Rules for Ox. No.

4. OX. NO. of OXYGEN in all cpd's = -2 EXCEPT in PEROXIDE where it is -1 and OF_2 where it is +2.

5) In cpd's the sum of the ox. no's of all atoms is equal to zero.

eg $H_2O = 2X(+1) + (-2) = 0$

6) In ions containings more than one element, the overall charge is equal to the sum of the ox. no's of the constituent elements.

eg $NH_4^+ = -3 + 4(+1) = +1$ $OH^- = -2 + (+1) = -1$

Examples

(1) Determine the oxidation no of Manganese in MnO_4^-.

(2) Determine the oxidation no of Iron in Fe_2O_3.

KLB pg 100.

NB/ some elements have variable oxygen no's

i.e.

species	ox. No
NO_3^-	+5
NO_2	+4
NO	+2
N_2O	+1
N_2	0
NH_3, NH_4^+, Mg_3N_2	-3

Others redox rxn: $Mg + 2HCl \longrightarrow Mg^{2+}(aq)$ ...

$2FeCl_2(aq) + Cl_2(g) \longrightarrow 2FeCl_3(aq)$

- The qnt. of electricity carried by one mole of electrons is a Constant called Faraday (F) & is equivalent to 96,500 Coulombs.

$$1mol = 1F = 96500\ C$$

Example

1. In an experiment to electrolyse $CuSO_4$ soln. using copper electrodes 0.2 amperes were passed thro' the soln. for 1,930 sec. The mass of copper at the cathode increased from 6.35 to 6.478.
Find the charge on a copper ion. (1F = 96,500C, Cu = 64)

✓ $Q = I \times t$
$= 0.2 \times 1930$
$= 386\ C$

mass of copper deposited = 6.478 – 6.350 = 0.128g

0.128g of copper were deposited by 386 C
64g = ?

✓ $= \left(\frac{64 \times 386}{0.128}\right) = 193,000\ C$

✓ 1F = 96500 C
? ← 193,000 C ✓ $= \frac{193,000}{96500} = 2F$

1F = 1mol of e's

NB/ (iii) Problem with this product

✓ The number of e's required to deposited a gi[...] is equivalent to the charge on the ion.

obtained **purified**. It is evaporated to give **pellets**.

Electrolysis

NB/ (iii) Problem with this product

✓ **Expensive** due to **high cost** of **mercury**.
✓ **Mercury** is **poisonous**.

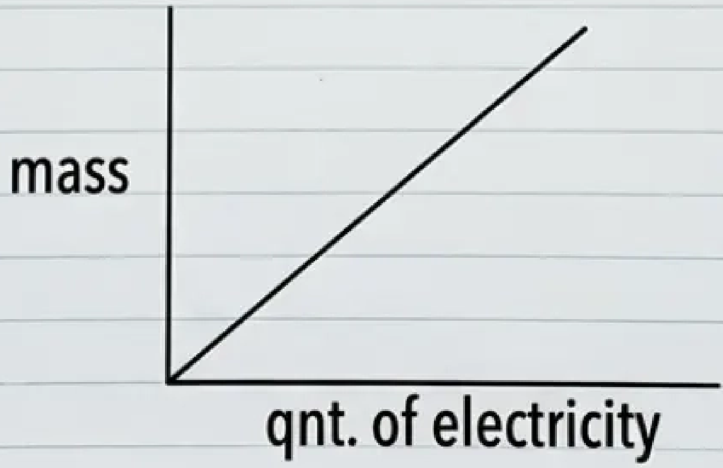
Quantitative treatment of Electrolysis

- The **charge** on an ion depends on its **valency** (the number of electrons lost or gained).

* (•) **The Faraday Constant:** The **qnt. of electricity** needed to produce **one mole** of atoms of any element is always **96,500 Coulombs.**

Faradays law of Electrolysis

Faradays law of Electrolysis: The **mass** is **electricity** to produced(+) erncet of the **current** to the **time**, **charge** and it on the **charge** is a osais.



qnt. of **electricity (Q)** is a coulombe spend where **Amperes (I)**, or **seconds (t)**, in a statent oity of **Coulombs (C)** ⇒ **Colcoulomb**

A **colcoulomb** is a qnt. of **electricity** passed when a **current** (I) of **1 ampere** flows for one **second**

NB/ The ease with which a metal lose e's depends on their position in the reactivity series.

- A metal which looses e's easily will acquire a much more negative change on the surface - vice-versa.

