

ELECTROCHEMISTRY

Electrochemistry can be defined as the study of the effects of electricity on a substance/ compound and how chemical reactions produce electricity.

Electrochemistry therefore deals mainly with:

- i) Reduction and oxidation
- ii) Electrochemical (voltaic) cell
- iii) Electrolysis (electrolytic) cell

(i) REDUCTION AND OXIDATION (REDOX)

1. In terms of oxygen transfer:

- i) Reduction is **removal** of oxygen.
- ii) Oxidation is **addition** of oxygen.
- iii) Redox is **simultaneous** addition and removal of oxygen.
- iv) Reducing agent is the species that undergoes **oxidation**, therefore **gains** oxygen.
- v) Oxidizing agent is the species that undergoes **reduction**, therefore **loses/donates** oxygen.

e.g. When hydrogen is passed through heated copper (II) oxide, it is **oxidised** to copper metal as in the equation below:

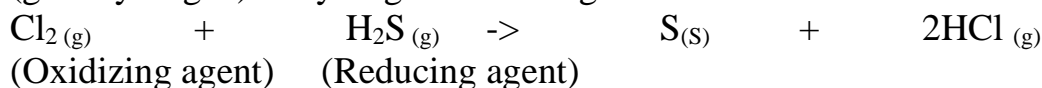


2. In terms of hydrogen transfer:

- i) Oxidation is the **removal** of hydrogen.
- ii) Reduction is the **addition** of hydrogen.
- iii) Redox is **simultaneous** addition and removal of hydrogen.
- iv) Reducing agent is the species that undergoes **oxidation**, therefore **loses/donates** hydrogen.

v) Oxidizing agent is the species that undergoes **reduction**, therefore **gains** hydrogen.

e.g. When hydrogen sulphide gas is bubbled into a gas jar containing chlorine gas it is oxidized (lose the hydrogen) to sulphur (yellow solid). The chlorine is reduced (gain hydrogen) to hydrogen chloride gas.



3. In terms of electron transfer:

i) Oxidation is **donation/ loss/ removal** of electrons.

ii) Reduction is **gain/ accept/ addition** of electrons.

iii) Redox is **simultaneous gain/ accept/ addition** and **donation/ loss/ removal** of electrons.

iv) Reducing agent is the species that undergoes **oxidation**, therefore **looses/ donates** electrons.

v) Oxidizing agent is the species that undergoes **reduction**, therefore **gains/ accepts** electrons.

Example

a) Displacement of metals from their solutions:

Place 5cm³ each of Iron (II) sulphate (VI) solution into three different test tubes. Add about 1g of copper tunings / powder into one test tube then zinc and magnesium powders separately into the other test tubes. Shake thoroughly for 2 minutes each. Record any colour changes in the table below.

Metal added to Iron (II) sulphate (VI) solution	Colour changes
Copper	Solution remains green
Zinc	Green colour fades
Magnesium	Green colour fades

Explanation

-When a more reactive metal is added to a solution of less reactive metal, it displaces it from its solution.

-When a less reactive metal is added to a solution of a more reactive metal, it does not displace it from its solution.

-Copper is less reactive than iron therefore cannot displace iron its solution.

-Zinc is more reactive than iron therefore can displace iron from its solution.

-Magnesium is more reactive than iron therefore can displace iron from its solution.

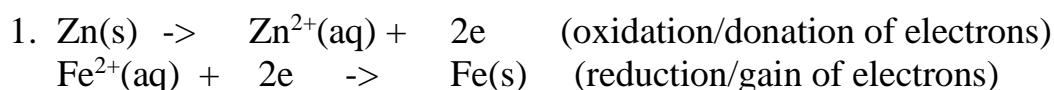
In terms of electron transfer:

- the more reactive metal undergoes oxidation (reducing agent) by **donating/loosing** electrons to form **ions**

-the less reactive metal undergoes reduction (oxidizing agent) by its ions in solution gaining /accepting/acquiring the electrons to form the metal.

-displacement of metals involves therefore electron transfer from a more reactive metal to ions of another less reactive metal.

Examples



Fe²⁺(aq) + Zn(s) -> Zn²⁺(aq) + Fe(s) (redox/both donation and gain of electrons)

2. **Mg(s) -> Mg²⁺(aq) + 2e** (oxidation/donation of electrons)

Fe²⁺(aq) + 2e -> Fe(s) (reduction/gain of electrons)

Fe²⁺(aq) + Mg(s) -> Mg²⁺(aq) + Fe(s) (redox/both donation and gain of electrons)

3. **Zn(s) -> Zn²⁺(aq) + 2e** (oxidation/donation of electrons)

Cu²⁺(aq) + 2e -> Cu(s) (reduction/gain of electrons)

Cu²⁺(aq) + Zn(s) -> Zn²⁺(aq) + Cu(s) (redox/both donation and gain of electrons)

4. **Fe(s) -> Fe²⁺(aq) + 2e** (oxidation/donation of electrons)

2Ag⁺(aq) + 2e -> 2Ag(s) (reduction/gain of electrons)

2Ag⁺(aq) + Fe(s) -> Fe²⁺(aq) + 2Ag(s) (redox/both donation and gain of electrons)

5. **Zn(s) -> Zn²⁺(aq) + 2e** (oxidation/donation of electrons)

Cl₂(g) + 2e -> 2Cl⁻(aq) (reduction/gain of electrons)

Cl₂(g) + Zn(s) -> Zn²⁺(aq) + 2Cl⁻(aq) (redox/both donation and gain of electrons)

6. **2Mg(s) -> 2Mg²⁺(aq) + 4e** (oxidation/donation of electrons)

O₂(g) + 4e -> 2O²⁻(aq) (reduction/gain of electrons)

O₂(g) + 2Mg(s) -> 2Mg²⁺(aq) + 2O²⁻(aq) (redox/both donation and gain of electrons)

Note

(i)The number of electrons donated/lost MUST be equal to the number of electrons gained/acquired.

(i)During displacement reaction, the colour of ions /salts fades but does not if displacement does not take place. e.g

a)Green colour of **Fe²⁺(aq) fades** if Fe²⁺(aq) ions are **displaced** from their solution. Green colour of **Fe²⁺(aq) appear** if Fe/iron **displaces** another salt/ions from their solution.

b)Blue colour of **Cu²⁺(aq) fades** if Cu²⁺(aq) ions are **displaced** from their solution and **brown** copper deposits appear. Blue colour of **Cu²⁺(aq) appear** if Cu/copper **displaces** another salt/ions from their solution.

c)Brown colour of **Fe³⁺(aq) fades** if Fe³⁺(aq) ions are **displaced** from their solution. Brown colour of **Fe³⁺(aq) appear** if Fe/iron **displaces** another salt/ions from their solution to form Fe³⁺(aq).

(iii) Displacement reactions also produce **energy/heat**. The **closer/nearer** the metals in the reactivity/electrochemical series the **less** energy/heat of displacement.

(iv) The **higher** the metal in the reactivity series therefore the **easier** to loose/donate electrons and thus the **stronger** the reducing agent.

4. (a) In terms of oxidation number:

i) Oxidation is increase in oxidation numbers.

ii) Reduction is decrease in oxidation numbers.

iii) Redox is simultaneous increase in oxidation numbers of one species/substance and a decrease in oxidation numbers of another species/substance.

iv) Reducing agent is the species that undergoes oxidation, therefore increases its oxidation number.

v) Oxidizing agent is the species that undergoes reduction, therefore increases its oxidation number.

(b) The idea/concept of oxidation numbers uses/applies the following simple guideline rules:

Guidelines /rules applied in assigning oxidation number

1. Oxidation number of combined Oxygen is always **-2** except in peroxides ($\text{Na}_2\text{O}_2/\text{H}_2\text{O}_2$) where its Oxidation number is **-1**

2. Oxidation number of combined Hydrogen is always **+1** except in Hydrides (NaH/KH) where its Oxidation number is **-1**

3. All **atoms** and **molecules** of elements have oxidation number **0** (zero)

Atom	Oxidation number	Molecule	Oxidation number
Na	0	Cl_2	0
O	0	O_2	0
H	0	H_2	0
Al	0	N_2	0
Ne	0	O_3	0
K	0	P_3	0
Cu	0	S_8	0

4. All **combined metals** and **non-metals** have oxidation numbers **equal** to their **valency /oxidation state** e.g.

Metal/non-metal ion	Valency	Oxidation state	Oxidation number
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Fe ²⁺	2	-2	-2
Fe ³⁺	3	-3	-3
Cu ²⁺	2	-2	-2
Cu ⁺	1	+1	+1
Cl ⁻	1	-1	-1
O ²⁻	2	-2	-2
Na ⁺	1	+1	+1
Al ³⁺	3	+3	+3
P ³⁻	3	-3	-3
Pb ²⁺	2	+2	+2

5. Sum of oxidation numbers of **atoms** of elements making a **compound** is **equal zero(0)** e.g.

Using this rule ,an unknown oxidation number of an atom in a compound can be determined as below:

a) CuSO₄ has-

- one atom of Cu with oxidation number +2(refer to Rule 4)
- one atom of S with oxidation number +6 (refer to Rule 4)
- six atoms of O each with oxidation number -2(refer to Rule 4)

Sum of oxidation numbers of **atoms** in CuSO₄ = (+2 + +6 + (-2 x 6)) = **0**

b) H₂SO₄ has-

- two atom of H each with oxidation number +1(refer to Rule 2)
- one atom of S with oxidation number +6 (refer to Rule 4)
- four atoms of O each with oxidation number -2(refer to Rule 4)

Sum of oxidation numbers of **atoms** in H₂SO₄ = (+2 + +6 + (-2 x 4)) = **0**

c) KMnO₄ has-

- one atom of K with oxidation number +1(refer to Rule 4)
- one atom of Mn with oxidation number +7 (refer to Rule 4)
- four atoms of O each with oxidation number -2(refer to Rule 4)

Sum of oxidation numbers of **atoms** in KMnO₄ = (+1 + +7 + (-2 x 4)) = **0**

Determine the oxidation number of:

I. Nitrogen in;

$$-\text{NO} \Rightarrow x + -2 = 0 \text{ thus } x = 0 - (-2) = +2$$

The chemical name of this compound is thus Nitrogen(II)oxide

$$-\text{NO}_2 \Rightarrow x + (-2 \times 2) = 0 \text{ thus } x = 0 - (-4) = +4$$

The chemical name of this compound is thus Nitrogen(IV)oxide

$$-\text{N}_2\text{O} \Rightarrow 2x + -2 = 0 \text{ thus } 2x = 0 - (-2) = +2/2 = +1$$

The chemical name of this compound is thus Nitrogen(I)oxide

II. Sulphur in;

$$-\text{SO}_2 \Rightarrow x + (-2 \times 2) = 0 \text{ thus } x = 0 - (-4) = +4$$

The chemical name of this compound is thus Sulphur(IV)oxide

$$-\text{SO}_3 \Rightarrow x + (-2 \times 3) = 0 \text{ thus } x = 0 - (-6) = +6$$

The chemical name of this compound is thus Sulphur(VI)oxide
 $-H_2SO_4 = ((+1 \times 2) + x + (-2 \times 4))$ thus $x = 0 - (+2 + -8) = +6$
 The chemical name of this compound is thus Sulphuric(VI)acid
 $-H_2SO_3 = ((+1 \times 2) + x + (-2 \times 3))$ thus $x = 0 - (+2 + -6) = +4$
 The chemical name of this compound is thus Sulphuric(IV)acid

III. Carbon in;

$-CO_2 \Rightarrow x + (-2 \times 2) = 0$ thus $x = 0 - (-4) = +4$
 The chemical name of this compound is thus carbon(IV)oxide
 $-CO \Rightarrow x + -2 = 0$ thus $x = 0 - -2 = +2$
 The chemical name of this compound is thus carbon(II)oxide
 $-H_2CO_3 = ((+1 \times 2) + x + (-2 \times 3))$ thus $x = 0 - (+2 + -6) = +4$
 The chemical name of this compound is thus Carbonic(IV)acid

IV. Manganese in;

$-MnO_2 \Rightarrow x + (-2 \times 2) = 0$ thus $x = 0 - (-4) = +4$
 The chemical name of this compound is thus Manganese(IV)oxide
 $-KMnO_4 = ((+1 + x + (-2 \times 4))$ thus $x = 0 - (+1 + -8) = +7$
 The chemical name of this compound is thus Potassium manganate(VII)

V. Chromium in;

$-Cr_2O_3 \Rightarrow 2x + (-2 \times 3) = 0$ thus $2x = 0 - (-6) = +6 / 2 = +3$
 The chemical name of this compound is thus Chromium(III)oxide
 $-K_2Cr_2O_7 \Rightarrow (+1 \times 2) + 2x + (-2 \times 7) = 0$
 thus $2x = 0 - +2 + -14 = +12 / 2 = +6$
 The chemical name of this compound is thus Potassium dichromate(VI)
 $-K_2CrO_4 \Rightarrow (+1 \times 2) + x + (-2 \times 4) = 0$
 thus $2x = 0 - +2 + -8 = +12 / 2 = +6$
 The chemical name of this compound is thus Potassium chromate(VI)

6. The **sum** of the oxidation numbers of atoms of elements making a charged radical/complex ion is **equal** to its **charge**.

Using this rule, the oxidation number of unknown atom of an element in a charged radical/complex ion can be determined as in the examples below;

a) SO_4^{2-} has-

- one atom of S with oxidation number +6(refer to Rule 4)
- four atoms of O each with oxidation number -2(refer to Rule 1)

Sum of oxidation numbers of **atoms** in $SO_4^{2-} = (+6 + (-2 \times 4)) = -2$

The chemical name of this radical is thus sulphate(VI) ion

b) NO_3^- has-

- one atom of N with oxidation number +4(refer to Rule 4)
- three atoms of O each with oxidation number -2(refer to Rule 1)

Sum of oxidation numbers of **atoms** in $NO_3^- = (+4 + (-2 \times 3)) = -1$

The chemical name of this radical is thus nitrate(IV) ion.

