive Cu^{2+} ions move to cathode (− electrode) and gains electrons: $Cu^{2+} + 2e^{-} \rightarrow Cu$ (reduction).
- The negative $C\ell^{-}$ ions move to the anode (+ electrode), an active Cu electrode. Cu is a stronger reducing agent than $C\ell^{-}$ and therefore Cu will be oxidised to Cu^{2+}: $Cu \rightarrow Cu^{2+} + 2e^{-}$ (oxidation).

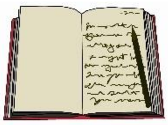
Oxidation half-reaction	$Cu \rightarrow Cu^{2+} + 2e^{-}$
Reduction half-reaction	$Cu^{2+} + 2e^{-} \rightarrow Cu$
Overall cell reaction	$Cu(s) + Cu^{2+}(aq) \rightarrow Cu^{2+}(aq) + Cu(s)$

The decomposition of copper (II) chloride



- The negative $C\ell^-$ ions move to the anode (+ electrode), an inactive carbon electrode, and loses electrons: $2C\ell^- \rightarrow C\ell_2 + 2e^-$ (oxidation).
- The positive Cu^{2+} ions move to cathode (− electrode) and gains electrons: $Cu^{2+} + 2e^- \rightarrow Cu$ (reduction).

Oxidation half-reaction	$2C\ell^- \rightarrow C\ell_2 + 2e^-$
Reduction half-reaction	$Cu^{2+} + 2e^- \rightarrow Cu$
Overall cell reaction	$2C\ell^-(aq) + Cu^{2+}(aq) \rightarrow C\ell_2(g) + Cu(s)$



ACTIVITY 1: MULTIPLE -CHOICE QUESTIONS

Various options are provided as possible answers to the following questions. Each question has only ONE correct answer. Choose the answer and write only the letter (A–D) next to the question numbers (1.1 to 1.6) in the ANSWER BOOK, e.g. 1.11 E.

1.1 A concentrated solution of sodium chloride, $NaC\ell(aq)$, undergoes electrolysis.

Which ONE of the combinations correctly shows the products formed at each electrode?

	CATHODE	ANODE
A	Na	$C\ell_2$
B	H_2	OH^-
C	$C\ell_2$	H_2 and OH^-
D	H_2 and OH^-	$C\ell_2$

(2)

1.2 Which ONE of the following statements is CORRECT for an ELECTROLYTIC CELL?

A The cell reaction is exothermic.

B The cell reaction is endothermic.

C The cell reaction is spontaneous.

D The heat of reaction is negative. (2)

1.3 Which ONE of the following statements is CORRECT for an ELECTROLYTIC CELL?

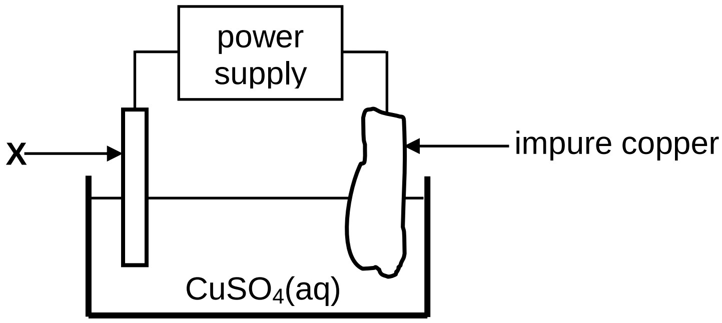
A The anode is the positive electrode.

B The cathode is the positive electrode.

C Oxidation takes place at the cathode.

D Reduction takes place at the anode. (2)

1.4 The diagram below represents a cell that is used for the refining of copper.



Which ONE of the following statements is TRUE?