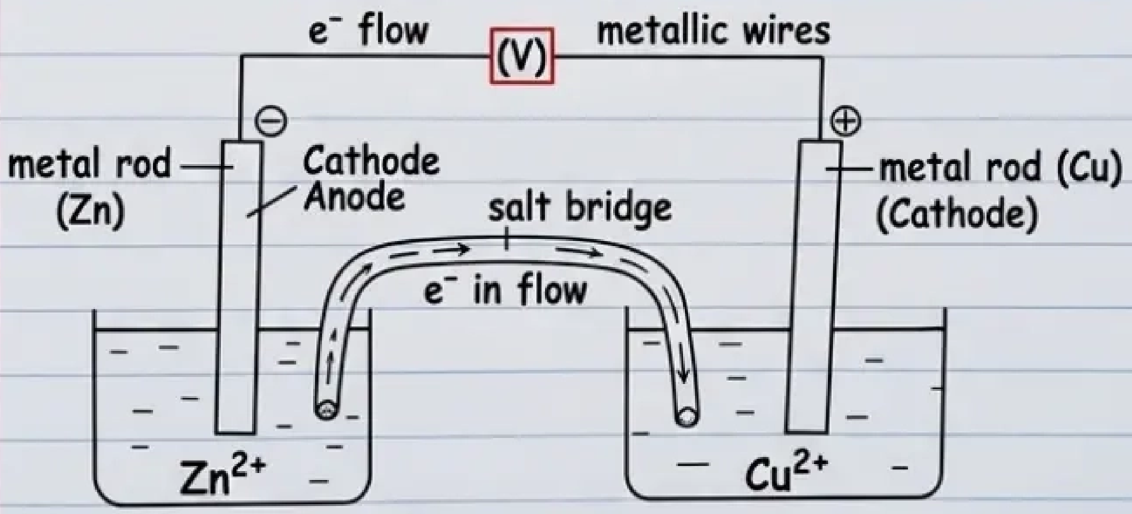
✓ It is not possible to measure the electrode potential of one electrode.

HOWEVER; We can measure the difference in potential of two electrodes.

Where:

The two electrodes are connected by metallic wires then - The solns are connected thro' a salt bridge.

i.e



Electrochemical Cell

NB/ The salt bridge is in form of a filter paper soaked in a saturated soln of KNO_3 or $NaNO_3$.

CAUTION:

The salts chosen for the salt bridge must not react with the salt soln in each half cell.

Imp - Electron will flow from the metal electrode with a higher conc. of electrons to the electrode with a lower conc. of e's.

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

i.e. $M_{(s)} \longrightarrow M^{n+}_{(aq)} + n\,e^-$ where $n = 1, 2, 3 \ldots$

- The **lost** e-s remain on the metal surface, while the metal **develops** a **negative** charge.
- The **negative** charges attracts the ions gains & accept e-s from the metal rod & form atoms once more.

i.e.

$M^{n+}_{(aq)} + n\,e^-s \longrightarrow M_{(s)}$ eg $Zn^{2+}_{(aq)} + 2\,e^-s \longrightarrow Zn_{(s)}$

NB/ The two rxns face place at the same time & rate i.e an **equilibrium** is established.

i.e.

Metal atom $\rightleftharpoons$ Metal ions

eg

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

Now, A **potential difference** is always created/generated btwn the metal rod & the freely charged ions in the solu. This normally occurs whenever there is a separation of +ve & –ve charges.

✓ The conc. of electrons on the metal rod is measured by a **quantities** called **electrode potential.**

✓ This kind of arrn shown by the figure above is called a **half cell.**

Representation:

Metal / metal ion

Where; The **vertical line** rep. change of pha

eg $Zn_{(s)} / Zn^{2+}_{(aq)}$ (Zn half cell) relationship of ... metal lose e's ... negativity ... much more negative change on the surface – vice-versa

more +ve (has highest tendency) — *strongest oxidizing agent*
more –ve (has lowest tendency) to lose e^-s — *strongest reducing agent* oxidation

Platinised - platinum electrode is immersed in 1M H^+ ions solution.

Function of platinised- platinum

1. It provides a large S.A. for dissociation of hydrogen molecules to be adsorbed on the platinum at **equilibrium**.

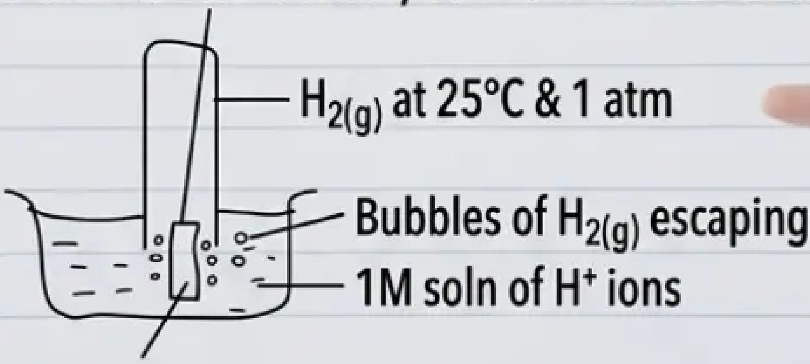
$$\tfrac{1}{2} H_{2(g)} \rightleftharpoons H^+_{(aq)} + e^- \quad E^\circ = 0.00V.$$

2. It is an inert electrode which provides a connection btwn $H_{2(g)} | H^+$.
   Theref; the std. hydrogen half cell is rep. as;

$$Pt_{(s)}, H_{2(g)} | H^+_{(aq)} \therefore \tfrac{1}{2} H_{2(g)} \text{------>} H^+_{(aq)} + e^- \quad E = 0.00V.$$

to allow attainment of equilibrium

3. Acts as an electrical conductor; to the external circuit.



NB/ The **std electrode potential** of any element is taken as the difference btwn its potential & that of **hydrogen**.