Determine the oxidation number of:

I. Nitrogen in;

$$-\text{NO}_2^- \Rightarrow x + (-2 \times 2) = -1 \text{ thus } x = -1 - (-4) = +3$$

The chemical name of this compound/ion/radical is thus Nitrate(III)ion

II. Sulphur in;

$$-\text{SO}_3^{2-} \Rightarrow x + (-2 \times 3) = -2 \text{ thus } x = -2 - (-6) = +4$$

The chemical name of this compound/ion/radical is thus Sulphate(IV)ion

III. Carbon in;

$$-\text{CO}_3^{2-} = x + (-2 \times 3) = -2 \text{ thus } x = -2 - (-6) = +4$$

The chemical name of this compound/ion/radical is thus Carbonate(IV)ion

IV. Manganese in;

$$-\text{MnO}_4^- = x + (-2 \times 4) = -1 \text{ thus } x = -1 - (-2 \times 4) = +7$$

The chemical name of this compound/ion/radical is thus manganate(VII) ion

V. Chromium in

$$-\text{Cr}_2\text{O}_7^{2-} \Rightarrow 2x + (-2 \times 7) = -2$$

$$\text{thus } 2x = -2 - (-2 \times 7) = +12 / 2 = +6$$

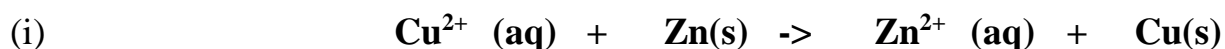
The chemical name of this compound/ion/radical is thus dichromate(VI) ion

$$-\text{CrO}_4^{2-} \Rightarrow x + (-2 \times 4) = -2$$

$$\text{thus } x = -2 - (-2 \times 4) = +6$$

The chemical name of this compound/ion/radical is thus chromate(VI) ion

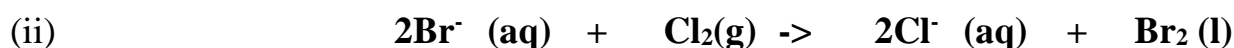
(c) Using the concept/idea of oxidation numbers as increase and decrease in oxidation numbers, the oxidizing and reducing species/agents can be determined as in the following examples;



Oxidation numbers \rightarrow $\begin{matrix} +2 & & 0 & & +2 & & 0 \end{matrix}$

Oxidizing species/agents \Rightarrow Cu^{2+} ; its oxidation number decrease from +2 to 0 in $\text{Cu}(\text{s})$

Reducing species/agents \Rightarrow Zn^{2+} ; its oxidation number increase from 0 to +2 in $\text{Zn}(\text{s})$



Oxidation numbers \rightarrow $\begin{matrix} -1 & & 0 & & -1 & & 0 \end{matrix}$

Oxidizing agent \Rightarrow $\text{Cl}_2(\text{g})$; its oxidation number decrease from 0 to -1 in $2\text{Cl}^- (\text{aq})$

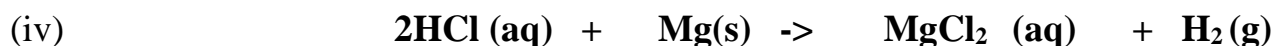
Reducing agents \Rightarrow Zn^{2+} ; its oxidation number increase from -1 to 0 in $\text{Zn}(\text{s})$



Oxidation numbers \rightarrow $\begin{matrix} 0 & & 0 & & +2 & & -1 \end{matrix}$

Oxidizing agent \Rightarrow $\text{Br}_2(\text{l})$; its oxidation number decrease from 0 to -1 in $2\text{Br}^- (\text{aq})$

Reducing agents \Rightarrow $\text{Zn}(\text{s})$; its oxidation number increase from 0 to +2 in Zn^{2+}



Oxidation numbers \rightarrow $\begin{matrix} 2(+1 -1) & & 0 & & +2 & 2(-1) & & 0 \end{matrix}$

Oxidizing agent \Rightarrow H^+ in HCl ; its oxidation number decrease from +1 to 0 in $\text{H}_2(\text{g})$

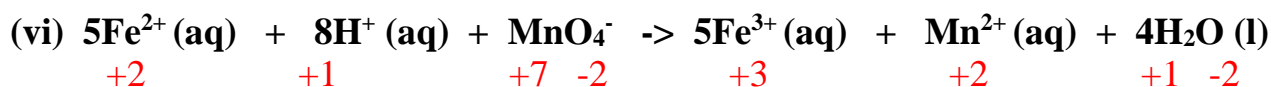
Reducing agents \Rightarrow $\text{Mg}(\text{s})$; its oxidation number increase from 0 to +2 in Mg^{2+}



Oxidation numbers \rightarrow $\begin{matrix} +1 & -2 & & 0 & & +1 & -2 & +1 & & 0 \end{matrix}$

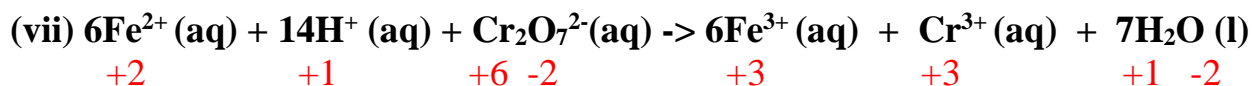
Oxidizing agent => **H⁺** in H₂O; its oxidation number decrease from +1 to 0 in **H₂(g)**

Reducing agents => **Na(s)**; its oxidation number increase from 0 to +1 in **Na⁺**



Oxidizing agent => **Mn** in **MnO₄⁻**; its oxidation number decrease from +7 to +2 in **Mn²⁺**

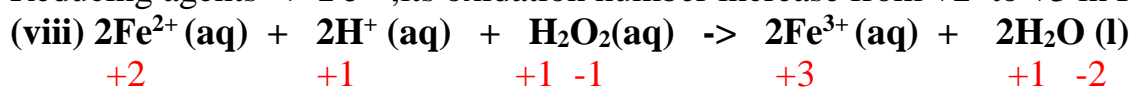
Reducing agents => **Fe²⁺**; its oxidation number increase from +2 to +3 in **Fe³⁺**



Oxidizing agent:

Cr in **Cr₂O₇²⁻**; its oxidation number decrease from +6 to +3 in **Cr³⁺**

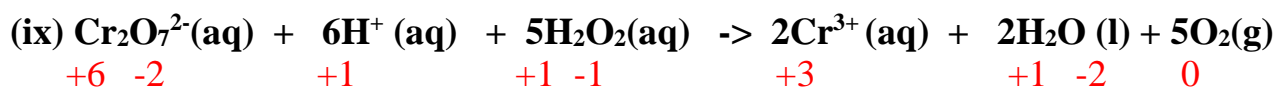
Reducing agents => **Fe²⁺**; its oxidation number increase from +2 to +3 in **Fe³⁺**



Oxidizing agent:

O in **H₂O₂**; its oxidation number decrease from -1 to -2 in **H₂O**

Reducing agents => **Fe²⁺**; its oxidation number increase from +2 to +3 in **Fe³⁺**



Oxidizing agents:

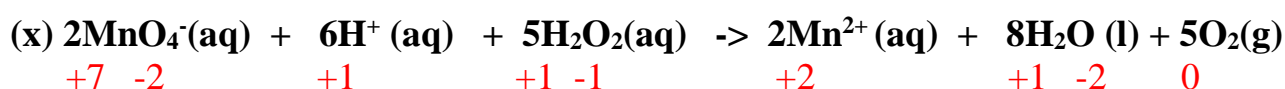
O in **H₂O₂**; its oxidation number decrease from -1 to -2 in **H₂O**

Cr in **Cr₂O₇²⁻**; its oxidation number decrease from +6 to +3 in **Cr³⁺**

Reducing agents

O in **H₂O₂**; its oxidation number increase from -1 to 0 in **O₂(g)**

O in **Cr₂O₇²⁻**; its oxidation number increase from -2 to 0 in **O₂(g)**



Oxidizing agents:

O in **H₂O₂**; its oxidation number decrease from -1 to -2 in **H₂O**

Mn in **MnO₄⁻**; its oxidation number decrease from +7 to +2 in **Mn²⁺**

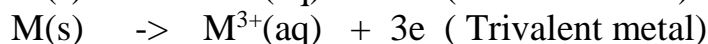
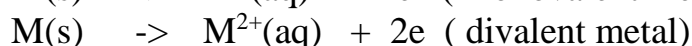
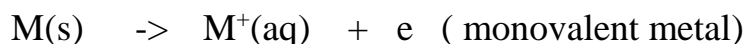
Reducing agents

O in **H₂O₂**; its oxidation number increase from -1 to 0 in **O₂(g)**

O in **MnO₄⁻**; its oxidation number increase from -2 to 0 in **O₂(g)**

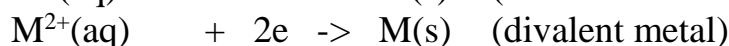
(ii)ELECTROCHEMICAL (VOLTAIC) CELL

1. When a metal rod/plate is put in a solution of its own salt, some of the metal ionizes and dissolve into the solution i.e.



The ions move into the solution leaving electrons on the surface of the metal rod/plate.

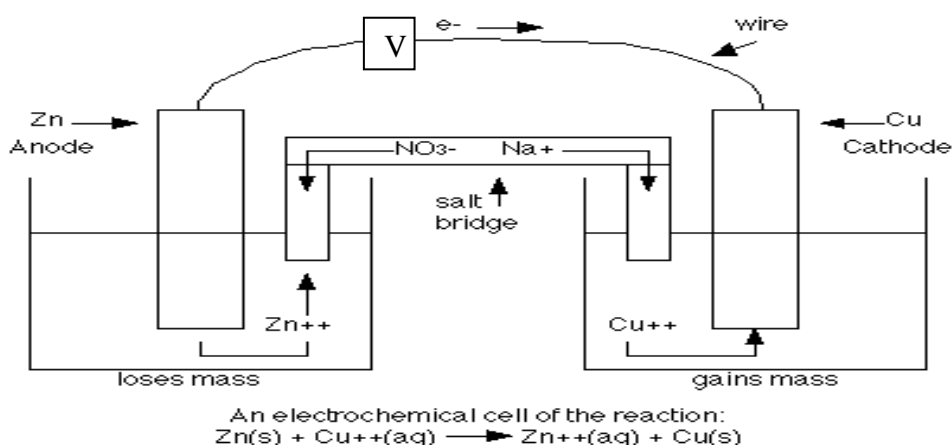
2.The **metal** rod becomes therefore **negatively** charged while its own **solution positively** charged. As the positive charges of the solution increase, some of them recombine with the electrons to form back the metal atoms



3. When a metal rod/plate is put in a solution of its own salt, it constitutes/forms a **half-cell**. The tendency of metals to ionize differ from one metal to the other. The difference can be measured by connecting two half cells to form an electrochemical/voltaic cell as in the below procedure:

To set up an electrochemical /voltaic cell
To compare the relative tendency of metals to ionize

Place 50cm³ of 1M Zinc(II) sulphate(VI) in 100cm³ beaker. Put a clean zinc rod/plate into the solution. Place 50cm³ of 1M Copper(II) sulphate(VI) in another 100cm³ beaker. Put a clean copper rod/plate of equal area (length x width) with Zinc into the solution. Connect/join the two metals(to a voltmeter) using connecting wires. Dip a folded filter paper into a solution of Potassium nitrate(V) or sodium(I) chloride(I) until it soaks. Use the folded soaked filter paper to connect/join the two solutions in the two beakers. The whole set up should be as below



Repeat the above procedure by replacing:

- (i) Zinc half cell with Magnesium rod/plate/ribbon dipped in 50cm³ of 1M magnesium (II) sulphate(VI) solution
- (ii) Zinc half cell with Silver rod/plate/coin dipped in 50cm³ of 1M silver(I) nitrate(V) solution
- (iii) Copper half cell with Iron rod/plate/spoon dipped in 50cm³ of 1M Iron (II) sulphate(VI) solution

Record the observations in the table below

Changes on the 1 st metal rod (A)	Changes on the 2 nd metal rod (B)	Changes on the 1 st solution (A(aq))	Changes on the 2 nd solution (B(aq))	Voltage/voltmeter reading(Volts)
Using Zn/Cu half cell -The rod decrease in size /mass /dissolves/ erodes	-copper rod /plate increase in size /mass/ deposited	Zinc(II)sulphate (VI)colour remain colourless	Blue Copper (II)sulphate (VI)colour fades. Brown solid/residue/ deposit	0.8 (Theoretical value= 1.10V)

Using Mg/Cu half cell -The rod decrease in size /mass /dissolves/ erodes	-copper rod /plate increase in size /mass/ deposited	Magnesium(II) sulphate(VI) colour remain colourless	Blue Copper (II)sulphate (VI)colour fades Brown solid/residue/ deposit	1.5 (Theoretical value= 2.04V)
Using Ag/Cu half cell -The rod increase in size /mass /deposited	-silver coin/ rod /plate increase in size /mass/ deposited	Blue Copper (II)sulphate (VI)colour remains	Silver(I)nirate (V)colour remain colourless	0.20 (Theoretical value= 0.46V)
Using Fe/Cu half cell -The rod decrease in size /mass /dissolves/ erodes	-copper rod /plate increase in size /mass/ deposited	Iron(II)sulphate (VI)colour becomes more green	Blue Copper (II)sulphate (VI)colour fades.Brown solid/residue/ deposit	0.60 (Theoretical value= 0.78V)

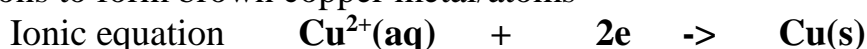
From the above observations ,it can be deduced that:

(i)in the Zn/Cu half-cell the;

-Zinc rod/plate ionizes /dissolves faster than the copper rod/plate to form Zn^{2+}



-blue copper ions in the Copper (II)sulphate solution gains the donated electrons to form brown copper metal/atoms



This reaction shows /imply the **Zinc** rod has a **higher** tendency to ionize than **copper**.The **Zinc** rod has a **higher net** accumulation of electrons and is more **negative** compared to the copper rod which has **lower** accumulation of electrons. The copper rod is therefore relatively more **positive** with respect to Zinc rod.

When the two half cells are connected , electrons therefore flow from the **negative** Zinc rod through the external wire to be gained by copper **ions**. This means a net accumulation/**increase** of Zn^{2+} positive ions on the negative half cell and a net **decrease** in Cu^{2+} positive ions on the positive half cell.

The purpose of the salt bridge therefore is:

(i)complete the circuit

(ii)maintain balance of charges /ions on both half cells.

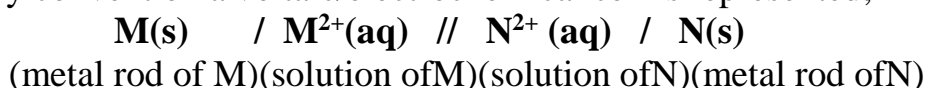
For the negative half cell the $\text{NO}_3^- / \text{Cl}^-$ from salt bridge **decrease/neutralise** the increased **positive**(Zn^{2+}) ion.

For the positive half cell the Na^+ / K^+ from salt bridge **increase** the decreased **positive**(Cu^{2+}) ion.

The voltmeter should theoretically register/read a 1.10Volts as a measure of the electromotive force (**e.m.f**) of the cell .Practically the voltage reading is lowered because the connecting wires have some **resistance** to be overcome.

A combination of two half cells that can **generate** an electric **current** from a **redox** reaction is called a voltaic/electrochemical cell.

By convention a voltaic/electrochemical cell is represented;



Note;

a)(i)Metal M must be the one **higher** in the reactivity series.

(ii)It forms the **negative** terminal of the cell.

(iii)It must diagrammatically be drawn **first** on the **left hand side** when illustrating the voltaic/electrochemical cell.

b)(i)Metal N must be the one **lower** in the reactivity series.