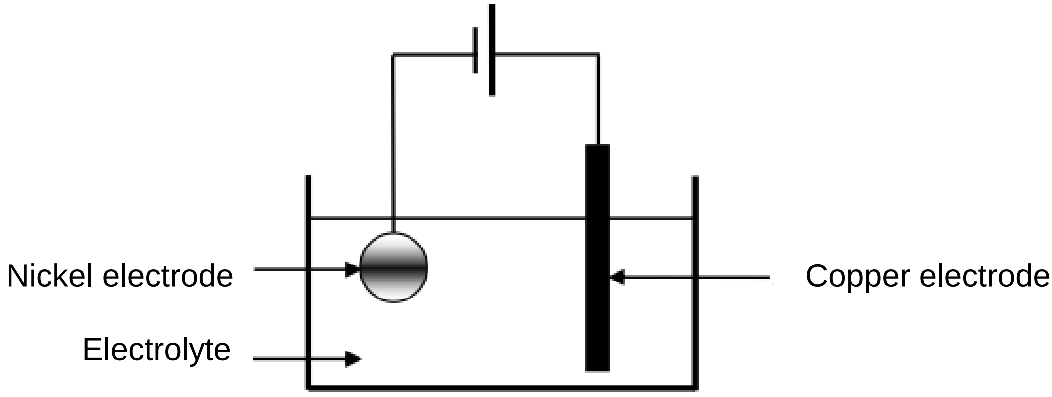
A **X** is made of platinum.

B The mass of **X** increases.

C **X** is the electrode where oxidation takes place.

D **X** is connected to the positive terminal of the power supply. (2)

1.5. The simplified diagram below represents an electrolytic cell used to electroplate a nickel (Ni) coin with copper (Cu).



Which ONE of the following reactions takes place at the anode?

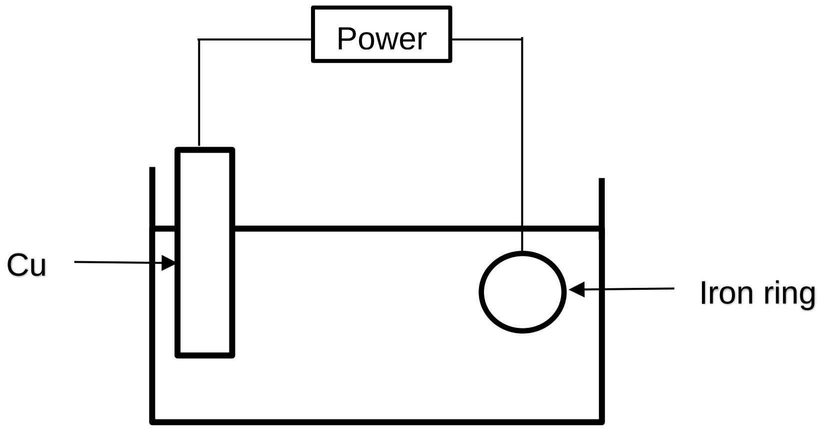
A $Ni^{2+} + 2e^- \rightarrow Ni$

B $Ni \rightarrow Ni^{2+} + 2e^-$

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

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

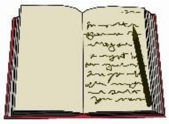
1.6 The electrolytic cell below is used during the electroplating of an iron ring with copper.



Which ONE of the following combinations is CORRECT about the ions in the electrolyte when the cell is operating?

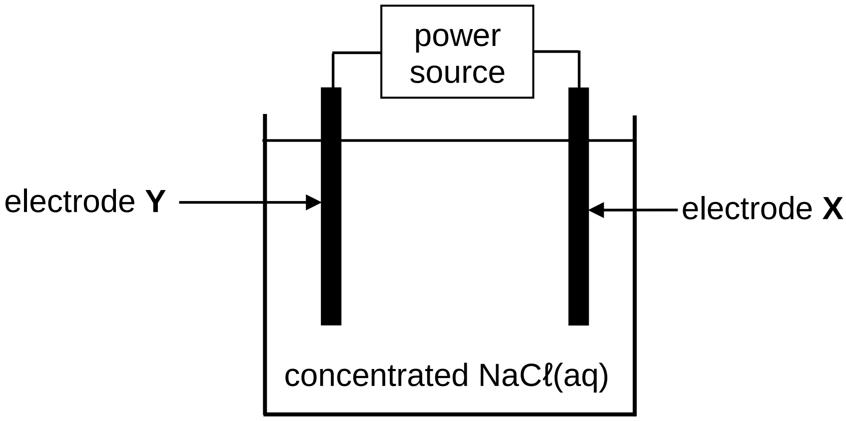
	Concentration	Positive ions
A	Remain constant	Cu^{2+}
B	Remain constant	Fe^{2+}
C	Increases	Fe^{3+}
D	Increases	Cu^{2+}

(2)
[12]



ACTIVITY 2

The simplified diagram below represents an electrochemical cell used for the electrolysis of a concentrated sodium chloride solution, $NaC\ell(aq)$. **X** and **Y** are carbon electrodes.