✓ IF the electrode has a **greater** tendency to **loose** e^-s than hydrogen then the **electrode potential** of its **half-cell** is **negative** with respect to the **hydrogen half cell**.
Vice versa.

[*]s the +ve eg copper.]

eg Zinc

Ionic half eqns for the rxns

$$Zn_{(s)} \text{---------->} Zn^{2+}_{(aq)} + 2e^- \qquad E^\circ = 0.76V.$$

Combined half eqn

$$Mg_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + Cu_{(s)}$$

(Blue) (Brown)

NB/ The more reactive elements are **stronger reducing agents** since they **loose e's faster**, vice versa.

i.e. The **less reactive metal**, e.g. lead (Pb) & copper (Cu), **loose e's less readily** & are **weakly reducing agents**.

Reactivity Series

Potassium	K	Strongest reducing agent.
Sodium	Na	
Calcium	Ca	
Magnesium	Mg	
Aluminium	Al	Decreasing reducing power ↓
Zinc	Zn	
Iron	Fe	
Lead	Pb	
Copper	Cu	
Silver	Ag	
Gold	Au	Weakest reducing agent.

Other displacement rxns

1. $2Al_{(s)} + 3Cu^{2+}{}_{(aq)} \rightarrow 2Al^{3+}{}_{(aq)} + 3Cu_{(s)}$
2. $Cu_{(s)} + 2Ag^{+}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Ag_{(s)}$

i.e. Asn

NB: Write the **ionic half equations** for each rxn

[E°+ve - E°-ve]

*relation lose e's ... negativity ... much more negative change on the surface - vice-versa

eg $Zn_{(s)}$ ---------> $Zn^{2+}_{(aq)} + 2e^-$ E° = 0.76V.

Cl_2O_7 -under special conditions

i.e.

$M_{(s)}$ -------> $M^{n+}_{(aq)} + n\,e^-$ *where* n = 1, 2, 3 ...

- The lost e-s remain on the metal surface, while the metal **deps a -ve charge.**
- The -ve charges attracts the ions gains & accept e-s from the metal rod & form atoms once more.

i.e. $M^{n+}_{(aq)} + n\,e^-s$ ---> $M_{(s)}$ eg $Zn^{2+}_{(aq)} + 2\,e^-s$ ---> $Zn_{(s)}$

<u>NB/</u> The two rxns face place at the same time & rate i.e an **equilibrium** is established.

i.e. Metal atom $\rightleftharpoons$ Metal ions

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

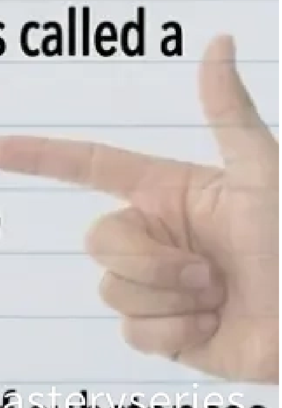
Now, A **potential difference** is always created/generated btwn the metal rod & the freely charged ions in the solu. This normally occurs whenever there is a separation of +ve & -ve charges.

✓ The conc. of e-s on the metal rod is measured by a qut. called **electrode potential.**

✓ This kind of arrn shown by the figure above is called a **half cell.**

Representation: Metal / metal ion

Where; The vertical line rep. change of phase (Zn half cell)



<u>DIAG : KLB - assignment</u>

✓ use of std electrode potentials

① To compare oxidizing & reducing powers of substances

④ To determine emf of a cell

③ To predict whether or not a rxn will take place or not

EXAMPLE : KLB pg 114 - 115

✓ uses of Electrochemical cells

[Above] are used as a source of energy

i.e a) Dry cell

- They are used in a variety of electrical appl[...] radius, watches, clocks, electric bells etc.
- They are cheap & convenient to use bec_3 they contain

Displacement Rxns

Is a rxn in which a **more** reactive element takes over the place of a less reactive element (**metal**).

Case: Rxns btwn. solid samples and aqueous solns.

ie.

	Soln of Metal Ion					
Metal	Ca^{2+}	Mg^{2+}	Zn^{2+}	Fe^{2+}	Pb^{2+}	Cu^{2+}
Calcium	no rxn			no rxn	no rxn	
Magnesium	no rxn		✓ rxn		no rxn	
Zinc	no rxn	✓ rxn		✓ rxn		
Iron	no rxn			✓ rxn		
Lead	no rxn			no rxn	no rxn	
Copper	no rxn		no rxn		no rxn	

It is noted that metals **higher** in the reactivity series displace these metals **lower** in the reactivity series.