(ii)It forms the **positive** terminal of the cell.

(iii)It must diagrammatically be drawn second/**after/ right hand side** when illustrating the voltaic/electrochemical cell.

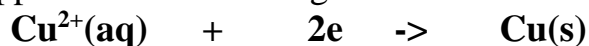
Illustration of the voltaic/electrochemical cell.

(i)Zn/Cu cell

1. Zinc rod ionizes /dissolves to form Zn^{2+} ions at the negative terminal



2. Copper ions in solution gain the donated electrons to form copper atoms/metal



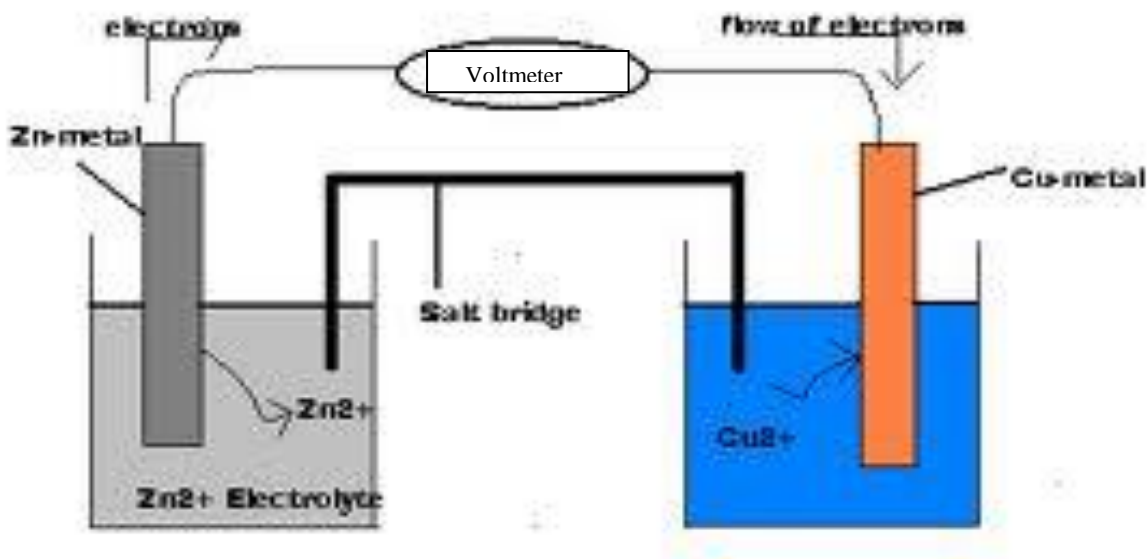
3.Overall redox equation



4.cell representation.



5.cell diagram

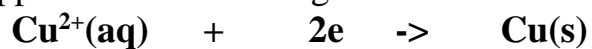


(ii) Mg/Cu cell

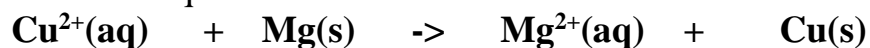
1. Magnesium rod ionizes /dissolves to form Mg^{2+} ions at the negative terminal



2. Copper ions in solution gain the donated electrons to form copper atoms/metal



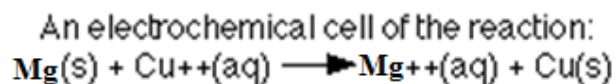
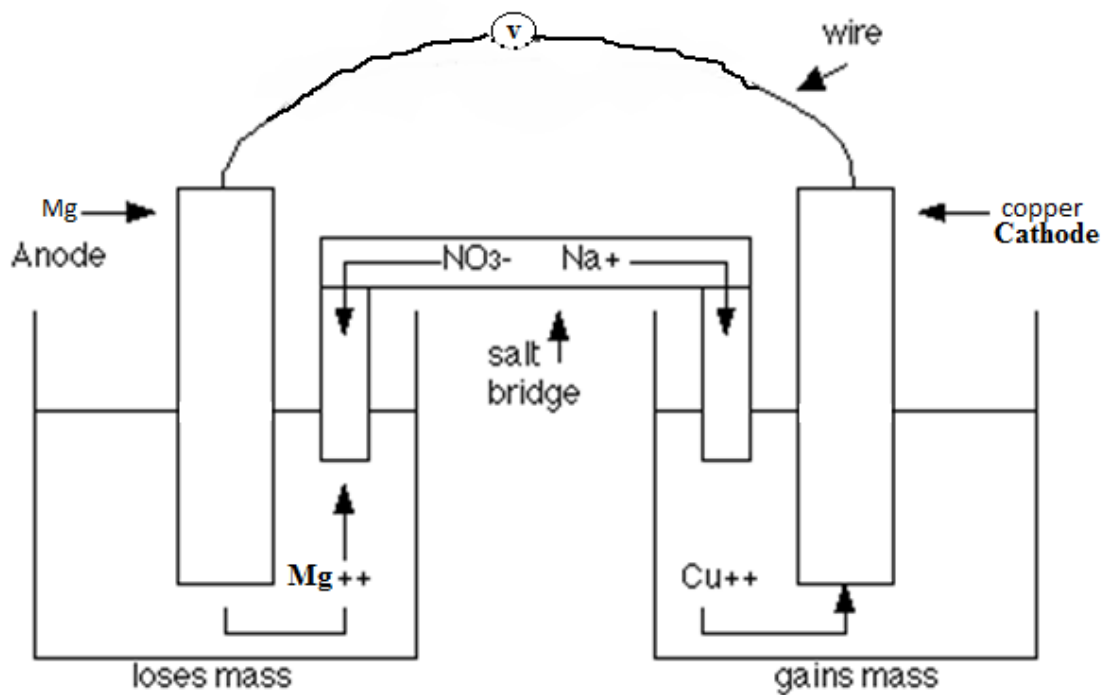
3. Overall redox equation



4. cell representation.



5. cell diagram.



(iii) Fe/Cu cell

1. Magnesium rod ionizes /dissolves to form Mg^{2+} ions at the negative terminal

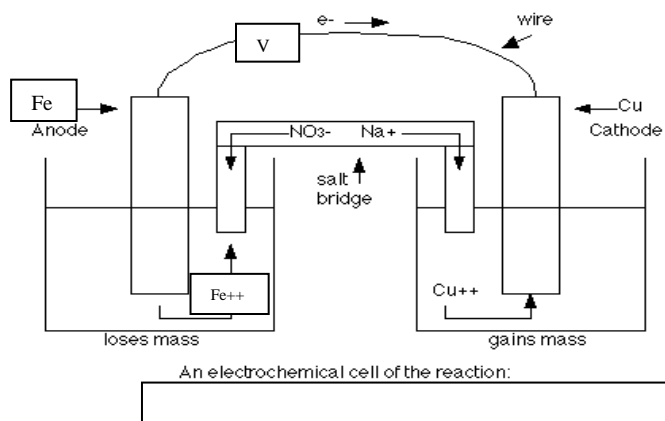
$$\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}$$
2. Copper ions in solution gain the donated electrons to form copper atoms/metal

$$\text{Cu}^{2+}(\text{aq}) + 2\text{e} \rightarrow \text{Cu(s)}$$
3. Overall redox equation

$$\text{Cu}^{2+}(\text{aq}) + \text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Cu(s)}$$
4. cell representation.

$$\text{Fe(s)} / 1\text{M, Fe}^{2+}(\text{aq}) // 1\text{M, Cu}^{2+}(\text{aq}) / \text{Cu(s)} \quad E^0 = +0.78 \text{ V}$$

5. cell diagram.

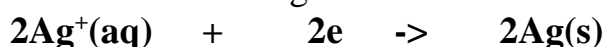


(iv) Ag/Cu cell

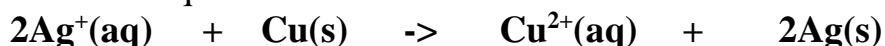
1. Copper rod ionizes /dissolves to form Cu^{2+} ions at the negative terminal



2. Silver ions in solution gain the donated electrons to form silver atoms/metal



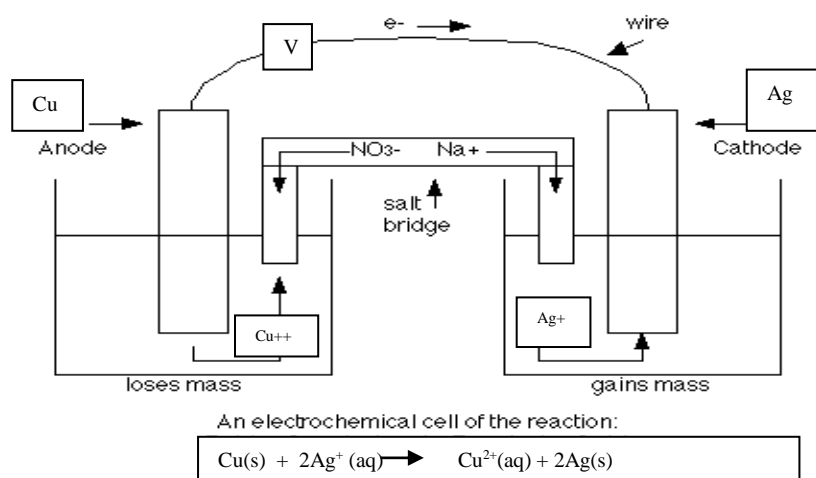
3. Overall redox equation



4. cell representation.



5. cell diagram.



Standard electrode potential (E^0)

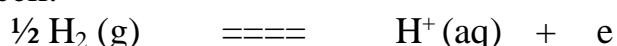
The **standard** electrode potential (E^0) is obtained if the **hydrogen** half cell is used as **reference**. The standard electrode potential (E^0) consist of inert platinum electrode immersed/dipped in 1M solution of (sulphuric(VI) acid) H^+ ions. Hydrogen gas is bubbled on the platinum electrodes at:

(i) a temperature of 25°C

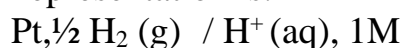
(ii) atmospheric pressure of $101300\text{Pa}/101300\text{Nm}^{-2}/1\text{atm}/760\text{mmHg}/76\text{cmHg}$

(iii) a concentration of $1\text{M}(1\text{moledm}^{-3})$ of sulphuric(VI) acid/ H^+ ions and $1\text{M}(1\text{moledm}^{-3})$ of the other half cell.

Hydrogen is **adsorbed** onto the **surface** of the platinum. An **equilibrium/balance** exist between the adsorbed layer of molecular hydrogen and H^+ ions in solution to form a half cell.



The half cell representation is:



The standard electrode potential (E^0) is thus defined as **the potential difference for a cell comprising of a particular element in contact with 1M solution of its own ions and the standard hydrogen electrode.**

If the other electrode has a **higher/greater** tendency to lose electrons than the hydrogen electrode, the electrode is therefore **negative** with respect to hydrogen electrode and its electrode potential has **negative (E^0)** values.

If the other electrode has a **lower/lesser** tendency to lose electrons than the hydrogen electrode, the electrode is therefore **positive** with respect to hydrogen electrode and its electrode potential has **positive (E^0)** values.

Table showing the standard electrode potential (E^0) of some reactions

Reaction	(E^0) values in volts
$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	+2.87
$H_2O_2(aq) + H^+(aq) + 2e^- \rightarrow H_2O(l)$	+1.77
$MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + H_2O(l)$	+1.70
$2HClO(aq) + 2H^+(aq) + 2e^- \rightarrow Cl_2(aq) + 2H_2O(l)$	+1.59
$MnO_4^-(aq) + 4H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + H_2O(l)$	+1.51
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$	+1.23
$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$2H^+(aq) + O_2(g) \rightarrow H_2O_2(aq)$	+0.68
$I_2(aq) + 2e^- \rightarrow 2I^-(aq)$	+0.54
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	+0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	-0.13
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	-0.44
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	-0.77
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	-2.37
$Na^+(aq) + e^- \rightarrow Na(s)$	-2.71
$K^+(aq) + e^- \rightarrow K(s)$	-2.92

Note:

(i) E^0 values generally show the **possibility/feasibility** of a **reduction** process/**oxidizing strength**.

(ii) The **element/species** in the half cell with the **highest** negative E^0 value easily **gain / acquire** electrons.

It is thus the **strongest oxidizing agent** and its reduction process is **highly possible/feasible**. The **element/species** in the half cell with the **lowest** positive E^0 value easily **donate / lose** electrons.

It is thus the **strongest reducing agent** and its reduction process is the **least possible/feasible**.

(iii) The **overall** redox reaction is **possible/feasible** if it has a **positive (+) E^{\ominus}** .

If the **overall** redox reaction is **not possible/ not feasible/ forced**, it has a **negative (-) E^{\ominus}**

Calculation examples on E^{\ominus}

Calculate the E^{\ominus} value of a cell made of:

a) Zn and Cu

From the table above:

$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu}(\text{s}) \quad E^{\ominus} = +0.34\text{V}$ (higher E^{\ominus} / **Right Hand Side** diagram)

$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s}) \quad E^{\ominus} = -0.77\text{V}$ (lower E^{\ominus} / **Left Hand Side** diagram)

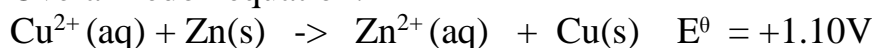
$\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \quad E^{\ominus} = +0.77$ (reverse lower E^{\ominus} to derive cell reaction / representation)

Overall $E^{\ominus} = E^{\ominus}$ higher- E^{\ominus} lower / E^{\ominus} RHS - E^{\ominus} LHS/ E^{\ominus} oxidized- E^{\ominus} reduced

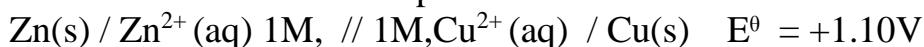
Substituting:

$$\text{Overall } E^{\ominus} = +0.34 - (-0.77) = \underline{+1.10\text{V}}$$

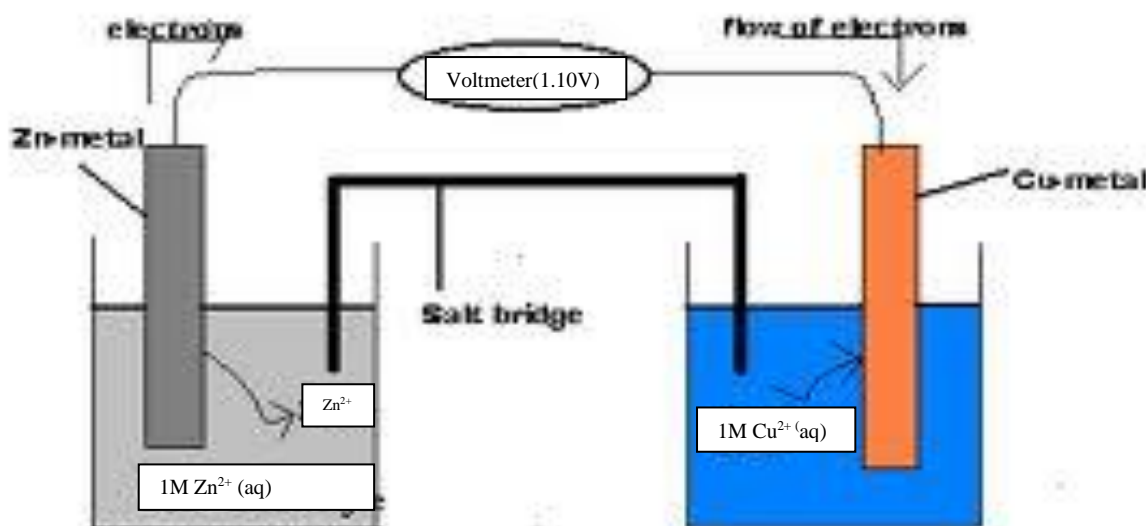
Overall redox equation:



Overall conventional cell representation:



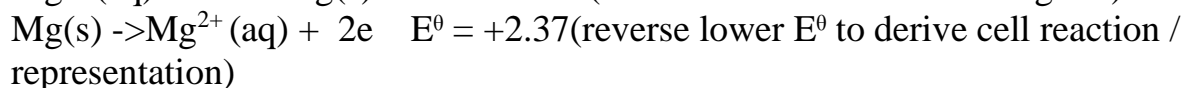
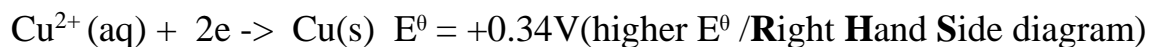
Overall conventional cell diagram:



Zinc and copper reaction has a **positive(+)** overall E^{θ} therefore is possible/feasible and thus Zinc can displace/reduce Copper solution.

b)Mg and Cu

From the table above:

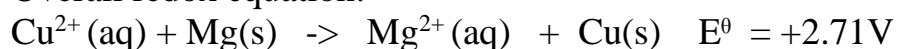


Overall $E^{\theta} = E^{\theta} \text{ higher- } E^{\theta} \text{ lower} / E^{\theta} \text{ RHS} - E^{\theta} \text{ LHS} / E^{\theta} \text{ oxidized- } E^{\theta} \text{ reduced}$

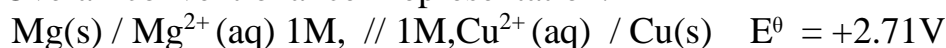
Substituting:

$$\text{Overall } E^{\theta} = +0.34 - (-2.37) = \underline{\underline{+2.71\text{V}}}$$

Overall redox equation:

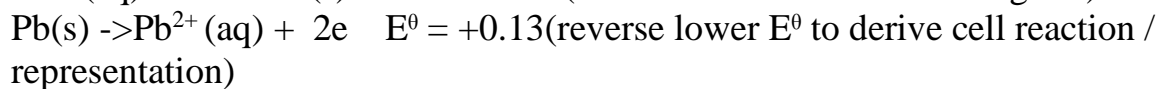


Overall conventional cell representation:



c)Ag and Pb

From the table above:

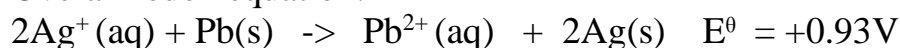


Overall $E^{\theta} = E^{\theta} \text{ higher- } E^{\theta} \text{ lower} / E^{\theta} \text{ RHS} - E^{\theta} \text{ LHS} / E^{\theta} \text{ oxidized- } E^{\theta} \text{ reduced}$

Substituting:

$$\text{Overall } E^{\theta} = +0.80 - (-0.13) = \underline{\underline{+0.93\text{V}}}$$

Overall redox equation:

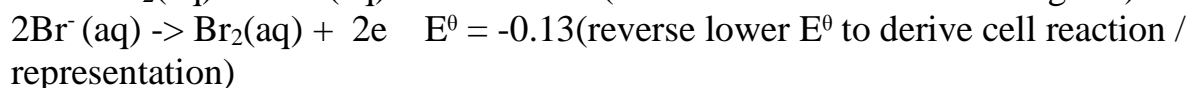
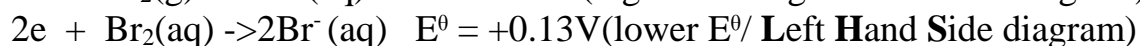
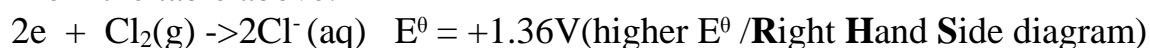


Overall conventional cell representation:



d)Chlorine and Bromine

From the table above:

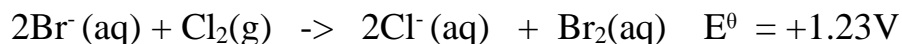


Overall $E^{\theta} = E^{\theta} \text{ higher- } E^{\theta} \text{ lower} / E^{\theta} \text{ RHS} - E^{\theta} \text{ LHS} / E^{\theta} \text{ oxidized- } E^{\theta} \text{ reduced}$

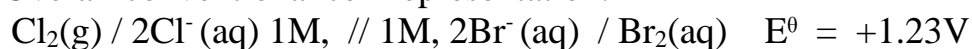
Substituting:

$$\text{Overall } E^{\theta} = -0.13 - (-1.36) = \underline{\underline{+1.23\text{V}}}$$

Overall redox equation:



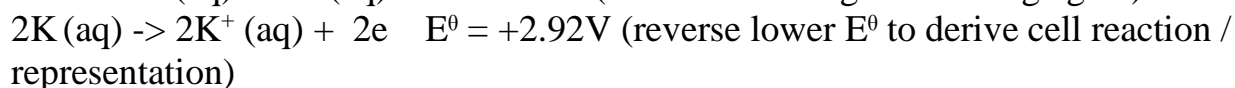
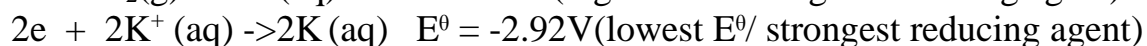
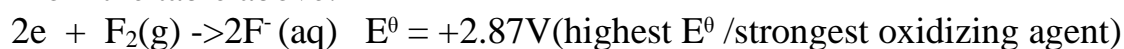
Overall conventional cell representation:



Chlorine displaces bromine from bromine water. When chlorine gas is thus bubbled in bromine water, the pale **green** colour **fades** as displacement takes place and a **brown** solution containing dissolved bromine liquid is **formed**. This reaction is feasible /possible because the overall redox reaction has a **positive** E^θ value.

e) Strongest oxidizing agent and the strongest reducing agent.

From the table above:

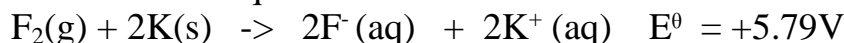


Overall $E^\theta = E^\theta$ higher- E^θ lower / E^θ RHS - E^θ LHS/ E^θ oxidized- E^θ reduced

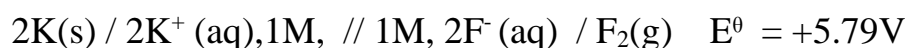
Substituting:

$$\text{Overall } E^\theta = +2.87 - (-2.92) = \underline{\underline{+5.79\text{V}}}$$

Overall redox equation:



Overall conventional cell representation:



The redox reactions in an electrochemical/voltaic is commercially applied to make the:

(a) Dry /primary/Laclanche cell.

(b) Wet /secondary /accumulators.

(a) Dry/primary/Laclanche cell

Examine a used dry cell.

Note the positive and the negative terminal of the cell. Carefully using a knife cut a cross section from one terminal to the other.

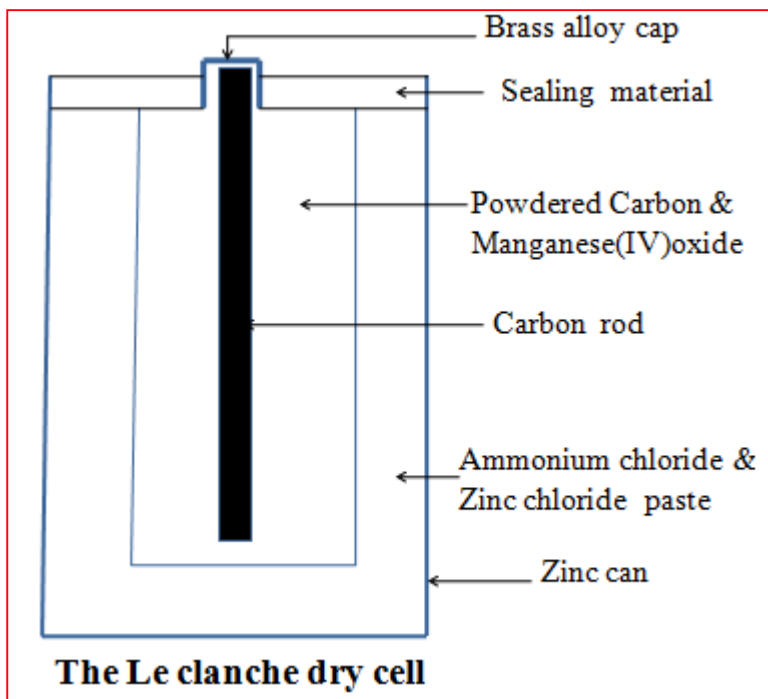
The dry cell consist of a **Zinc can** containing a **graphite rod** at the **centre** surrounded by a **paste** of;

-Ammonium chloride

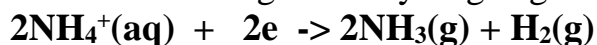
-Zinc chloride

-powdered manganese (IV) oxide mixed with Carbon.

Zinc acts/serve as the negative terminal where it ionizes/dissociates:



Ammonium ions in ammonium chloride serve as the positive terminal where it is converted to ammonia gas and hydrogen gas.



Ammonia forms a complex salt / compound / $(\text{Zn}(\text{NH}_3)_4)^{2+}(\text{aq})$ / tetramminezinc(II) complex with the Zinc chloride in the paste.

Manganese (IV) oxide oxidizes the hydrogen produced at the electrodes to water preventing any bubbles from coating the carbon terminal which would reduce the efficiency of the cell.

Ammonium chloride is used as paste because the solid does not conduct electricity because the ions are fused/not mobile.

Since the reactants are used up, the dry /primary /Laclanche cell cannot provide continuous supply of electricity. The process of restoring the reactants is called **recharging**.

b)Wet/Secondary/Accumulators

1. Wet/Secondary/Accumulators are **rechargeable** unlike dry /primary /Laclanche cells. Wet/Secondary/Accumulators are made up of:

(i)**Lead** plate that forms the **negative** terminal

(ii)**Lead(IV) oxide** that forms the **positive** terminal

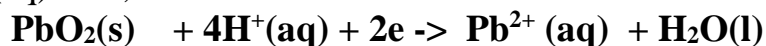
2. The two electrodes are dipped in concentrated sulphuric(VI) acid of a relative density 1.2/1.3

3. At the negative terminal, lead ionizes /dissolves;

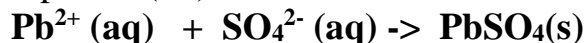


4. At the positive terminal,

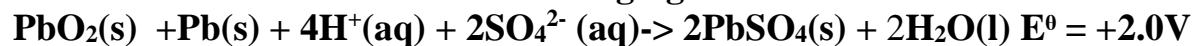
(i) Lead(IV) oxide **reacts** with the hydrogen ions in sulphuric(VI)acid to form Pb^{2+} (aq) ions;



(ii) Pb^{2+} (aq) ions formed **instantly** react with sulphate (VI) ions/ SO_4^{2-} (aq) from sulphuric (VI)acid to form **insoluble** Lead(II) sulphate (VI).



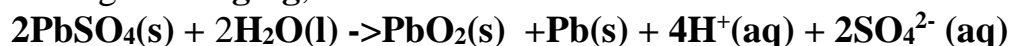
5.The **overall** cell reaction is called **discharging**



6.The insoluble Lead(II) sulphate (VI) formed should not be left for long since fine Lead(II) sulphate (VI) will change to a coarse non-reversible and inactive form making the cell less efficient.

As the battery discharges ,lead and lead(IV)oxide are depleted/finished/reduced and the concentration of sulphuric(VI)acid decreases.

7. During **recharging**, the electrode reaction is reversed as below:



8. A car battery has six Lead-acid cells making a total of 12 volts.

(iii)ELECTROLYSIS (ELECTROLYTIC CELL)

1.Electrolysis is defined simply as the **decomposition** of a **compound** by an electric current/**electricity**.

A compound that is decomposed by an electric current is called an electrolyte. Some electrolytes are **weak** while others are **strong**.

2.Strong electrolytes are those that are fully ionized/dissociated into (many) ions.

Common strong electrolytes include:

(i)all **mineral** acids

(ii)all strong **alkalis**/sodium hydroxide/potassium hydroxide.

(iii)all soluble **salts**

3. Weak electrolytes are those that are partially/partly ionized/dissociated into (few) ions.

Common weak electrolytes include:

- (i) all **organic** acids
- (ii) all **bases** except sodium hydroxide/potassium hydroxide.
- (iii) **Water**

4. A compound that is **not** decomposed by an electric current is called non-electrolyte. Non-electrolytes are those compounds /substances that exist as molecules and thus cannot ionize/dissociate into (any) ions .

Common non-electrolytes include:

- (i) most organic solvents (e.g. petrol/paraffin/benzene/methylbenzene/ethanol)
- (ii) all hydrocarbons (alkanes /alkenes/alkynes)
- (iii) Chemicals of life (e.g. proteins, carbohydrates, lipids, starch, sugar)

5. An electrolyte in **solid** state has **fused** /joined ions and therefore does **not** conduct electricity but the **ions** (cations and anions) are **free** and **mobile** in **molten** and **aqueous** (solution, dissolved in water) state.

6. During electrolysis, the free ions are attracted to the **electrodes**. An electrode is a rod through which current enters and leaves the electrolyte during electrolysis. An electrode that does not influence/alter the products of electrolysis is called an **inert electrode**.

Common inert electrodes include:

- (i) **Platinum**
- (ii) **Carbon graphite**

Platinum is not usually used in a school laboratory because it is very **expensive**. Carbon graphite is **easily**/readily and **cheaply** available (from used dry cells).

7. The **positive** electrode is called **Anode**. The anode is the electrode through which **current enters** the electrolyte/**electrons leave** the electrolyte

8. The **negative** electrode is called **Cathode**. The cathode is the electrode through which **current leaves** the electrolyte / **electrons enter** the electrolyte

9. During the electrolysis, free **anions** are attracted to the **anode** where they **lose** /**donate** electrons to form **neutral** atoms/molecules. i.e.