2.1 Define the term *electrolysis*. (2)

2.2 Chlorine gas, $C\ell_2(g)$, is released at electrode **X**.

Write down the:

2.2.1 Letter (**X** or **Y**) of the electrode where oxidation takes place (1)

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

2.2.3 Direction in which electrons flow in the external circuit

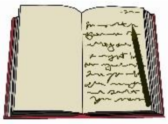
Choose from **X** to **Y** OR **Y** to **X**. (1)

2.2.4 Balanced equation for the net (overall) cell reaction that takes place in the cell (3)

2.3 How will the pH of the electrolyte change during the reaction?

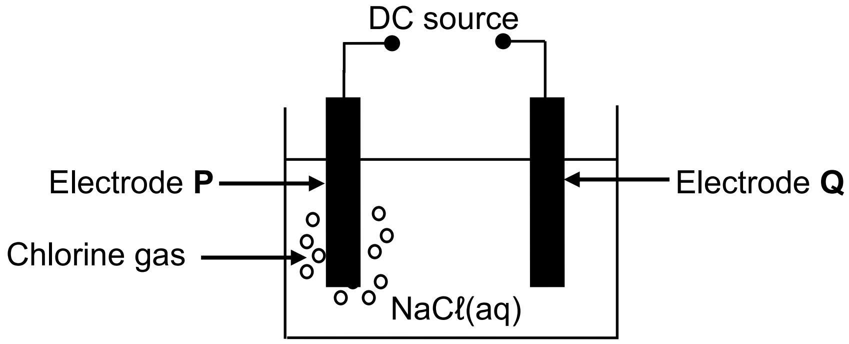
Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

2.4 Give a reason for the answer to QUESTION 2.3. (1)
[11]



ACTIVITY 3

The simplified diagram below represents a cell used for the electrolysis of a concentrated sodium chloride solution, NaCℓ(aq).
.



3.1 Define the term *electrolysis*. (2)

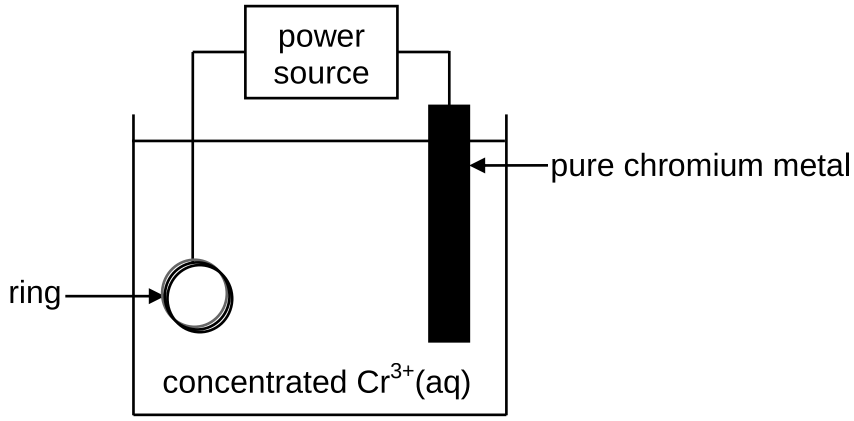
3.2 Which electrode (**P** or **Q**) is positive? (1)

3.3 Write down the half-reaction taking place at electrode **P**. (2)

3.4 Write down the overall (net) cell reaction that takes place in this cell. (3)
[8]

ACTIVITY 4

The diagram below shows a simplified electrolytic cell used to electroplate a ring.