Example:

$$Mg_{(s)} + \underset{\text{(Blue)}}{CuSO_{4(aq)}} \longrightarrow MgSO_{4(aq)} + \underset{\text{(Brown)}}{Cu_{(s)}}$$

i.e.

✓ Magnesium looses 2e⁻s to copper **ie** it is oxidized & its ox. state increases from 0 to +2.

$$Mg_{(s)} \longrightarrow Mg^{2+}{}_{(aq)} + 2e^-$$

✓ Copper is then displaced by magnesium ions & the colour of soln. changes from blue to brown **ie** it is redu copper metal.

i.e.

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

*CAUTION: Halogens are poisonous when inhaled and the reactions should be performed in fume cupboard or in open.

Cl_2O_7 -under special conditions

Halogens Reactivity & Oxidizing Power

✓ The above rxns shows that **chlorine** has a **greater** tenden of accepting **e's** unlike bromine and Iodine.
i.e.

$Cl_2 > Br_2 > I_2$; **Reactivity reduces down the group** in halogens.

In conclusion, the **greater** the tendency to accept **e's**, the **higher** is the **oxidizing power**. Therefore, **Chlorine** is the **strongest oxidizing agent.**
i.e.

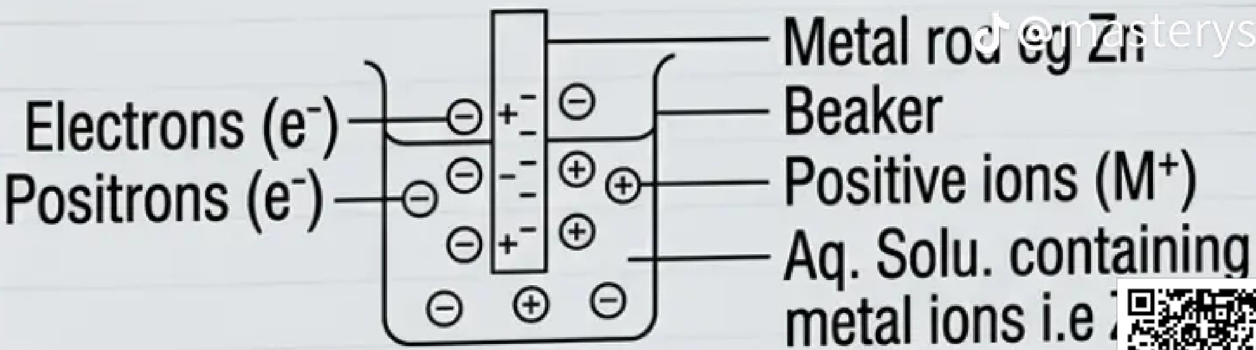
Cl (Chlorine)
Br (Bromine)
I (Iodine)

↓ **Decreasing oxidizing power**

ELECTROCHEMICAL CELL

In an electrochemical cell, chemical rxn generate electric current.

The tendency of metal's to form ions

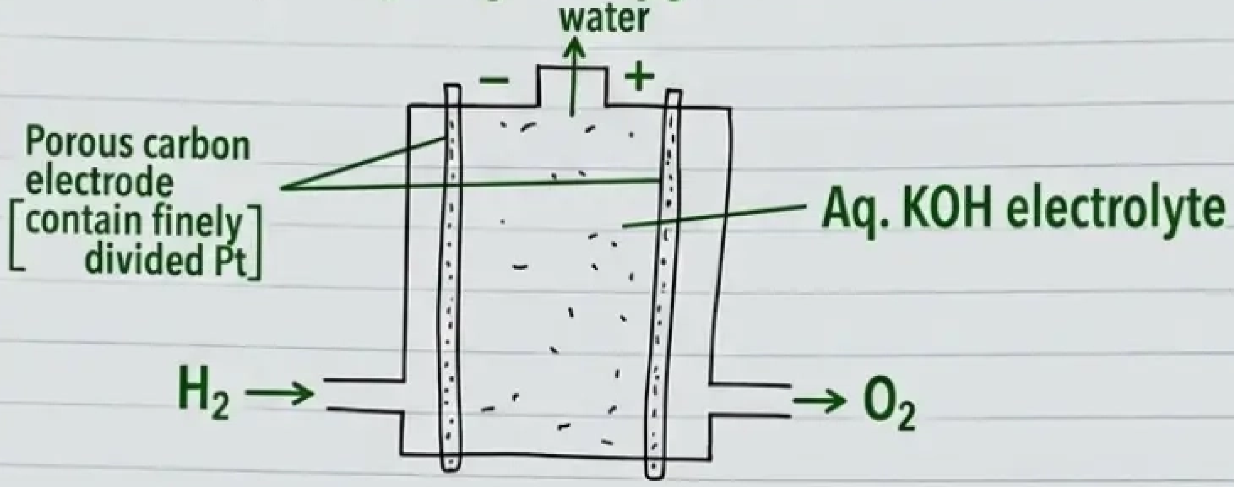


IF; a metal is dipped into an aq. soln. containin
it tends to loose e's and forms an ion.