$M(l) \rightarrow M^+(l) + e$ (for cations from molten electrolytes)

$M(s) \rightarrow M^+(aq) + e$ (for cations from electrolytes in aqueous state / solution / dissolved in water)

The neutral atoms /molecules form the **products** of electrolysis at the anode. This is called **discharge** at anode

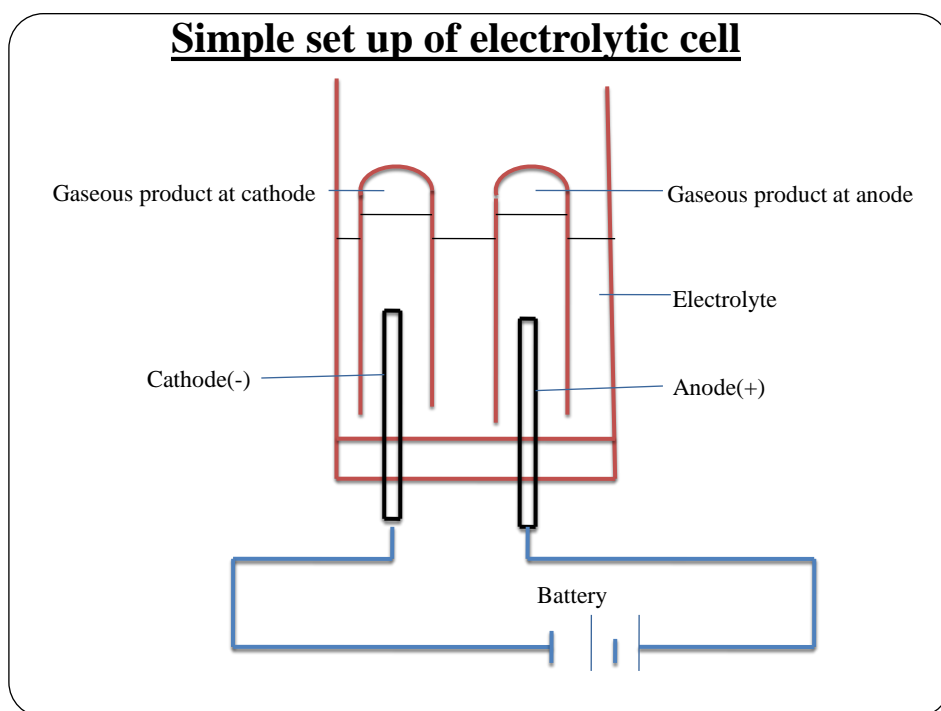
10. During electrolysis, free **cations** are attracted to the **cathode** where they **gain** /**accept/acquire** electrons to form **neutral** atoms/molecules.

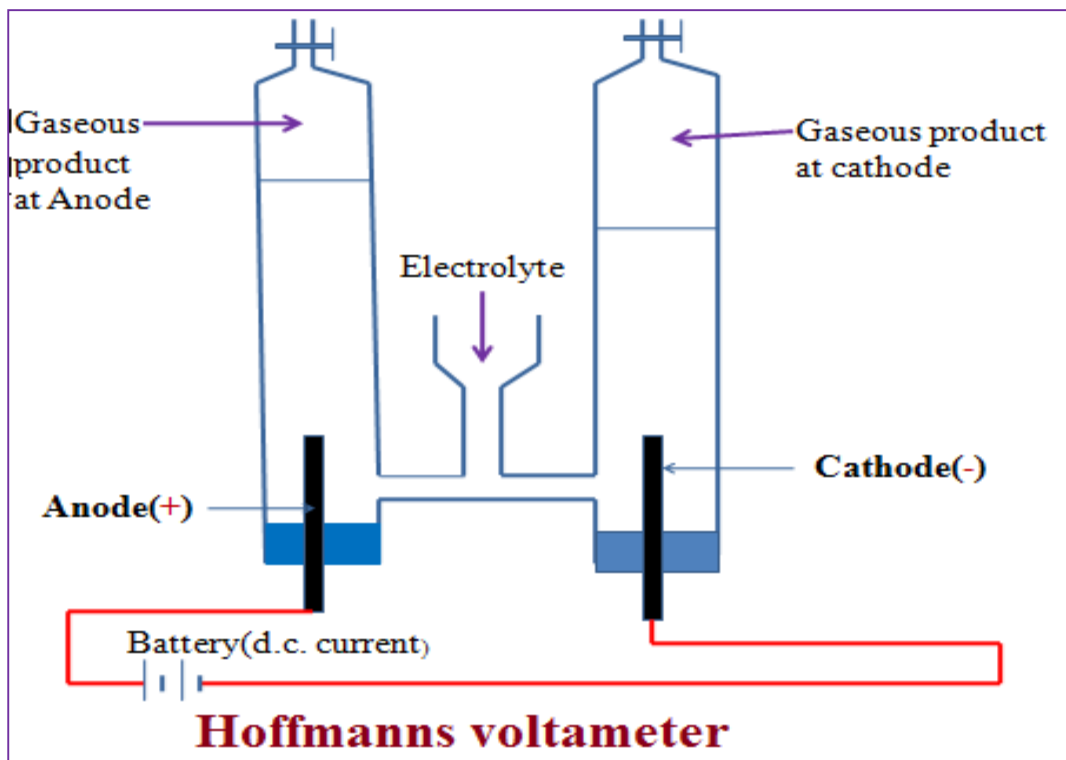
$X^+ (aq) + 2e \rightarrow X(s)$ (for cations from electrolytes in aqueous state / solution / dissolved in water)

$2X^+ (l) + 2e \rightarrow X (l)$ (for cations from molten electrolytes)

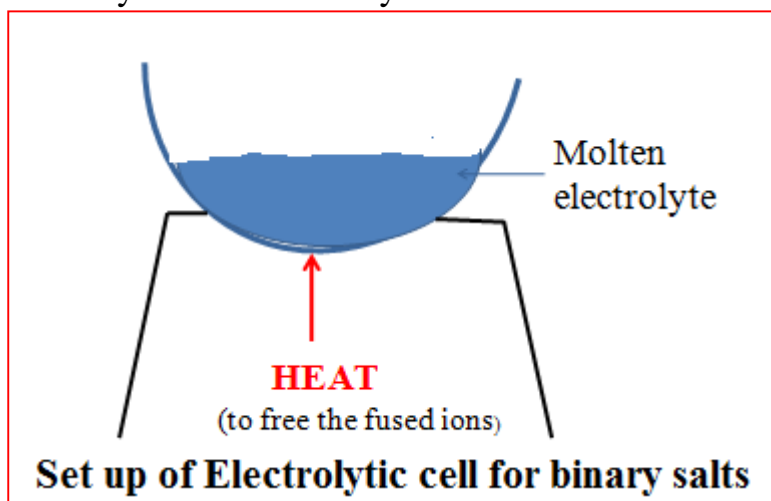
The neutral atoms /molecules form the **products** of electrolysis at the cathode. This is called **discharge** at cathode.

11. The below set up shows an electrolytic cell.





12. For a compound /salt containing only two ion/binary salt the products of electrolysis in an electrolytic cell can be determined as in the below examples:



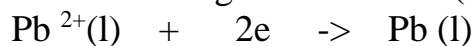
a) To determine the products of electrolysis of molten Lead(II)chloride

(i) Decomposition of electrolyte into free ions;



(Compound decomposed into free cation and anion in **liquid** state)

(ii) At the cathode/negative electrode(-);



(Cation / Pb^{2+} gains / accepts / acquires electrons to form free **atom**)

(iii)At the anode/positive electrode(+);



(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule**)

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid lead metal.

II.At the anode pale green chlorine gas.

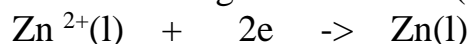
b)To determine the products of electrolysis of molten Zinc bromide

(i)Decomposition of electrolyte into free ions;



(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);



(Cation / Zn^{2+} gains / accepts / acquires electrons to form free **atom**)

(iii)At the anode/positive electrode(+);



(Anion / Br^- donate/lose electrons to form free **atom** then a liquid **molecule** which change to **gas** on heating)

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid Zinc metal.

II.At the anode **red** bromine **liquid** / **red/brown** bromine **gas**.

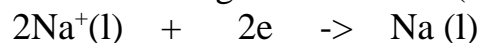
c)To determine the products of electrolysis of molten sodium chloride

(i)Decomposition of electrolyte into free ions;



(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);



(Cation / Na^+ gains / accepts / acquires electrons to form free **atom**)

(iii)At the anode/positive electrode(+);



(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule**)

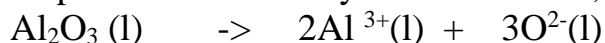
(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid sodium metal.

II.At the anode pale green chlorine gas.

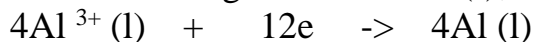
d) To determine the products of electrolysis of molten Aluminium (III)oxide

(i) Decomposition of electrolyte into free ions;



(Compound decomposed into free cation and anion in **liquid** state)

(ii) At the cathode/negative electrode(-);



(Cation / Al^{3+} gains / accepts / acquires electrons to form free **atom**)

(iii) At the anode/positive electrode(+);



(Anion / 6O^{2-} donate/lose 12 electrons to form free **atom** then three gas **molecule**)

(iv) Products of electrolysis therefore are;

I. At the cathode grey beads /solid aluminium metal.


II. At the anode colourless gas that relights/rekindles glowing splint.

13. For a compound /salt mixture containing **many** ions in an electrolytic cell, the **discharge** of ions in the cell depend on the following **factors**:

a) Position of cations and anions in the electrochemical series

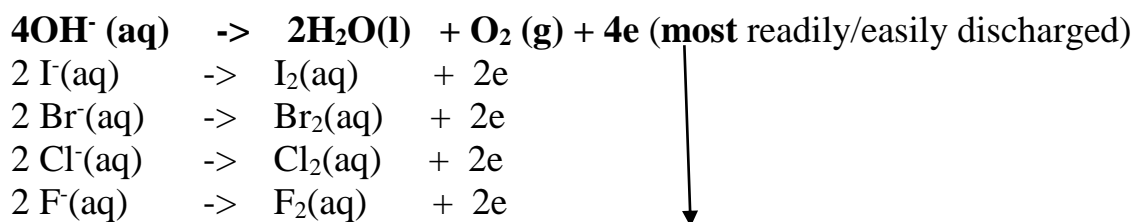
1. Most electropositive cations require more energy to reduce (gain electrons) and thus not readily discharged. The **higher** elements /metals in the electrochemical series the **less easily/readily** it is **discharged** at the cathode in the electrolytic cell.

Table I showing the relative ease of discharge of **cations** in an electrolytic cell

$\text{K}^+(\text{aq})$	$+ \text{e}^- \rightarrow \text{K}(\text{s})$	<div style="text-align: center;"> <p>(least readily/easily discharged)</p>  <p>(hydrogen is usually “metallic”)</p> <p>(most readily/easily discharged)</p> </div>
$\text{Na}^+(\text{aq})$	$+ \text{e}^- \rightarrow \text{Na}(\text{s})$	
$\text{Ca}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Ca}(\text{s})$	
$\text{Mg}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Mg}(\text{s})$	
$\text{Al}^{3+}(\text{aq})$	$+ 3\text{e}^- \rightarrow \text{Al}(\text{s})$	
$\text{Zn}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	
$\text{Fe}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	
$\text{Pb}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	
$2\text{H}^+(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	
$\text{Cu}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	
$\text{Hg}^{2+}(\text{aq})$	$+ 2\text{e}^- \rightarrow \text{Hg}(\text{s})$	
$\text{Ag}^+(\text{aq})$	$+ \text{e}^- \rightarrow \text{Ag}(\text{s})$	(most readily/easily discharged)

2. The **OH^-** ion is the **most** readily/easily **discharged** anion . All the other anionic **radicals**(SO_4^{2-} , SO_3^{2-} , CO_3^{2-} , HSO_4^- , HCO_3^- , NO_3^- , PO_4^{3-})are not/never discharged. The ease of discharge of halogen ions increase down the group.

Table II showing the relative ease of discharge of **anions** in an electrolytic cell



SO_4^{2-} , SO_3^{2-} , CO_3^{2-} , HSO_4^- , HCO_3^- , NO_3^- , PO_4^{3-} **not/never/rarely** discharged.

3.(a) When **two** or more **cations** are attracted to the **cathode**, the ion **lower** in the electrochemical series is discharged **instead** of that which is **higher** as per the table I above. This is called selective/preferential discharge at cathode.

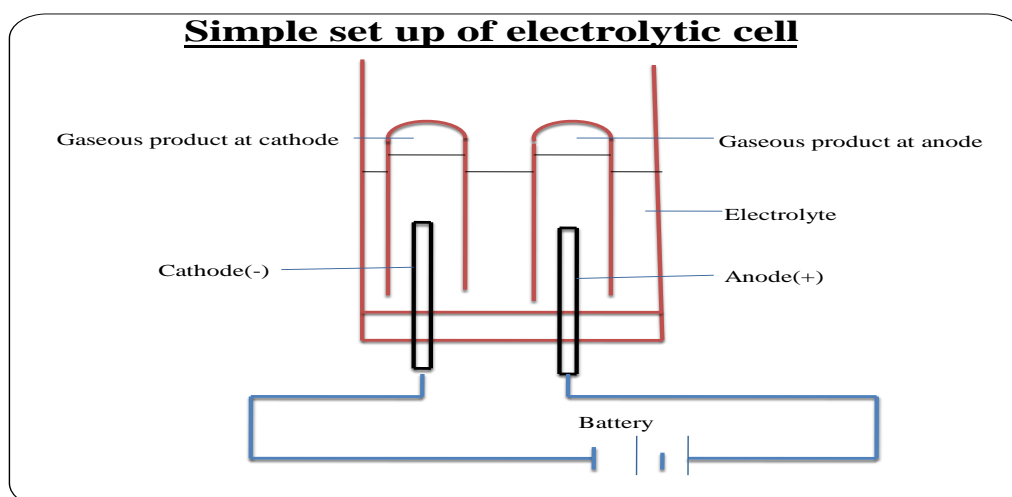
(b) When **two** or more **anions** are attracted to the **anode**, the ion **higher** in the electrochemical series is discharged **instead** of that which is **lower** as per the table I above. This is called selective/preferential discharge at anode.

4. The following experiments show the influence /effect of selective/preferential discharge on the products of electrolysis:

(i) Electrolysis of acidified water/dilute sulphuric(VI) acid

Fill the Hoffmann voltameter with dilute sulphuric(VI) acid. Connect the Hoffmann voltameter to a d.c. electric supply. Note the observations at each electrode.

Electrolytic cell set up during electrolysis of acidified water/dilute sulphuric(VI) acid



Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.