4.1 Define the term *electrolyte*. (2)

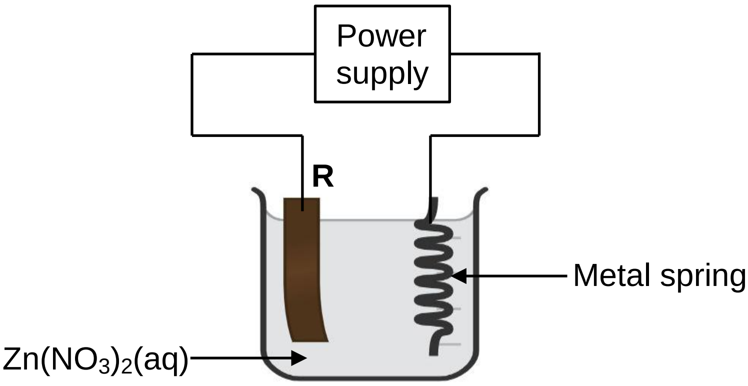
4.2 Is the pure chromium metal the ANODE or the CATHODE of the cell? Give a reason for the answer. (2)

4.3 Write down the half-reaction that takes place at the ring. (2)

4.4 Calculate the total charge transferred when the mass of the pure chromium changes by 2 g. (5)
[11]

ACTIVITY 5

The simplified electrolytic cell below is used to electroplate a metal spring. Zinc nitrate, $Zn(NO_3)_2(aq)$, is used as an electrolyte and **R** is an electrode.



5.1 Define the term *electrolytic cell.* (2)

5.2 Which electrode (**R** or **METAL SPRING**) is the ANODE? Give a reason for the answer. (2)

5.3 Write down the:

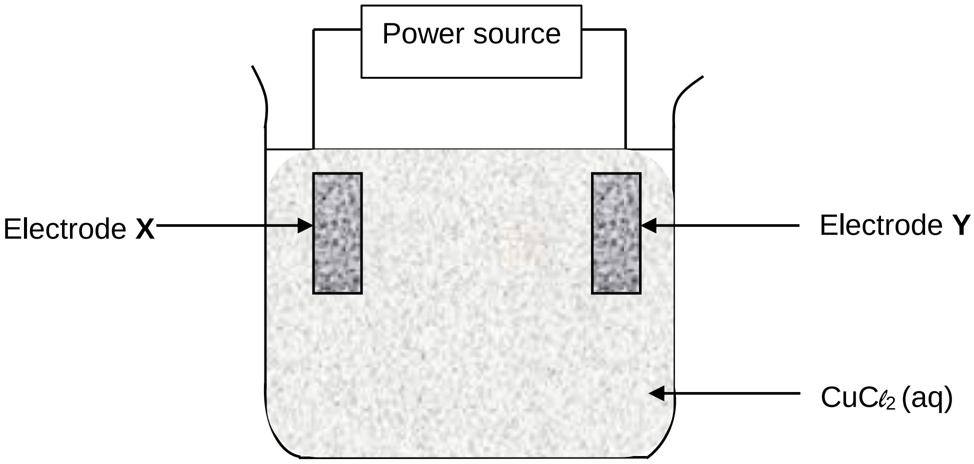
5.3.1 Equation for the half-reaction occurring at the metal spring (2)

5.3.2 NAME or FORMULA of a suitable metal that can be used as electrode **R** (1)

5.4 Explain the answer to QUESTION 5.3.2. (2)
[9]

ACTIVITY 6

The diagram below represents an electrochemical cell that is used for the refining of copper.



One of the electrodes consists of *pure copper* and the other of *impure copper*.

6.1 When the cell is allowed to operate, it is found that the mass of electrode **X increases** over time.

6.1.1 Define the term *electrolyte*. (2)

6.1.2 Is electrode **X** the CATHODE or ANODE of the cell? (1)

6.1.3 Write down a half-reaction to justify the answer to QUESTION 6.1.2 above. (2)

6.2 State whether electrode **Y** is connected to the POSITIVE terminal or the NEGATIVE terminal of the power source. (1)

6.3 The impure copper contains *zinc* and *silver*. State what will happen to the:

6.3.1 Silver that is in the impure copper (1)

6.3.2 Zinc that is in the impure copper (1)

6.4 Refer to the *relative strength of oxidising agents* to explain why zinc in the impure copper will not be deposited on the pure copper electrode. (2)

6.5 After the purification of the impure copper was completed, it was found that $3,75 \times 10^{-2}$ mol of copper was formed. Calculate the percentage of copper that was present in the impure sample if the initial mass of the impure copper was 4 g. (4)

Electrodes **X** and **Y** are now replaced with *graphite* (C) electrodes.