(c) Fuel cell

These are electrochemical cells which convert the chemical energy of a fuel directly to electrical energy.

Example: Hydrogen Oxygen cell



At the –ve terminal, hydrogen reacts with OH^- ions to form water & e^-s are released.

i.e.

$$2H_2 + 4OH^-_{(aq)} \text{-------->} 4H_2O_{(l)} + 4e^-$$

At the +ve terminal, oxygen & water acquire e^-s to form OH^- ions

i.e.

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

overall rxn: $2H_{2(g)} + O_{2(g)} \text{----->} 2H_2O_{(l)}$

The cell produces electricity continuously as long as hydrogen & oxygen are fed into it. However; it does not store energy unlile every cell unlike a dry cell.

relationship of ... metal lose e's ... negativity ... much more negative change on the surface – vice-versa

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

Where; **/** : rep. change of phase.

: || : rep the **salt bridge**

NB/ The **half cell** in which oxidation takes place is always on the **left hand** cell. [NB] **electron flow** from the L.H.S to the R.H.S

Example

1. Write the cell rep developed when $Cu_{(s)}/Cu^{2+}_{(aq)}$ is connected to:

   i) $Fe_{(s)}/Fe^{2+}_{(aq)}$
   ii) $Ag_{(s)}/Ag^{+}_{(aq)}$

2. Draw the electrochemical cells in 1 and show the **dir** of flow of e^-s.

TABLE: KLB pg 108: Electrode potentials obtained when $Cu_{(s)}/Cu^{2+}_{(aq)}$ half cell is used as a **reference** half cell.

Standard electrode potential

Hydrogen is chosen as the reference electrode is usually r.t.a the **Standard hydrogen half cell.**

-* Being the ref. electrode, it is assigned an **electrode potential** value of **0.00V**

Std. conditions for measuring electrode potential

✓ Temp = **25°C.**

✓ ALL: soln hv a conc. of **1M.**

✓ pressure = **1 atm.**

C.R : $4H^+_{(aq)} + 4e^- \text{---------->} 2H_{2(g)}$

NB/ The amt. of copper oxidized at the anode is equal to the amt. of copper deposited on the cathode of they the conc. of copper (11) ions in the soln. remains the same.

Factors affecting preferential discharge

1) Position in the electrochemical series
The cations high in the reactivity series req. more energy to be reduced. AND ; ANIONS high in the reactivity series req. more energy to be oxidized.

i.e. Anion
KLB pg 127

2) Conc. of the electrolyte
A cation/anion high in concn is preferentially is the ions are close in the electrochemical series.

✓ Nature of electrode used

3) Nature of electrode used

Applc of electrolysis

*relation lose e's ... negativity ... much more negative change on the surface – vice-versa

[E°+ve - E°-ve] Cl_2O_7

B Accumulators

Lead acid accumulator contains a lead paste which is the **negative terminal** & the lead (IV) oxide plate which is the **positive terminal.**

✓ The electrodes are dipped in an aq. soln of **sulphuric (VI) acid.**

At the **negative terminal** lead atoms loose e-s to form lead (II) ions.
i.e.

$$Pb_{(s)} \text{ --------> } Pb^{2+}_{(aq)} + 2e^-$$

At the **positive terminal; PbO_2** reacts with H^+ ions in H_2SO_4 to form **lead (IV) ions.**

i.e. $$PbO_{2(s)} + 4H^+_{(aq)} + 2e^- \text{ -----> } Pb^{2+}_{(aq)} + 2H_2O_{(l)}$$

Then;

The **Pb^{2+}** ions formed reacts instantly with **SO_4^{2-}** ions to form **lead sulphate.**
i.e.

$$Pb^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \text{ ------> } PbSO_{4(s)}$$

✓ **Overall rxn:**

$$Pb_{(s)} + PbO_{2(s)} + 4H^+_{(aq)} + 2SO_4^{2-}_{(aq)} \xrightarrow{\text{discharging}} 2PbSO_{4(s)} + 2H_2O_{(l)}$$

During recharging, the reverse rxn occurs.
i.e.

$$2PbSO_{4(s)} + 2H_2O_{(l)} \text{ --------> } Pb_{(s)} + PbO_{2(s)} + 4H^+_{(aq} \ldots q)$$

The **difference** btwn the electrode potentials of the two electrodes is called **electromotive force (emf)** & is measured by a **voltmeter**.

Func. of Salt Bridge

1. Complete the circuit by providing contact btwn the two electrolytes
2. Maintains **balance of charges** in the electrolytes by providing ions to replace those ions used up in the ions formed.

The Zinc-Copper Electrochemical Cell

DIAG

Pg 107

What is observed when copper-copper ions half cell **$Cu_{(s)}/Cu^{2+}_{(aq)}$** ... is connected to **Zinc-zinc** ion half cell, **$Zn_{(s)}/Zn^{2+}_{(aq)}$**?