II. Name the ions in acidified water that are attracted/move to:

Cathode- $\text{H}^+(\text{aq})$ from either sulphuric(VI) acid (H_2SO_4) or water (H_2O)

Anode- $\text{SO}_4^{2-}(\text{aq})$ from sulphuric (VI) acid (H_2SO_4) and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2(\text{g})$

Anode $4\text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + 4\text{e}^-$

(4OH^- ions selectively discharged instead of SO_4^{2-} ions at the anode)

IV. Name the products of electrolysis of acidified water.

Cathode- Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound)

Anode- Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

The **four**(4) electrons donated/lost by OH^- ions to form **1** molecule/**1**volume/**1**mole of **oxygen** (O_2)gas at the **anode** are gained/acquired/accepted by the four $\text{H}^+(\text{aq})$ ions to form **2** molecule/**2**volume/**2**mole of **Hydrogen** (H_2)gas at the **cathode**.

The volume of **Oxygen** gas at the anode is thus a **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is thus a **twice** the volume of **Oxygen** produced at the anode.

VI. Why is electrolysis of dilute sulphuric(VI) acid called "electrolysis of (acidified) water"?

The ratio of $\text{H}_2(\text{g})$: $\text{O}_2(\text{g})$ is **2:1** as they are combined in water. This implies/means that water in the electrolyte is being decomposed into hydrogen and Oxygen gases. The electrolysis of dilute sulphuric acid is therefore called "electrolysis of acidified water."

VI. Explain the changes in concentration of the electrolyte during electrolysis of acidified water"

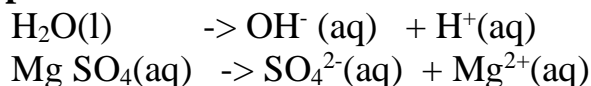
The concentration of dilute sulphuric (VI) acid **increases**. Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape. The concentration /mole of acid present in a given volume of solution thus continue increasing/rising.

(ii)Electrolysis of Magnesium sulphate(VI) solution

Fill the Hoffmann voltameter with dilute sulphuric(VI) acid. Connect the Hoffmann voltameter to a d.c. electric supply. Note the observations at each electrode.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

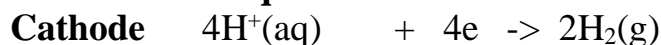


II. Name the ions in Magnesium sulphate(VI) solution that are attracted/move to:

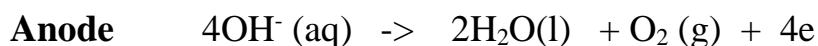
Cathode- $\text{Mg}^{2+}(\text{aq})$ from Magnesium sulphate(VI) solution (Mg SO_4) and $\text{H}^+(\text{aq})$ from water (H_2O)

Anode- $\text{SO}_4^{2-}(\text{aq})$ from Magnesium sulphate(VI) solution (Mg SO_4) and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:



H^+ ions selectively discharged instead of Mg^{2+} ions at the cathode)



(4OH^- ions selectively discharged instead of SO_4^{2-} ions at the anode)

IV. Name the products of electrolysis of Magnesium sulphate(VI) solution

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound)

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

The **four**(4) electrons donated/lost by OH^- ions to form **1** molecule/**1**volume/**1**mole of **oxygen** (O_2)gas at the **anode** are gained/acquired/accepted by the four $\text{H}^+(\text{aq})$ ions to form **2** molecule/**2**volume/**2**mole of **Hydrogen** (H_2)gas at the **cathode**.

The volume of **Oxygen** gas at the anode is thus a **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is thus a **twice** the volume of **Oxygen** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of Magnesium sulphate(VI) solution

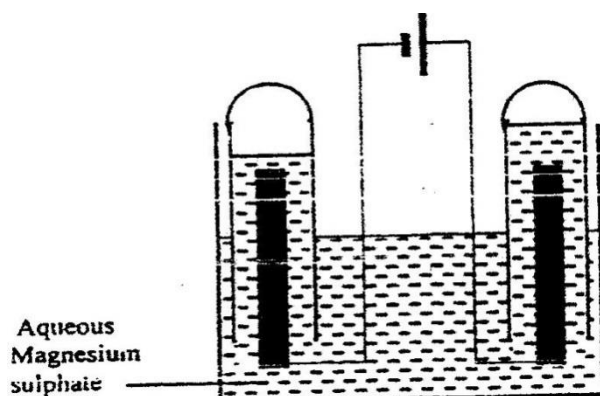
The concentration of dilute Magnesium sulphate(VI) solution **increases**.

The ratio of $\text{H}_2(\text{g})$: $\text{O}_2(\text{g})$ is **2:1** as they are combined in water.

Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape as products.

The concentration /mole of acid present in a given volume of Magnesium sulphate(VI) solution thus continue increasing/rising.

The set – up below was used during the electrolysis of aqueous magnesium sulphate using inert electrodes.



Name a suitable pair of electrodes for this experiment

Identify the ions and cations in the solution

On the diagram label the cathode

Write ionic equations for the reactions that took place at the anode.

Explain the change that occurred to the concentration of magnesium sulphate solution during the experience.

During the electrolysis a current of 2 amperes was passed through the solution for 4 hours. Calculate the volume of the gas produced at the anode. (1 faraday 96500 coulombs and volume of a gas at room temperature is 24000cm³)

One of the uses of electrolysis is electroplating

What is meant by electroplating?

Give two reasons why electroplating is necessary.

b) Concentration of the electrolytes

1. **High** concentrations of cations and/or anions at the electrodes **block** the ion/s that is likely to be discharged at the electrode. This is called **over voltage**. A concentrated solution therefore produces different products of electrolysis from a dilute one.

2. The following experiments show the influence/effect of concentration of electrolyte on the products of electrolysis.

(i) **Electrolysis of dilute and concentrated (brine) sodium chloride solution**

I. Dissolve about 0.5 g of pure sodium chloride crystals in 100cm³ of water. Place the solution in an electrolytic cell. Note the observations at each electrode for 10 minutes. Transfer the set up into a **fume chamber/open** and continue to make observations for a further 10 minute.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

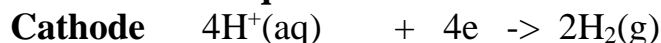


II. Name the ions in sodium chloride solution that are attracted/move to:

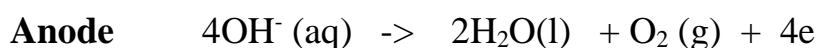
Cathode- $\text{Na}^+(\text{aq})$ from Sodium chloride solution (NaCl) and $\text{H}^+(\text{aq})$ from water (H_2O)

Anode- $\text{Cl}^-(\text{aq})$ from sodiumchloride solution (NaCl) and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:



(H^+ ions selectively discharged instead of Na^+ ions at the cathode)



(4OH^- ions selectively discharged instead of Cl^- ions at the anode)

IV. Name the products of electrolysis of dilute sodium chloride solution

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound)

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

Four(4) electrons donated/lost by OH^- ions to form **1** molecule/1volume/1mole of **oxygen (O_2)**gas at the **anode** are gained/acquired/accepted by four $\text{H}^+(\text{aq})$ ions to form **2** molecule/2volume/2mole of **Hydrogen (H_2)**gas at the **cathode**.

The volume of **Oxygen** gas at the anode is **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **twice** the volume of **Oxygen** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of sodium chloride solution

The concentration of dilute sodium chloride solution **increases**.

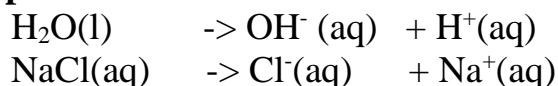
The ratio of **$\text{H}_2(\text{g})$: $\text{O}_2(\text{g})$** is **2:1** as they are combined in water. Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape as products.

The concentration /moles of salt present in a given volume of sodium chloride solution continue increasing/rising.

II. Dissolve about 20 g of pure sodium chloride crystals in 100cm³ of water. Place the solution in an electrolytic cell. Note the observations continuously at each electrode for 30 minutes in a fume chamber/open.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

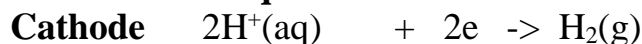


II. Name the ions in sodium chloride solution that are attracted/move to:

Cathode- $\text{Na}^+ \text{(aq)}$ from Sodium chloride solution (NaCl) and $\text{H}^+ \text{(aq)}$ from water (H_2O)

Anode- $\text{Cl}^- \text{(aq)}$ from sodium chloride solution (NaCl) and $\text{OH}^- \text{(aq)}$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:



H^+ ions selectively discharged instead of Na^+ ions at the cathode)



(Cl^- ions with a **higher** concentration **block** the discharge of OH^- ions at the anode)

IV. Name the products of electrolysis of concentrated sodium chloride solution/brine

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode-Chlorine gas(pale green gas that bleaches damp/moist/wet litmus papers)

V. Explain the difference in volume of products at the cathode and anode.

Two (2) electrons donated/lost by Cl^- ions to form **1** molecule/1volume/1mole of **Chlorine** (Cl_2)gas at the **anode** are gained/acquired/accepted by two $\text{H}^+ \text{(aq)}$ ions to form **1** molecule/1volume/1mole of **Hydrogen** (H_2)gas at the **cathode**.

The volume of **Chlorine** gas at the anode is **equal to** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **equal to** the volume of **Chlorine** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of concentrated sodium chloride solution/brine

The concentration of concentrated sodium chloride solution/brine **increases**.

The ratio of $\text{Cl}_2 \text{(g)}$: $\text{H}_2 \text{(g)}$ is **1:1** as they are combined in water.

Water in the electrolyte is decomposed into only Hydrogen gas that escapes as products at cathode.

The concentration /moles of OH^- (aq) and Na^+ ion (as NaOH) present in a given volume of electrolyte continue increasing/rising.

This makes the electrolyte strongly alkaline with **high pH**.

As the electrolysis of brine continues the concentration of Cl^- ions **decrease** and **oxygen** gas start being **liberated** at **anode**.

The electrolyte pH is thus lowered and the concentration of brine starts again increasing.

(ii)Electrolysis of dilute and concentrated Hydrochloric acid solution

I. Prepare about 50cm³ of 0.05 M of dilute Hydrochloric acid in 100cm³ solution. Place the solution in an electrolytic cell. Note the observations at each electrode for 10 minutes.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

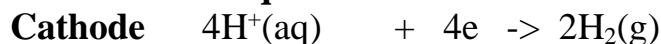


II. Name the ions in dilute Hydrochloric acid solution that are attracted/move to:

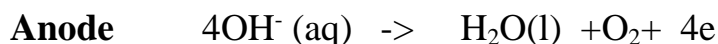
Cathode- $\text{H}^+(\text{aq})$ from dilute Hydrochloric acid (HCl) and $\text{H}^+(\text{aq})$ from water (H_2O)

Anode- $\text{Cl}^-(\text{aq})$ from dilute Hydrochloric acid (HCl) and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:



H^+ ions selectively discharged instead of Na^+ ions at the cathode)



(4OH^- ions selectively discharged instead of Cl^- ions at the anode)

IV. Name the products of electrolysis of dilute Hydrochloric acid

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

Four(4) electrons donated/lost by OH^- ions to form **1** molecule/1volume/1mole of **oxygen (O_2)** gas at the **anode** are gained/acquired/accepted by four $\text{H}^+(\text{aq})$ ions to form **2** molecule/2volume/2mole of **Hydrogen (H_2)** gas at the **cathode**.

The volume of **Oxygen** gas at the anode is **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **twice** the volume of **Oxygen** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of dilute Hydrochloric acid

The concentration of dilute Hydrochloric acid **increases**.

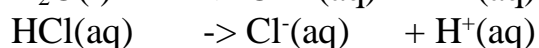
The ratio of H_2 (g): O_2 (g) is **2:1** as they are combined in water. Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape as products.

The concentration /moles of HCl present in a given volume of dilute Hydrochloric acid continue increasing/rising.

II. Prepare about 50cm³ of 2M of Hydrochloric acid in 100cm³ solution. Place the solution in an electrolytic cell. Note the observations at each electrode for 30 minutes
Caution This experiment should be done in the open/fume chamber.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

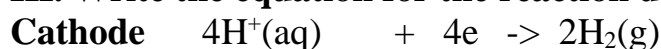


II. Name the ions in 2M Hydrochloric acid solution that are attracted/move to:

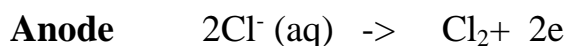
Cathode- $\text{H}^+(\text{aq})$ from dilute Hydrochloric acid (HCl) and $\text{H}^+(\text{aq})$ from water (H_2O)

Anode- $\text{Cl}^-(\text{aq})$ from dilute Hydrochloric acid (HCl) and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:



H^+ ions selectively discharged instead of Na^+ ions at the cathode)



(OH^- ions concentration is **low**. Cl^- ions concentration is **higher** at the anode thus cause **over voltage/block** discharge of OH^- ions)

IV. Name the products of electrolysis of 2M Hydrochloric acid

Cathode- Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode- Chlorine gas (Pale green gas that bleaches blue/red moist/wet/damp litmus papers)

V. Explain the difference in volume of products at the cathode and anode.

Two(2) electrons donated/lost by Cl^- ions to form **1** molecule/**1** volume/**1** mole of **Chlorine** (Cl_2) gas at the **anode** are gained/acquired/accepted by two $\text{H}^+(\text{aq})$ ions to form **1** molecule/**1** volume/**1** mole of **Hydrogen** (H_2) gas at the **cathode**.

The volume of **Chlorine** gas at the anode is **equal to** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **twice** the volume of **Chlorine** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of 2M Hydrochloric acid

The concentration of Hydrochloric acid **decreases**.

The ratio of H_2 (g): Cl_2 (g) is **1:1** as they are combined in Hydrochloric acid.

Water in the electrolyte is decomposed only into Hydrogen gas that escapes as products at the cathode.

There is a net accumulation of excess OH^- (aq) ions in solution.

This makes the electrolyte strongly alkaline with high pH.

c) Nature of electrodes used in the electrolytic cell

Inert electrodes (carbon-graphite and platinum) **do not** alter the expected products of electrolysis in an electrolytic cell. If another/different electrode is used in the electrolytic cell it alters/influences/changes the expected products of electrolysis.

The examples below illustrate the influence of the nature of electrode on the products of electrolysis:

(i) **Electrolysis of copper(II) sulphate(VI) solution**

I. Using carbon-graphite electrodes

Weigh Carbon -graphite electrodes. Record the masses of the electrodes in table I below. Place the electrodes in 1M copper(II) sulphate(VI) solution in a beaker. Set up an electrolytic cell.

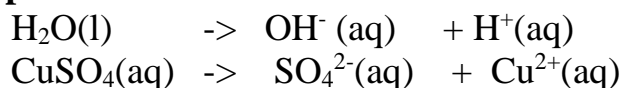
Close the switch and pass current for about 20 minutes. Observe each electrode and any changes in electrolyte. Remove the electrodes from the electrolyte. Wash with acetone/propanone and allow them to dry. Reweigh each electrode.

