



ELECTROCHEMISTRY

TERMINOLOGY

- **Oxidation**-the loss of electrons /an increase in oxidation number
- **Reduction**- the gain of electrons/ a decrease in oxidation number
- **Oxidising agent** - a substance that is reduced/gains electrons OR a substance that undergoes reduction
- **Reducing agent** - a substance that is oxidized/loses electrons OR a substance that undergoes oxidation

- **Electrode** - a solid electric conductor through which an electric current enters or leaves an electrolyte
- **Anode** - the electrode where oxidation takes place
- **Cathode** - the electrode where reduction takes place
- **Galvanic cells** - Is the electrochemical cell that **converts chemical energy to electrical energy**.
- **electrolytic cells** - Is the electrochemical cell that **converts electrical energy to chemical energy**.

- **Electrolysis** - The chemical process in which electrical energy is converted to chemical energy OR

the use of electrical energy to produce a chemical change

- **Electrolyte** – a substance of which the aqueous solution contains ions OR a substance that dissolves in water to give a solution that conducts electricity.

oxidation numbers

In order to keep track of the electrons during a chemical reaction oxidation numbers are assigned to each element in the reaction. The oxidation number is similar to the valency of the element.

Rules for assigning oxidation numbers

1. Pure elements and diatomic elements (eg. O_2) = 0.
2. Hydrogen = +1 (except when bonded to a metal, then -1).
3. Oxygen = -2 (except peroxides, eg. H_2O_2 , = -1). (When bonded to fluorine, = +2).
4. Metals = group number. group 1=+1; group 2 =+2; group 3=+3
5. The oxidation number of monoatomic ions is the same as its charge. eg. Zn^{2+} : Zn = +2
6. Group 17 elements = -1.
7. The oxidation number of transition metals is indicated by stock notation. eg. iron (III) chloride : Fe = +3
8. In a neutral compound, the sum of the oxidation numbers is 0.
9. In a polyatomic ion the sum of the oxidation numbers is equal to the charge of the ion.

REPRESENTING REDOX REACTIONS

Redox reactions can be shown in two half-reactions showing the transfer of electrons. Oxidation half-reaction: $X \rightarrow X^{2+} + 2e^{-}$ (electrons are shown as products)

Reduction half-reaction: $Y^{2+} + 2e^{-} \rightarrow Y$ (electrons are shown as reactants)

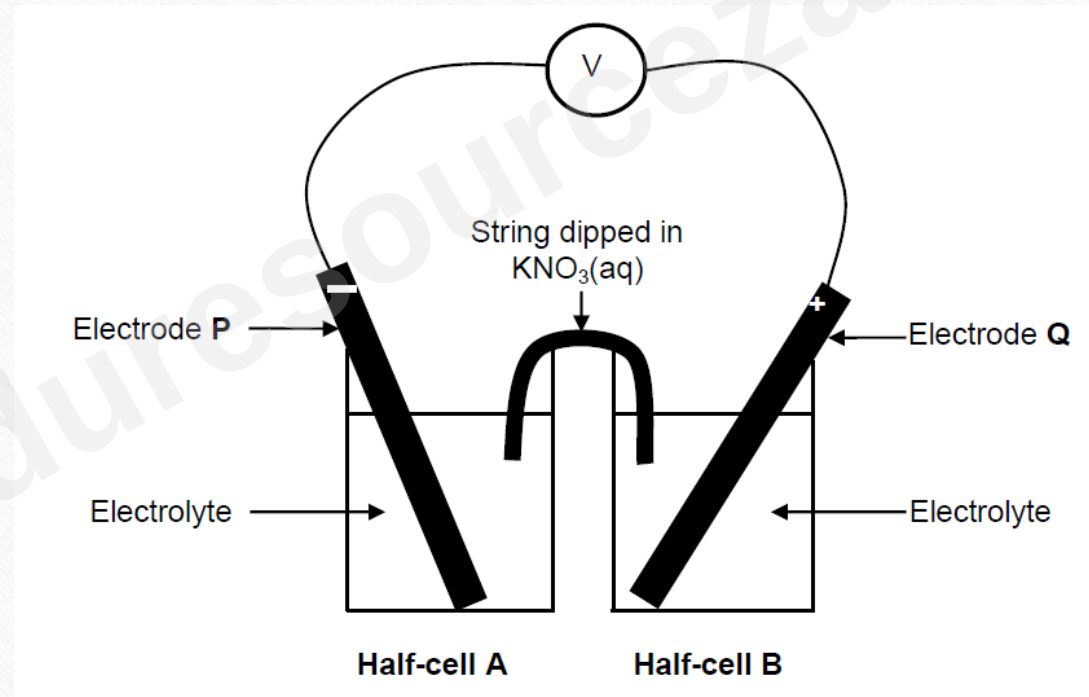
Net ionic reaction: $X + Y^{2+} \rightarrow X^{2+} + Y$ (no electrons are shown in the reaction)