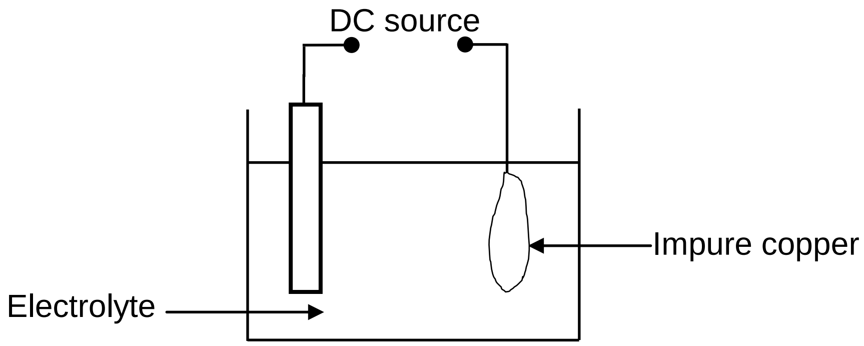
6.6 State what will be observed at electrode **Y** (now graphite). (1)

6.7 Describe what will happen to the blue colour of the electrolyte as the reaction proceeds/progresses. (2)

[17]

ACTIVITY 7

The simplified diagram below represents an electrochemical cell used for the purification of copper. The impure copper contains small amounts of silver (Ag) and zinc (Zn) as the only impurities.



7.1 Define the term *electrolysis*. (2)

7.2 Write down the NAME or FORMULA of TWO positive ions present in the electrolyte. (2)

7.3 Write down the half-reaction that takes place at the cathode. (2)

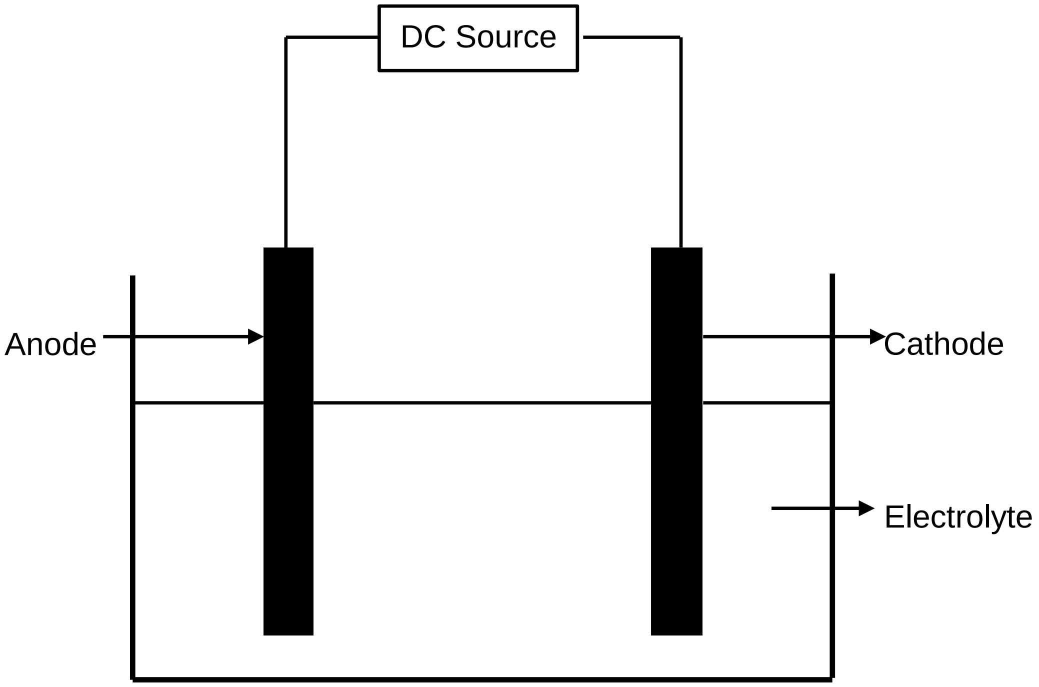
7.4 Refer to the Table of Standard Reduction Potentials and explain why the purified copper will NOT contain any zinc. (3)

7.5 Calculate the maximum mass of Cu formed if 0,6 moles of electrons are transferred. (3)

[12]

ACTIVITY 8

Copper metal can be purified by electrolysis, using the electrochemical shown below.