Sample results

Mass of cathode before electrolysis	23.4 g	Mass of anode before electrolysis	22.4 g
Mass of cathode after electrolysis	25.4 g	Mass of anode after electrolysis	22.4 g
Brown solid deposit at the cathode after electrolysis	-	Bubbles of colourless gas that relight splint	-
Blue colour of electrolyte fades /become less blue	-	Blue colour of electrolyte fades /become less blue	-

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.



II. Name the ions in 1M copper(II) sulphate(VI) solution that are attracted/move to:

Cathode- Cu^{2+} (aq) from copper(II) sulphate(VI) solution and H^+ (aq) from water (H_2O)

Anode- SO_4^{2-} (aq) from copper(II) sulphate(VI) solution and OH^- (aq) from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $2\text{Cu}^{2+}(\text{aq}) + 4\text{e}^- \rightarrow 2\text{Cu}(\text{s})$

Cu^{2+} ions are **lower** than H^+ ions in the electrochemical series therefore selectively discharged at the cathode.)

Anode $4\text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + 4\text{e}^-$

(OH^- ions are **higher** than SO_4^{2-} ions in the electrochemical series therefore selectively discharged at the cathode.)

IV. Name the products of electrolysis of 1M copper(II) sulphate(VI) solution

Cathode- 2 moles of **copper** metal as brown solid coat

Anode- Oxygen gas (Colourless gas that relights /rekindles glowing splint)

V. Explain the changes that take place at the cathode and anode.

Four(4) electrons donated/lost by OH^- ions to form **1** molecule/**1** volume/**1** mole of **Oxygen** (O_2) gas at the **anode** are gained/acquired/accepted by two Cu^{2+} (aq) ions to form **2** moles of **brown copper** solid that deposit itself at the **cathode**.

The moles of **oxygen** gas at the anode is **equal to** the moles of **copper** produced at the cathode

VI. Explain the changes in electrolyte during electrolysis of 1M copper (II) sulphate(VI) solution.

(i) The **pH** of copper(II) sulphate(VI) solution lowers/**decreases**. The salt becomes **more acidic**. Water in the electrolyte is decomposed only into Oxygen gas (from the OH^- ions) that **escapes** as products at the **anode**. There is a net accumulation of **excess** H^+ (aq) ions in solution. This makes the electrolyte strongly **acidic** with **low** pH.

(ii) Cu^{2+} (aq) ions are responsible for the **blue** colour of the electrolyte/ copper(II) sulphate (VI) solution. As electrolysis continues, **blue** Cu^{2+} (aq) ions gain electrons to form **brown Copper**. The blue colour of electrolyte therefore fades/become less blue.

(iii) Copper is deposited at the cathode. This increases the mass of the cathode. OH^- ions that produce Oxygen gas at anode come from water. Oxygen escapes out/away without increasing the mass of anode.

II. Using copper electrodes

Weigh clean copper plates electrodes. Record the masses of the electrodes in table I below. Place the electrodes in 1M copper(II) sulphate(VI) solution in a beaker. Set up an electrolytic cell.

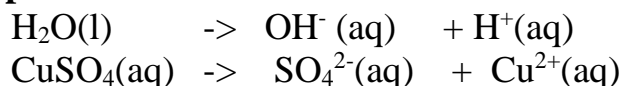
Close the switch and pass current for about 20 minutes. Observe each electrode and any changes in electrolyte. Remove the electrodes from the electrolyte. Wash with acetone/propanone and allow them to dry. Reweigh each electrode.

Sample results

Mass of cathode before electrolysis	23.4 g	Mass of anode before electrolysis	22.4 g
Mass of cathode after electrolysis	25.4 g	Mass of anode after electrolysis	20.4 g
Brown solid deposit at the cathode after electrolysis	-	Anode decrease in size/erodes/wear off	-
Blue colour of electrolyte remain blue	-	Blue colour of electrolyte remain blue	-

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

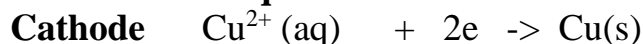


II. Name the ions in 1M copper(II) sulphate(VI) solution that are attracted/move to:

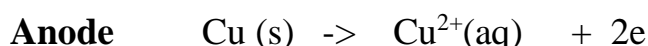
Cathode- $\text{Cu}^{2+}(\text{aq})$ from copper(II) sulphate(VI) solution and $\text{H}^+(\text{aq})$ from water (H_2O)

Anode- $\text{SO}_4^{2-}(\text{aq})$ from copper(II) sulphate(VI) solution and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:



Cu^{2+} ions are **lower** than H^+ ions in the electrochemical series therefore selectively discharged at the cathode.)



(Both OH^- ions and SO_4^{2-} ions move to the anode but none is discharged. The copper anode itself ionizes/dissolves/dissociate because less energy is used to remove an electron/ionize /dissociate copper atoms than OH^- ions.

IV. Name the products of electrolysis of 1M copper(II) sulphate(VI) solution using copper electrodes.

Cathode-1 moles of **copper** metal as brown solid coat (Cathode increase/deposits)

Anode-Anode erodes/decrease in size

V. Explain the changes that take place during the electrolytic process

(i)Cathode

-**Cu²⁺** ions are **lower** than **H⁺** ions in the electrochemical series therefore selectively discharged at the cathode. **Cu²⁺** ions have greater tendency to accept/gain/acquire electrons to form brown **copper** atoms/solid that deposit itself and increase the mass/size of the **cathode**.The copper deposited at the cathode is **pure**

-**H⁺** ions accumulate around the cathode. Electrolyte thus becomes strongly acidic around the cathode.

-**Cu²⁺** ions in solution are responsible for the blue colour of electrolyte. Blue colour of electrolyte **fade** around the cathode.

(ii)Anode

Copper atom at the anode easily ionizes to release electrons. The anode therefore keeps decreasing in mass/eroding. The amount of copper that dissolve/erode is **equal** to the mass of copper deposited. This is called **electrode ionization**.

Electrode ionization is where the anode erodes/decrease and the cathode deposits/increase during electrolysis. The overall **concentration** of the electrolyte remains **constant**

14.In industries electrolysis has the following uses/applications:

(a)Extraction of reactive metals from their ores.

Potassium, sodium ,magnesium, and aluminium are extracted from their ores using electrolytic methods.

(b)Purifying copper after extraction from copper pyrites ores.

Copper obtained from copper pyrites ores is not pure. After extraction, the copper is refined by electrolysing copper(II)sulphate(VI) solution using the **impure** copper as **anode** and a thin strip of **pure** copper as **cathode**. Electrode ionization take place there:

(i)At the cathode; **Cu²⁺ (aq) + 2e -> Cu(s)** (Pure copper deposits on the strip

(ii)At the anode; **Cu(s) ->Cu²⁺ (aq) + 2e** (impure copper erodes/dissolves)

(c)Electroplating

The label EPNS(**E**lectro **P**lated **N**ickel **S**ilver) on some steel/metallic utensils mean they are plated/coated with silver and/or Nickel to **improve** their **appearance**(add their **aesthetic** value)and **prevent**/slow **corrosion**(**rusting** of iron). Electroplating is the process of coating a metal with another metal using an electric current. During electroplating,the **cathode** is made of the metal to be **coated**/impure.

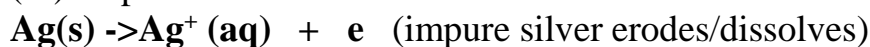
Example:

During the electroplating of a spoon with silver

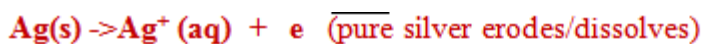
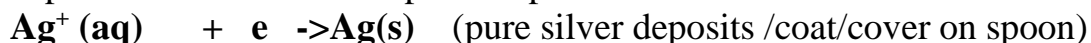
(i) the spoon/impure is placed as the cathode (negative terminal of battery)

(ii) the pure silver is placed as the anode (positive terminal of battery)

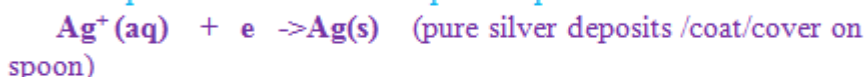
(iii) the pure silver erodes/ionizes/dissociates to release electrons:



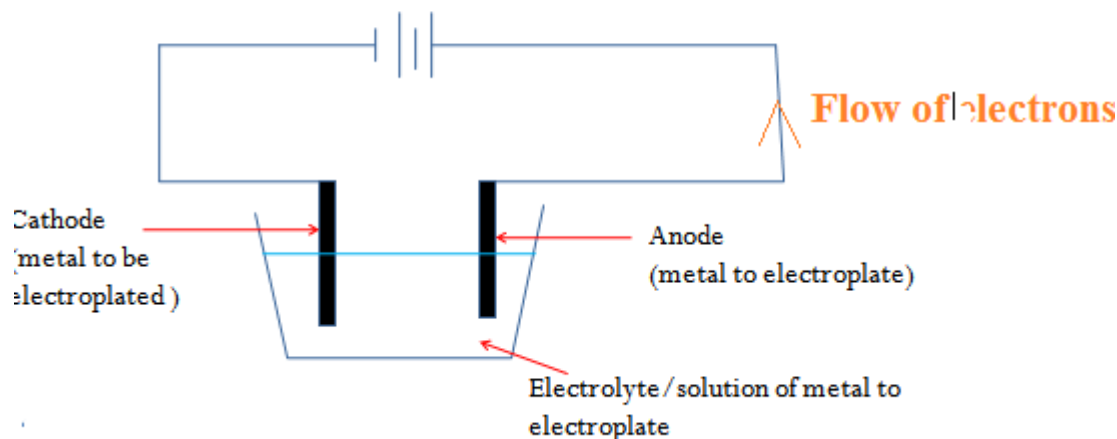
(iv) silver (Ag^+) ions from electrolyte gain electrons to form pure silver deposits / coat / cover the spoon/impure



(iv) silver (Ag^+) ions from electrolyte gain electrons to form pure silver deposits / coat / cover the spoon/impure



Electrolytic set up during electroplating



15. The quantitative amount of products of electrolysis can be determined by applying Faradays 1st law of electrolysis.

Faradays 1st law of electrolysis states that “the mass/amount of substance liberated/produced/used during electrolysis is directly proportional to the quantity of of electricity passed/used.”

(a) The SI unit of quantity of electricity is the coulomb(C). The coulomb may be defined as the quantity of electricity passed/used when a current of one ampere flow for one second.i.e;

$$1\text{Coulomb} = 1\text{Ampere} \times 1\text{Second}$$

The Ampere is the SI unit of current(I)

The Second is the SI unit of time(t) therefore;

Quantity of electricity(in Coulombs) = Current(I) x time(t)

Practice examples

1. A current of 2 amperes was passed through an electrolytic cell for 20 minutes. Calculate the quantity of electric charge produced.

Working:

$$\begin{aligned}\text{Quantity of electricity(in Coulombs)} &= \text{Current(I)} \times \text{time(t)} \\ \text{Substituting /converting time to second} &= 2 \times (20 \times 60) \\ &= \underline{\underline{2400 \text{ C}}}\end{aligned}$$

2. A current of 2 amperes was passed through an electrolytic.96500 coulombs of charge were produced. Calculate the time taken.

Working:

$$\begin{aligned}\text{Time(t) in seconds} &= \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Current(I) in amperes}} \\ \text{Substituting} &= \frac{96500}{2} \\ &= \underline{\underline{48250 \text{ seconds}}}\end{aligned}$$

3. 96500 coulombs of charge were produced after 10 minutes in an electrolytic cell . Calculate the amount of current used.

Working:

$$\begin{aligned}\text{Current(I) in amperes} &= \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Time(t) in seconds}} \\ \text{Substituting/converting time to second=} &= \frac{96500}{10 \times 60} \\ &= \underline{\underline{160.8333 \text{ Amperes}}}\end{aligned}$$

(b)The quantity of electricity required for **one mole of electrons** at the anode/cathode is called the **Faraday constant(F)**. It is about **96500 Coulombs**.i.e

The number of **Faradays** used /required is equal to the number of **electrons** used at cathode/anode during the electrolytic process. e.g.

Cu^{2+} require to gain **2** moles of electrons=**2** Faradays =**2** x 96500 coulombs of electricity at the cathode.

Al^{3+} require to gain **3** moles of electrons=**3** Faradays =**3** x 96500 coulombs of electricity at the cathode

Na^+ require to gain **1** moles of electrons=**1** Faradays =**1** x 96500 coulombs of electricity at the cathode

2H^+ require to gain **2** moles of electrons=**2** Faradays =**2** x 96500 coulombs of electricity at the cathode to form 1 molecule of hydrogen gas

2O^{2-} require to lose/donate **4** moles of electrons=**4** Faradays =**4** x 96500 coulombs of electricity at the anode to form 1 molecule of Oxygen O_2 gas.

4OH^- require to lose/donate **4** moles of electrons=**4** Faradays =**4** x 96500 coulombs of electricity at the anode to form 1 molecule of Oxygen gas and 2 molecules of water.