- ✓ The **zinc rod** wears out.
- ✓ The blue colour of **$CuSO_4$** on the Cu electrode fades away to form **brown deposits** of **copper**.
- ✓ A voltage of **1.10V** is registered by the **voltmeter**.

NB/ **oxidation** takes place at the **anode** while **reduction** occurs at the **cathode**.

Anode

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

cathode

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

✓ From these rxns **release** of e's take place from the **$Zn_{(s)}/Zn^{2+}_{(aq)}$** half cell, while **$Cu_{(s)}/Cu^{2+}_{(aq)}$** half cell **gain** e's.

Overall electrochemical cell:

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

✓ The **ionic eqn** & the **emf** of an electrochemical cel **rep** by what is r.t.a **a cell notation / cell diagram**

Manufacture of NaOH & Cl_2 from Electrolysis of Conc. sodium chloride [Brine]

The **method** used to make **chlorine** & **NaOH** is by use of **Mercury cell** i.e **flowing mercury cathode cell.**
It consists of **conc. NaCl** as the **electrolyte, graphite** or **titanium rods** as the **anode** and **mercury** as the **cathode** since it is **agood conductor** of **electricity.**

✓ When an **electric current** is allowed to pass thro' the electrolyte, **Cl^- ions** & **OH^-** move to the **anode** & **Cl^- ions** are **preferentially discharged** due to high conc.
i.e

$$2Cl^-(aq) \text{ ---------> } Cl_2(g) + 2e^-$$

✓ Both **Na^+** & **H^+** move to the **cathode**; instead of **H^+** being discharged, a **mercury cathode** allows preferentially discharge of **Na^+ ions.**
i.e

$$Na^+(aq) + e^- \text{ --------> } Na(s)$$

✓ The **Na** formed dissolves in **mercury** to form **sodium amalgam.**
i.e

$$Na(l) + Hg(l) \text{ -------> } NaHg(l)$$

sodium amalgam

✓ The **sodium amalgam** flows thro' a separate tank called **soda cell** that contains **distilled water.** Then the **sodium** in the amalgam **reacts** with **water** to form **NaOH & H_2 gas.**
✓ i.e

$$2NaHg(l) + 2H_2O(l) \text{ -----> } 2NaOH(aq) + 2Hg(l)$$

✓ The mercury is **recycled** [Thus **cut cost** & min. pollution]. H_2 gas is **pumped out** thro' **apipe** at the **top** of the **tank.** The **NaOH**

*E. Chem: - is the branch of chemist[ry] that deal with the r/ship btwn electrical current & chemical rxns

ELECTROCHEMISTRY

In form II we did electrolysis (effect of electric curent on subst). This topic deals with how chemical rxns produces electrical energy & how electricl energy produces chemical rxns.

Redox rxns

It is an abbreviation of reduction-oxidation rxn.
i.e. It involves both reduction & oxidation & involves electron gain & electron loss.

4 Instance:
In a rxn of iron filing & Copper II sulphate soltn.
i.e. $Fe_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

Observation:- A brown solid (copper) is formed.
- A green ppt of iron II sulphate is formed.

ExplN:
The rxn involves transfer of e⁻s from iron to copper while iron itself looses e⁻s to copper.
i.e.
Oxidation: Loss of e⁻s or/ (increase in oxidation number)

$$Fe_{(s)} \longrightarrow Fe^{2+}_{(aq)} + 2e^-$$

reducing agent

Reduction: Gain of e⁻s or/ (decrease in oxidation number)

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

oxidizing agent

NB/ The species that gains an e⁻ is the oxidizi[ng agent]
Example
When sodium hydroxide (is added to) aq. ammo[nium hydr]oxi[de] - is added to freshly prepared Iron II sulphate, a green ppt of Iron II hydroxide is formed.

Displacement rxns in Halogens.

Observation: when halogens are added to solns containing halide ions.

	Halide Solns		
	KCl	KBr	KI
$Cl_{2(g)}$		orange soln	Dark brown soln
$Br_{2(l)}$	No visible Δ	No visible Δ	Dark brown soln
$I_{2(s)}$	No visible Δ	No visible Δ	No visible Δ

NB/ Halogens are poisonous. The rxns should be carried in a ~~fume~~ cupboard or in the open.

EXPLN:

1. Chlorine is yellow in colour, bromine is Red-brown while Iodine is purple in colour.
2. Chlorine displaces bromine and Iodine

oxidation

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

Reduction

Oxidation step: $2Br^-_{(aq)} \longrightarrow Br_{2(l)} + 2e^-$

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

✓ Bromine has been oxidised since the Ox. no. increased from -1 to zero (0) while chlorine is reduced sin [illegible] o. reduces from 0 to -1

Therefore: Chlorine is the oxidizing agent while bromine is the

(iii) Hydrolysis of dil. H2SO4

Ions present: **Anode:** OH^-, SO_4^{2-}
Cathode: H^+

A.R : $4OH^-(aq)$ ----------> $2H_2O(l) + O_2(g) + 4e^-$

C.R : $4H^+(aq) + 4e^-$ -------> $2H_2(g)$

NB / For every **mole of oxygen** produced at the **anode, two moles of hydrogen** are formed at the **cathode.**
Hence;
The **vol. of H_2** is **twice the vol. of oxygen**