8.1 Define the term *electrolysis.* (2)

8.2 Write down the CHEMICAL NAME or FORMULA of the electrolyte. (1)

8.3 On which electrode will copper be formed? Write down only ANODE or CATHODE. Support your answer by writing down the relevant half reaction. (3)

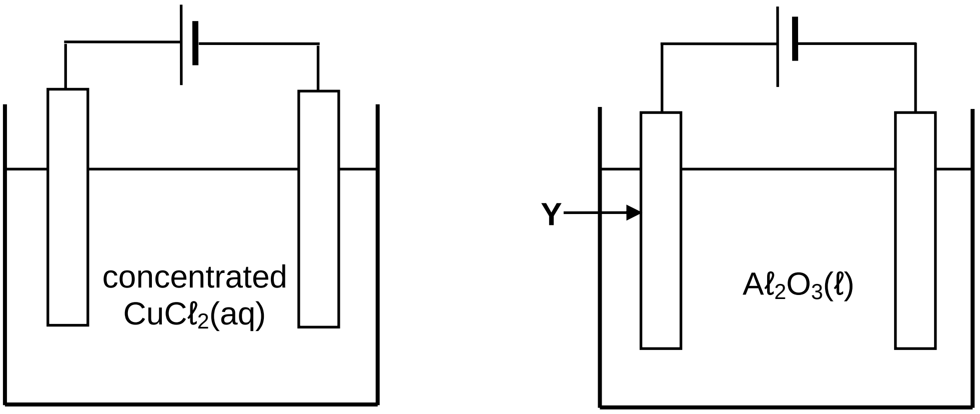
8.4 The solid impurities which form during the electrolysis contain silver.

Refer to the relative strength of reducing agents to explain why silver metal does not react with the electrolyte mentioned in QUESTION 8.2. (3)

[9]

ACTIVITY 9

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**

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

9.2 Write down the:

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

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

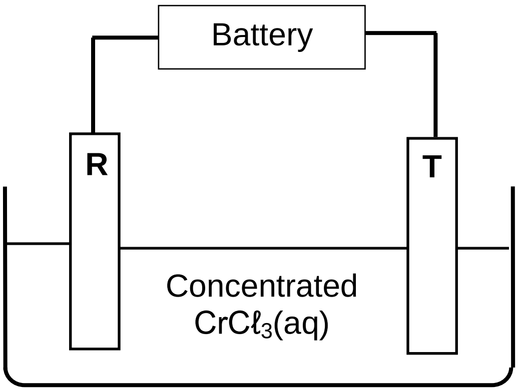
9.2.3 NAME or FORMULA of the product formed at the cathode of cell **A** (1)

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

[8]

ACTIVITY 10

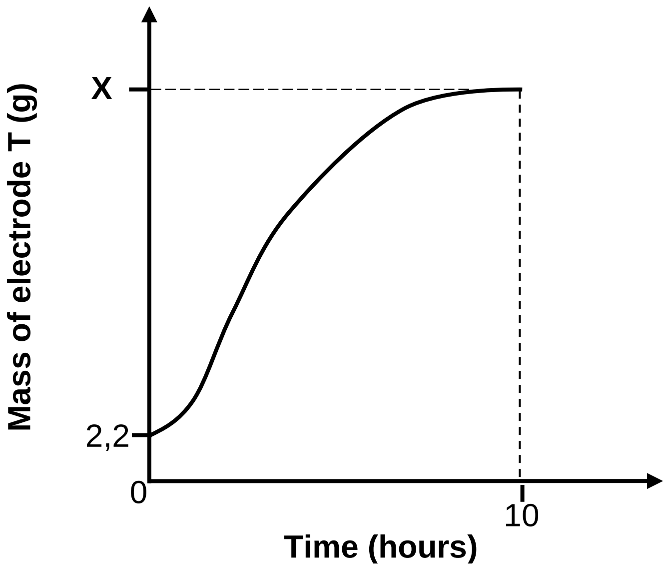
The diagram below represents a simplified cell used for the electrolysis of CONCENTRATED chromium(III) chloride, $CrC\ell_3(aq)$. Electrodes **R** and **T** are made of carbon.