(c) The mass/amount of products at the cathode/anode is related to the molar mass of the substance and/or the volume of gases at standard/room temperature and pressure as in the below examples:

Practice examples

1. Calculate the mass of copper deposited at the cathode when a steady current of 4.0 amperes is passed through copper(II)sulphate(VI) for 30 minutes in an electrolytic cell. (Cu=63.5, 1F = 96500C)

Working:

$$\begin{aligned} \text{Quantity of electricity (in Coulombs)} &= \text{Current(I)} \times \text{time(t)} \\ \text{Substituting /converting time to second} &= 4 \times (30 \times 60) \\ &= \underline{\underline{7200 \text{ C}}} \end{aligned}$$



2 mole of electrons = 2 Faradays = 2 x 96500 C produce a mass = molar mass of copper thus;

$$\begin{aligned} 2 \times 96500\text{C} &\rightarrow 63.5 \text{ g} \\ 72000\text{C} &\rightarrow \frac{7200 \times 63.5}{2 \times 96500} = \underline{\underline{2.3689 \text{ g}}} \text{ of copper} \end{aligned}$$

2.a) If 3.2 g of Lead were deposited when a current of 2.5 amperes was passed through an electrolytic cell of molten Lead(II)bromide for 20 minutes, determine the Faraday constant. (Pb = 207)

Working:

$$\begin{aligned} \text{Quantity of electricity (in Coulombs)} &= \text{Current(I)} \times \text{time(t)} \\ \text{Substituting /converting time to second} &= 2.5 \times (20 \times 60) \\ &= \underline{\underline{3000 \text{ C}}} \end{aligned}$$

$$\begin{aligned} \text{If 3.2g of Lead} &\rightarrow 3000\text{C} \\ \text{Then 207 g of Lead} &\rightarrow \frac{207 \times 3000}{3.2} = \underline{\underline{194062.5 \text{ C}}} \end{aligned}$$



$$\begin{aligned} \text{From the equation: 2 moles of electrons} &= 2 \text{ Faradays} = 194062.5 \text{ C} \\ \text{1mole of electrons} &= 1 \text{ Faraday} \Rightarrow \frac{194062.5}{2} = \underline{\underline{97031.25 \text{ C}}} \end{aligned}$$

b) What is the volume of bromine vapour produced at the anode at room temperature (1mole of gas at room temperature and pressure = 24000cm³)

Method 1



$$\begin{aligned} \text{From the equation: } 2 \text{ moles of electrons} &= 2 \text{ Faradays} = 194062.5 \text{ C} \rightarrow 24000 \text{ cm}^3 \\ 3000 \text{ C} &\rightarrow \frac{3000 \times 24000}{194062.5} \\ &= \mathbf{371.0145 \text{ cm}^3} \end{aligned}$$

Method 2



Mole ratio of products at Cathode: anode = 1:1

Moles of Lead at cathode = $\frac{3.2}{207} = 0.0155 \text{ moles} = \text{moles of Bromine}$

1 moles of bromine vapour $\rightarrow 24000 \text{ cm}^3$

0.0155 moles of Bromine $\rightarrow 0.0155 \times 24000 = \mathbf{372 \text{ cm}^3}$

Method 3



Ratio of Faradays used to form products at Cathode: anode = 2:2

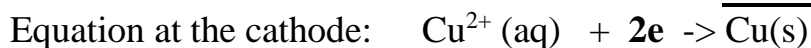
$\Rightarrow 2 \times 97031.25 \text{ C produce } 24000 \text{ cm}^3 \text{ of bromine vapour}$

Then: $3000 \text{ C} \rightarrow \frac{3000 \times 24000 \text{ cm}^3}{2 \times 97031.25} = \mathbf{371.0145 \text{ cm}^3}$

3. What mass of copper remain from 2.0 at the anode if a solution of copper(II)sulphate(VI) is electrolysed using a current of 1 ampere flowing through an electrolytic cell for 20 minutes. (Cu= 63.5, 1Faraday = 96487 coulombs)

Working:

Quantity of electricity (in Coulombs) = Current(I) x time(t)
 Substituting /converting time to second = $1 \times (20 \times 60)$
 = $\mathbf{1200 \text{ C}}$



2 mole of electrons = 2 Faradays = $2 \times 96500 \text{ C}$ erode/dissolve a mass = molar mass of copper thus;

$$\begin{aligned} 2 \times 96500 \text{ C} &\rightarrow 63.5 \text{ g} \\ 1200 \text{ C} &\rightarrow \frac{1200 \times 63.5}{2 \times 96500} = \mathbf{0.3948 \text{ g}} \text{ of copper deposited} \end{aligned}$$

Mass of copper remaining = Original mass – mass dissolved/eroded

$\Rightarrow 2.0 - 0.3948 = \mathbf{1.6052 \text{ g}}$ of copper remain

4. Calculate the current passed if a mass of 0.234 g of copper is deposited in 4 minutes during electrolysis of a solution of copper (II)sulphate(VI). (Cu= 63.5 ,1F = 96500C)

Working:



2 mole of electrons = 2 Faradays = $2 \times 96500 \text{ C}$ produce a mass = molar mass of copper thus;

$$63.5 \text{ g} \rightarrow 2 \times 96500 \text{ C}$$

$$0.234 \text{ g} \rightarrow \frac{0.234 \times 2 \times 96500}{63.5} = \underline{711.2126 \text{ C}}$$

$$\text{Current(I) in amperes} = \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Time(t) in seconds}}$$

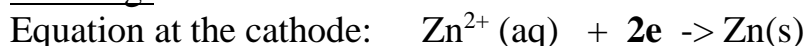
$$\text{Substituting/convertng time to second} = \frac{711.2126 \text{ C}}{4 \times 60}$$

$$= \underline{2.9634 \text{ Amperes}}$$

5. (a) What quantity of electricity will deposit a mass of 2.43 g of Zinc during electrolysis of a solution of Zinc (II)sulphate(VI).

(Zn= 65 ,1F = 96500C)

Working:



2 mole of electrons = 2 Faradays = 2 x 96500 C erode/dissolve a mass =molar mass of Zinc thus;

$$65 \text{ g} \rightarrow 2 \times 96500$$

$$2.43 \text{ g} \rightarrow \frac{2.43 \times 2 \times 96500}{65} = \underline{7215.2308 \text{ C}}$$

(b) Calculate the time (in minutes) it would take during electrolysis of the solution of Zinc (II)sulphate(VI) above if a current of 4.0 Amperes is used.

$$\text{Time(t) in seconds} = \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Current(I) in amperes}}$$

$$\text{Substituting} = \frac{7215.2308}{4} = \frac{1803.8077 \text{ seconds}}{60} = \underline{30.0635 \text{ minutes}}$$

6. When a current of 1.5 amperes was passed through a cell containing M^{3+} ions of metal M for 15 minutes, the mass at cathode increased by 0.26 g.(Faraday constant = 96500C

a) Calculate the quantity of electricity used.

$$\text{Quantity of electricity (in Coulombs)} = \text{Current(I)} \times \text{time(t)}$$

$$\text{Substituting /convertng time to second} = 1.5 \times (15 \times 60)$$

$$= \underline{1350 \text{ C}}$$

b) Determine the relative atomic mass of metal M



$$1350 \text{ C of electricity} \rightarrow 0.26 \text{ g of metal M}$$

3 mole of electrons = 3 Faradays = 3 x 96500 C produce a mass =molar mass of M thus;

$$\text{RAM of M} = \frac{0.26 \text{ g} \times 3 \times 96500}{1350} = \underline{55.7556} (\text{No units})$$

7. An element "P" has a relative atomic mass 88. When a current of 0.5 amperes was passed through fused chloride of "P" for 32 minutes and 10seconds ,0.44 g

of "P" was deposited at the cathode. Determine the charge on an ion of "P"(Faraday constant = 96500C)

Working:

$$\begin{aligned} \text{Quantity of electricity (in Coulombs)} &= \text{Current(I)} \times \text{time(t)} \\ \text{Substituting /converting time to second} &= 0.5 \times ((32 \times 60) + 10) \\ &= \underline{965C} \end{aligned}$$

0.44 g of metal "P" are deposited by 965C

$$88 \text{ g of metal "P" are deposited by: } \frac{88 \times 965}{0.44} = \underline{193000 C}$$

96500 C = 1 mole of electrons = 1 Faradays = single charge

$$193000 C \quad \rightarrow \quad \frac{193000}{96500} = \underline{2} \text{ moles/Faradays/charges} \Rightarrow \text{symbol of ion} = \text{P}^{2+}$$

8. During purification of copper by electrolysis 1.48 g of copper was deposited when a current was passed through aqueous copper (II)sulphate(VI) for 2 ½ hours. Calculate the amount of current that was passed. (Cu= 63.5 ,1F = 96500C)

Working:

Equation at the cathode: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$

2 mole of electrons = 2 Faradays = 2 x 96500 C produce a mass =molar mass of copper thus;

$$\begin{aligned} 63.5 \text{ g} &\rightarrow 2 \times 96500\text{C} \\ 1.48 \text{ g} &\rightarrow \frac{1.48 \times 2 \times 96500}{63.5} = \underline{4255.1181 C} \end{aligned}$$

$$\text{Current(I) in amperes} = \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Time(t) in seconds}}$$

$$\begin{aligned} \text{Substituting/converting time to second=} & \frac{4255.1181\text{C}}{((2 \times 60) + 30) \times 60} \\ &= \underline{0.4728 \text{ Amperes}} \end{aligned}$$

17. Practically Faraday 1st law of electrolysis can be verified as below.

Verifying Faraday 1st law of electrolysis

Procedure.

Weigh clean copper plates electrodes. Record the masses of the electrodes in table I below. Place the electrodes in 1M copper(II) sulphate(VI) solution in a beaker. Set up an electrolytic cell.

Close the switch and pass a steady current of 2 amperes by adjusting the rheostat for exactly 20 minutes. Remove the electrodes from the electrolyte. Wash with acetone/ propanone and allow them to dry. Reweigh each electrode.

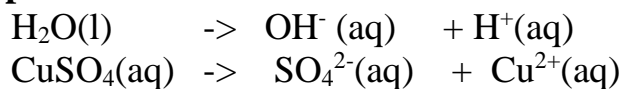
Sample results

Mass of cathode before electrolysis	7.00 g	Mass of anode before electrolysis	7.75 g
Mass of cathode after	8.25 g	Mass of anode after	6.50 g

electrolysis		electrolysis	
Change in mass at cathode after electrolysis	1.25 g	Change in mass at anode after electrolysis	1.25 g

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.



II. Name the ions in 1M copper(II) sulphate(VI) solution that are attracted/move to:

Cathode- $\text{Cu}^{2+}(\text{aq})$ from copper(II) sulphate(VI) solution and $\text{H}^+(\text{aq})$ from water (H_2O)

Anode- $\text{SO}_4^{2-}(\text{aq})$ from copper(II) sulphate(VI) solution and $\text{OH}^-(\text{aq})$ from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

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

Cu^{2+} ions are **lower** than H^+ ions in the electrochemical series therefore selectively discharged at the cathode.)

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

(Both OH^- ions and SO_4^{2-} ions move to the anode but none is discharged. The copper anode itself ionizes/dissolves/dissociate as less energy is used to remove an electron/ionize /dissociate copper atoms than OH^- ions.

IV. Name the products of electrolysis of 1M copper(II) sulphate(VI) solution using copper electrodes.

Cathode- 1.25 g of **copper** metal as brown solid coat/deposits

Anode- 1.25 g of **copper** metal erodes/decrease in size

V. (i)How many moles of electrons are used to deposit/erode one mole of copper metal at the cathode/anode?

From the equation at anode/cathode= 2 moles

(ii)How many Faradays are used to deposit/erode one mole of copper metal at the cathode/anode?

From the equation at anode/cathode : 2 moles = 2 Faradays

(iii)Calculate the quantity of electric charge used

Working:

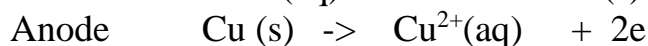
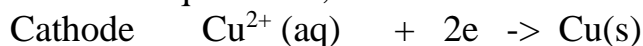
$$\begin{aligned} \text{Quantity of electricity (in Coulombs)} &= \text{Current(I)} \times \text{time(t)} \\ \text{Substituting /converting time to second} &= 2 \times 20 \times 60 \\ &= \underline{2400\text{C}} \end{aligned}$$

VI. (i) Calculate the quantity of electricity required to deposit/erode one mole of copper at the cathode/anode(Cu=63.5)

$$\begin{aligned} \text{Since } 1.25 \text{ g of copper} &\rightarrow 2400\text{C} \\ \text{Then } 63.5 \text{ g (1mole of copper)} &\rightarrow \frac{63.5 \times 2400}{1.25} = \underline{121920 \text{ C}} \end{aligned}$$

(ii) Determine the Faraday constant from the results in V(i) above

From the equation at;



$$2 \text{ moles} = 2 \text{ Faradays} \rightarrow 121920 \text{ C}$$

$$1 \text{ moles} = 1 \text{ Faradays} \rightarrow \frac{121920}{2} = \underline{60960 \text{ C}}$$

(iii) The faraday constant obtained above is far lower than theoretical. Explain

- high resistance of the wires used.
- temperatures at 25°C were not kept constant
- plates/electrodes used were not made of pure copper
- plates/electrodes used were not thoroughly clean copper

Further practice

1. An element P has a relative atomic mass of 88. When a current of 0.5 amperes was passed through the fused chloride of P for 32 minutes and 10 seconds, 0.44g of P were deposited at the cathode. Determine the charge on an ion of P. (1 faraday = 96500 Coulombs).

2. During electrolysis of aqueous copper (II) sulphate, 144750 coulombs of electricity were used. Calculate the mass of copper metal that was obtained (Cu = 64 ; 1 Faraday = 96500 coulombs) (3 mks)

3. A nitrate of a metal M was electrolysed .1.18 g of metal was deposited when a current of 4 amperes flow for 16 minutes. Determine the formula of the sulphate(VI)salt of the metal.

(Faraday constant = 96500 , RAM of X = 59.0)

Working

$$Q = It \Rightarrow (4 \times 16 \times 60) = 3840 \text{ C}$$

$$1.18 \text{ g of X} \Rightarrow 3840 \text{ C}$$

$$59.0 \text{ g} \Rightarrow \frac{59.0 \times 3840}{1.18} = 192000 \text{ C}$$

$$96500 \text{ C} = 1 \text{ Faraday}$$

$$192000 \text{ C} = \frac{192000 \text{ C} \times 1}{96500 \text{ C}} = 2\text{F} \text{ thus charge of M} = \text{M}^{2+}$$

Valency of M is 2 thus formula of sulphate(VI)salt MSO_4

4. Below is the results obtained when a current of 2.0amperes is passed through copper(II)sulphate(VI)solution for 15 minutes during electrolysis using copper electrode.

Initial mass of cathode = 1.0 g

Final mass of cathode = 1.6 g

Change in mass of cathode = 0.60 g

(i)Determine the change in mass at the anode. Explain your answer.

Mass decrease = 0.6g.

Electrode ionization take place where the cathode increase in mass form the erosion of the anode

(ii)Calculate the quantity of electricity required to deposit one mole of copper.(Cu =63.5)

$$Q = It \Rightarrow 2 \times 15 \times 60 = 1800 \text{ coulombs}$$

Method 1

$$\begin{array}{l} 0.60 \text{ g of copper} \quad \rightarrow 1800 \text{ coulombs} \\ 63.5 \text{ g} \quad \rightarrow \frac{63.5 \times 1800}{0.60} = \underline{\underline{190500}} \text{ Coulombs} \end{array}$$

Method 2

$$\text{Moles of Copper} = \frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{0.60}{63.5} = \underline{\underline{9.4488 \times 10^{-3} \text{ moles}}}$$

$$\begin{array}{l} 9.4488 \times 10^{-3} \text{ moles} \quad \rightarrow 1800 \text{ coulombs} \\ 1 \text{ Mole} \quad \rightarrow \frac{1 \times 1800 \text{ coulombs}}{9.4488 \times 10^{-3} \text{ moles}} = \underline{\underline{190500.381}} \text{ coulombs} \end{array}$$

(iii)Determine the oxidation number of copper produced at the cathode and hence the formula of its nitrate (V)salt (1 Faraday = 96500 Coulombs)

$$\begin{array}{l} 96500 \text{ Coulombs} \quad \rightarrow 1 \text{ Faraday} \\ 190500.381 \text{ coulombs} \quad \rightarrow \frac{190500.381 \text{ coulombs} \times 1}{96500 \text{ Coulombs}} \end{array}$$

$$= 1.9741 \text{ Faradays} \Rightarrow \underline{\underline{2F}}(\text{whole number})$$

Charge of copper = 2+ = **Oxidation number**

=> Valency of copper = **2** hence chemical formula of nitrate (V)salt = **Cu (NO₃)₂**