(v) Electrolysis of aq. $MgSO_4$

Ions present: **Cathode:** Mg^{2+}, H^+
Anode: OH^-, SO_4^{2-}

A.R: $4OH^-(aq)$ ---------> $2H_2O(l) + O_2(g) + 4e^-$

C.R: $4H^+(aq) + 4e^-$ -------> $2H_2(g)$

(vi) Electrolysis of $CuSO_4(aq)$ diff. electrodes eg carbon, copper electrodes

Ions present: **Anode:** SO_4^{2-}, OH^-
Cathode: Cu^{2+}, H^+

relationship of Fe^{+2}/Fe^{+3}... oxidizing agent...

REDOX REACTION OF HYDROGEN PEROXIDE

This is due to the presence of Fe^{2+} ions.

HOWEVER, IF H_2O_2 is added to acidified Iron II sulphate, a red-brown ppt of Iron III hydroxide is formed.
Reason
Hydrogen peroxide oxidizes Fe^{2+} ions to iron III Fe^{3+} ions

Oxidation Half-Reaction

i.e. $Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^-$

i.e. Oxidation no. has increased from +2 to +3 i.e. Inc. in oxidation number.

Reduction Half-Reaction

On the other hand; H_2O_2 has been reduced to water.

i.e. $2H^+_{(aq)} + H_2O_{2(aq)} + 2e^- \longrightarrow 2H_2O_{(l)}$

$$2Fe^{2+}_{(aq)} + H_2O_{2(aq)} + 2H^+_{(aq)} \longrightarrow 2Fe^{3+}_{(aq)} + 2H_2O_{(l)}$$

Oxidation Number

Is the apparent charge that an element has in a compound or on an ion.
Significance:

- To track movement of e^-s in redox reactions.
- To understand naming of inorganic compounds.

Rules for Determining Oxidation Numbers

1. Oxidation no. of uncombined element = 0
   e.g. in molecules of free atoms.
   (number drops) ↓
2. The charge on an ion with one element is equal to the Ox. no. of that element. oxidation no. (same value)
   e.g. Mg^{2+} : ox no. = +2
3. Ox. no. of hydrogen in ALL compounds = +1 E[illegible] metal hydrides where it is -1.
   e.g. HCl = +1. NaH = -1

[E°+ve - E°-ve] *relation lose e's ... negativity ... much more negative change on the surface - vice-versa

eg $Zn_{(s)}$ ---------> $Zn^{2+}_{(aq)} + 2e^-$ E° = 0.76V.

Cl_2O_7 -under special conditions

i.e.

$M_{(s)}$ --------> $M^{n+}_{(aq)} + n\ e^-$ *where* n = 1, 2, 3 ...

- The lost e-s remain on the metal surface, while the metal deps a -ve charge.
- The -ve charges attracts the ions gains & accept e-s from the metal rod & form atoms once more.

i.e. $M^{n+}_{(aq)} + n\ e^-s$ ---> $M_{(s)}$ eg $Zn^{2+}_{(aq)} + 2\ e^-s$ ---> $Zn_{(s)}$

NB/ The two rxns face place at the same time & rate i.e an **equilibrium** is established.

i.e. Metal atom ⇌ Metal ions

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

Now, A **potential difference** is always created/generated btwn the metal rod & the freely charged ions in the solu. This normally occurs whenever there is a separation of +ve & -ve charges.

✓ The conc. of e-s on the metal rod is measured by a qut. called **electrode potential.**

✓ This kind of arrn shown by the figure above is called a **half cell.**

Representation: Metal / metal ion

Where; The vertical line rep. change of phase

(Zn half cell)

DIAG : KLB - assignment

✓ use of std electrode potentials

① To compare oxidizing & reducing powers of substances

④ To determine emf of a cell

③ To predict whether or not a rxn will take place or not

EXAMPLE : KLB pg 114 - 115

✓ uses of Electrochemical cells

[Above] are used as a source of energy

i.e a) Dry cell

- They are used in a variety of electrical appliance eg radius, watches, clocks, electric bells etc.
- They are cheap & convenient to use bec, they contain

Using Carbon / platinum electrodes

- **A.R:** OH^- ions have a greater tendency to loose es unless SO_4^{2-} & hence is preferentially discharged.

  i.e

  $$4OH^-(aq) \text{ -------->} 2H_2O + O_2 + 4e^-$$

- **C.R:**

  copper has a greater tendency to gain es than Hydrogen & hence is preferentially discharged. The atoms are deposited on the cathode as red-brown coating.

  i.e

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

Observations

✓ The conc. of Cu^{2+} dec. & the blue colour fades & finally becomes colourless.

✓ H^+ ions accumulate & hence the soln. becomes acidic.