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

10.1 Define the term *electrolysis*. (2)

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



10.2.1 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:

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

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

[13]

DATA FOR PHYSICAL SCIENCES GRADE 12
PAPER 2 (CHEMISTRY)

GEGEWENS VIR FISIESE WETENSKAPPE GRAAD 12
VRAESTEL 2 (CHEMIE)

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* $c = \frac{m}{MV}$	$n = \frac{V}{V_m}$
$\frac{c_aV_a}{c_bV_b} = \frac{n_a}{n_b}$	$pH = -\log[H_3O^+]$
$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}$ 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* $E^{\theta}_{cell} = E^{\theta}_{oxidising agent} - E^{\theta}_{reducing agent}$ / $E^{\theta}_{sel} = E^{\theta}_{oksideermiddel} - E^{\theta}_{reduseermiddel}$	
$I = \frac{Q}{\Delta t}$	$n = \frac{Q}{q_e}$ where n = number of electrons

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

KEY/*SLEUTEL*

Electronegativity / *Elektronegatiwiteit*	Atomic number / *Atoomgetal*	Symbol / *Simbool*	Approximate relative atomic mass / *Benaderde relatiewe atoommassa*
1,9	29	Cu	63,5

1 (I)	2 (II)	3	4	5	6	7	8	9	10	11	12	13 (III)	14 (IV)	15 (V)	16 (VI)	17 (VII)	18 (VIII)
1 2,1 **H** 1																	2 **He** 4
3 1,0 **Li** 7	4 1,5 **Be** 9											5 2,0 **B** 11	6 2,5 **C** 12	7 3,0 **N** 14	8 3,5 **O** 16	9 4,0 **F** 19	10 **Ne** 20
11 0,9 **Na** 23	12 1,2 **Mg** 24											13 1,5 **Aℓ** 27	14 1,8 **Si** 28	15 2,1 **P** 31	16 2,5 **S** 32	17 3,0 **Cℓ** 35,5	18 **Ar** 40
19 0,8 **K** 39	20 1,0 **Ca** 40	21 1,3 **Sc** 45	22 1,5 **Ti** 48	23 1,6 **V** 51	24 1,6 **Cr** 52	25 1,5 **Mn** 55	26 1,8 **Fe** 56	27 1,8 **Co** 59	28 1,8 **Ni** 59	29 1,9 **Cu** 63,5	30 1,6 **Zn** 65	31 1,6 **Ga** 70	32 1,8 **Ge** 73	33 2,0 **As** 75	34 2,4 **Se** 79	35 2,8 **Br** 80	36 **Kr** 84
37 0,8 **Rb** 86	38 1,0 **Sr** 88	39 1,2 **Y** 89	40 1,4 **Zr** 91	41 **Nb** 92	42 1,8 **Mo** 96	43 1,9 **Tc**	44 2,2 **Ru** 101	45 2,2 **Rh** 103	46 2,2 **Pd** 106	47 1,9 **Ag** 108	48 1,7 **Cd** 112	49 1,7 **In** 115	50 1,8 **Sn** 119	51 1,9 **Sb** 122	52 2,1 **Te** 128	53 2,5 **I** 127	54 **Xe** 131
55 0,7 **Cs** 133	56 0,9 **Ba** 137	57 **La** 139	72 1,6 **Hf** 179	73 **Ta** 181	74 **W** 184	75 **Re** 186	76 **Os** 190	77 **Ir** 192	78 **Pt** 195	79 **Au** 197	80 **Hg** 201	81 1,8 **Tℓ** 204	82 1,8 **Pb** 207	83 1,9 **Bi** 209	84 2,0 **Po**	85 2,5 **At**	86 **Rn**
87 0,7 **Fr**	88 0,9 **Ra** 226	89 **Ac**															

58 **Ce** 140	59 **Pr** 141	60 **Nd** 144	61 **Pm**	62 **Sm** 150	63 **Eu** 152	64 **Gd** 157	65 **Tb** 159	66 **Dy** 163	67 **Ho** 165	68 **Er** 167	69 **Tm** 169	70 **Yb** 173	71 **Lu** 175
90 **Th** 232	91 **Pa**	92 **U** 238	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

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

Increasing strength of oxidising agents /*Toenemende sterkte van oksideermiddels* (↑)