Using copper electrodes

Observation: ✓ Mass of the anode decreases while that of the cathode increases.

i.e

OH^- & SO_4^{2-} ions move to the anode but none of them is discharged; instead the copper anode is gradually oxidized & goes into soln.

i.e

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

Reason: copper atoms have a less std. reduction po an OH^- & SO_4^{2-} hen e are readily oxid

This explains why the loss of the anode is readily oxidized.

discharged since it is lower in the electrochemical series than Na^+ (i.e. low emf) i.e. H^+ has a less –ve std. reduction potential [more readily reduced]

i.e.

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

Therefore, electrolysis of dil. NaCl involves oxygen gas at the anode & hydrogen gas at the cathode. It is essentially electrolysis of water.

(ii) Electrolysis of brine (conc. NaOH)

; Anode : OH^-, Cl^-

A.R: $2Cl^- \text{---------}> Cl_2 + 2e^-$

preferent

chloride ions are preferent discharged since it has a higher conc.

C.R:

Ions present : Na^+, H^+

H^+ ions are preferentially discharged due to their low emf of 0.000V i.e have a higher tendency to gain e's

i.e

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

NB/ 1. The e's released at the anode are the same ones accepted at the cathode.

2. This process is used in the production of N

Electrolysis

This is the process in which electrical energy is used to cause non-spontaneous chemical rxns to occur.

- In electrolysis, the electrolyte undergoes chemical decomposition.
- In electrolysis of aqueous solns, there are more than two ions since water also ionises unlike molten electrolytes.
- For example: In electrolysis of aqueous $MgCl_2$, there are 2 diff cations & 2 diff anions

i.e. Mg^{2+} & H^+, Cl^-, OH^-
[Cations] [Anions]

Preferential discharge of ions dur. electrolysis

① Electrolysis of dil. NaCl
Ions present:

Na^+, Cl^- : From NaCl
H^+, OH^- : From water

IF; an electrical current is passed thro' it, OH^- ion & Cl^- ions migrate to the anode(+) since they are -vely charged. (vice versa)
OH^- ions is preferentially discharged than chlorine ions since it's lower than Cl^- in the electrochemical series. (Higher tendency to loose e^-s) ⇒ OH^- ions are more readily oxidised

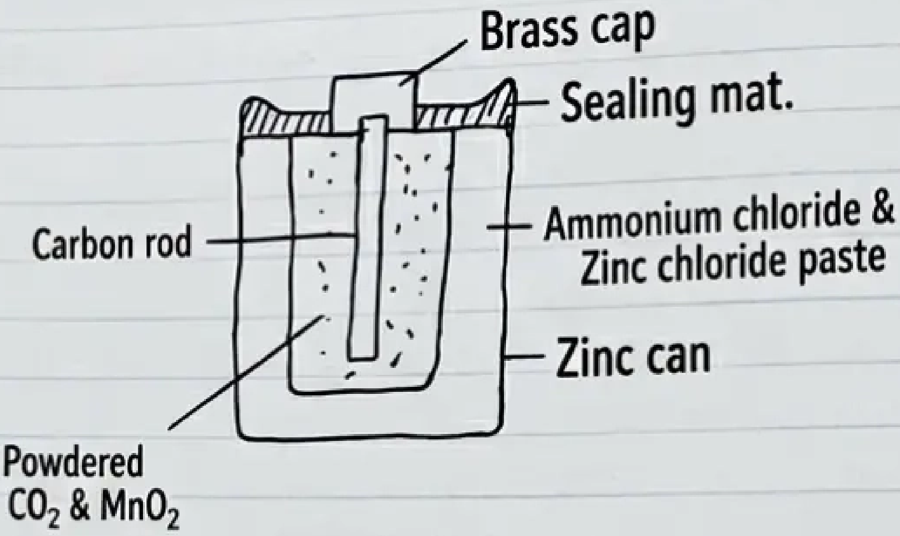
A R: $4OH^-_{(aq)}$ --------> $2H_2O_{(l)} + O_{2(g)} + 4e$

✓ At the cathode, Na^+ & H^+ are present & H^+ are preferentially

[E°+ve - E°-ve] *relation lose e's ... negativity ... much more negative change on the surface - vice-versa

* The electrolyte is paste form rather than liquid.
* Therefore, it cannot spill or leak.

Cl_2O_7



The Le-clanche cell

Rxn at the -ve terminal: $Zn_{(s)}$ -----> $Zn^{2+}_{(aq)} + 2e^-$

At the positive terminal, ammonia ions are converted to ammonia & hydrogen.
i.e.

$$2NH_4^+{}_{(aq)} + 2e^- \text{----->} 2NH_{3(g)} + H_{2(g)}$$

✓ These gases do not escape but are immediately used in other rxns.

✓ Manganese (IV) oxide oxidizes hydrogen to water, while ammonia forms a complex ion with zinc chloride paste.

✓ NB/ Any NH_4Cl paste does not conduct electricity: they are recommended.

✓ A dry cell is an example of a 1° cell; since the c... are used up & thereg the cell is discarded.
- others, can be reused after recharge. - They are cont...