Half-reactions/*Halfreaksies*			E^θ (V)
$F_2(g) + 2e^-$	⇌	$2F^-$	+ 2,87
$Co^{3+} + e^-$	⇌	Co^{2+}	+ 1,81
$H_2O_2 + 2H^+ + 2e^-$	⇌	$2H_2O$	+1,77
$MnO_4^- + 8H^+ + 5e^-$	⇌	$Mn^{2+} + 4H_2O$	+ 1,51
$C\ell_2(g) + 2e^-$	⇌	$2C\ell^-$	+ 1,36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	⇌	$2Cr^{3+} + 7H_2O$	+ 1,33
$O_2(g) + 4H^+ + 4e^-$	⇌	$2H_2O$	+ 1,23
$MnO_2 + 4H^+ + 2e^-$	⇌	$Mn^{2+} + 2H_2O$	+ 1,23
$Pt^{2+} + 2e^-$	⇌	Pt	+ 1,20
$Br_2(\ell) + 2e^-$	⇌	$2Br^-$	+ 1,07
$NO_3^- + 4H^+ + 3e^-$	⇌	$NO(g) + 2H_2O$	+ 0,96
$Hg^{2+} + 2e^-$	⇌	$Hg(\ell)$	+ 0,85
$Ag^+ + e^-$	⇌	Ag	+ 0,80
$NO_3^- + 2H^+ + e^-$	⇌	$NO_2(g) + H_2O$	+ 0,80
$Fe^{3+} + e^-$	⇌	Fe^{2+}	+ 0,77
$O_2(g) + 2H^+ + 2e^-$	⇌	H_2O_2	+ 0,68
$I_2 + 2e^-$	⇌	$2I^-$	+ 0,54
$Cu^+ + e^-$	⇌	Cu	+ 0,52
$SO_2 + 4H^+ + 4e^-$	⇌	$S + 2H_2O$	+ 0,45
$2H_2O + O_2 + 4e^-$	⇌	$4OH^-$	+ 0,40
$Cu^{2+} + 2e^-$	⇌	Cu	+ 0,34
$SO_4^{2-} + 4H^+ + 2e^-$	⇌	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + e^-$	⇌	Cu^+	+ 0,16
$Sn^{4+} + 2e^-$	⇌	Sn^{2+}	+ 0,15
$S + 2H^+ + 2e^-$	⇌	$H_2S(g)$	+ 0,14
$\mathbf{2H^+ + 2e^-}$	**⇌**	$\mathbf{H_2(g)}$	**0,00**
$Fe^{3+} + 3e^-$	⇌	Fe	– 0,06
$Pb^{2+} + 2e^-$	⇌	Pb	– 0,13
$Sn^{2+} + 2e^-$	⇌	Sn	– 0,14
$Ni^{2+} + 2e^-$	⇌	Ni	– 0,27
$Co^{2+} + 2e^-$	⇌	Co	– 0,28
$Cd^{2+} + 2e^-$	⇌	Cd	– 0,40
$Cr^{3+} + e^-$	⇌	Cr^{2+}	– 0,41
$Fe^{2+} + 2e^-$	⇌	Fe	– 0,44
$Cr^{3+} + 3e^-$	⇌	Cr	– 0,74
$Zn^{2+} + 2e^-$	⇌	Zn	– 0,76
$2H_2O + 2e^-$	⇌	$H_2(g) + 2OH^-$	– 0,83
$Cr^{2+} + 2e^-$	⇌	Cr	– 0,91
$Mn^{2+} + 2e^-$	⇌	Mn	– 1,18
$A\ell^{3+} + 3e^-$	⇌	$A\ell$	– 1,66
$Mg^{2+} + 2e^-$	⇌	Mg	– 2,36
$Na^+ + e^-$	⇌	Na	– 2,71
$Ca^{2+} + 2e^-$	⇌	Ca	– 2,87
$Sr^{2+} + 2e^-$	⇌	Sr	– 2,89
$Ba^{2+} + 2e^-$	⇌	Ba	– 2,90
$Cs^+ + e^-$	⇌	Cs	- 2,92
$K^+ + e^-$	⇌	K	– 2,93
$Li^+ + e^-$	⇌	Li	– 3,05

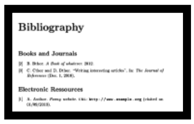
Increasing strength of reducing agents /*Toenemende sterkte van reduseermiddels* (↓)

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

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

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

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
$A\ell^{3+} + 3e^-$	⇌	$A\ell$	− 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
$C\ell_2(g) + 2e^-$	⇌	$2C\ell^-$	+ 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



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