NOTE: The number of e^{-} must balance

Balance redox reactions by using half reaction from the table of Standard reduction potentials (Table 4B)

THE GALVANIC CELL

The galvanic cell - Is a cell in which chemical energy is converted to electrical energy.



Galvanic Cell Cont.....

- The REDOX reaction is spontaneous and electrical energy is produced.
- The two half-cell reactions are set up in two different containers, being connected through the salt bridge.
- The anode is the negative electrode.
- The cathode is positive electrode.
- The half reaction at the anode is oxidation
- The half reaction at the cathode is reduction.
- Electrons move from anode to cathode in the external circuit.

- **Standard conditions which cells operates:**

1. Temperature = 25°C / 298K
2. Concentration = 1mol.dm⁻³,
3. Pressure = 1atm/ 1,013 x10⁵ Pa/101,3kPa

- **Functions of the salt bridge**

1. Complete the circuit
2. Maintains electrical neutrality of the electrolytes

- **Half reactions for both reduction and oxidation reactions**

- as oxidation half- reaction: $\text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(aq)} + 2e^{-}$
- as reduction half- reaction: $\text{Cu}^{2+}_{(aq)} + 2e^{-} \rightarrow \text{Cu}$
- as net ionic equation: $\text{Zn}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Zn}^{2+}_{(aq)} + \text{Cu}_{(s)}$

- **Standard cell notation**
- The cell notation starts with an anode then double vertical lines that represent a salt bridge then a cathode $\text{Zn}_{(s)} / \text{Zn}^{2+}_{(1 \text{ mol.dm}^{-3})} \parallel \text{Cu}^{2+}_{(1 \text{ mol.dm}^{-3})} / \text{Cu}_{(s)}$
- emf of the battery using the formula $E_{\text{cell}}^{\theta} = E_{\text{cathode}}^{\theta} - E_{\text{anode}}^{\theta}$

CATHODE	ANODE
• Electrode where reduction occurs • Positive electrode • Cu is the cathode • Cu (II) ions (Cu^{2+}) are attracted to cathode where they gain electrons to form Cu atoms (solid copper) • Cu atoms (Cu) get deposited on the cathode • Cu gains mass	• Electrode where oxidation occurs • Negative electrode • Zn is the anode • Zinc loses electrons to form zinc ions (Zn^{2+}) • Zn loses mass

Cathode half reaction (reduction)



Oxidising agent: copper(II) ions (Cu^{2+})

Copper(II) ions have a stronger oxidising ability than zinc ions.

Observation:

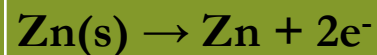
- Reddish-brown copper atoms deposit on the electrode.
- Electrode will experience an increase in mass.
- Blue electrolyte will gradually become colourless.

Net ionic equation: $\text{Zn}(\text{s}) + \text{Cu}^{2+} \rightarrow \text{Cu}(\text{s}) + \text{Zn}^{2+}$

Cell notation: $\text{Zn}(\text{s}) \mid \text{}^{2+}(\text{aq})(1 \text{ mol.dm}^{-3}) \parallel \text{Cu}^{2+}(\text{aq})(1 \text{ mol.dm}^{-3}) \mid \text{Cu}(\text{s})$

Movement of electrons: Electrons flow from anode to cathode in the external circuit.

Anode half reaction (oxidation)



Reducing agent: zinc atoms (Zn)

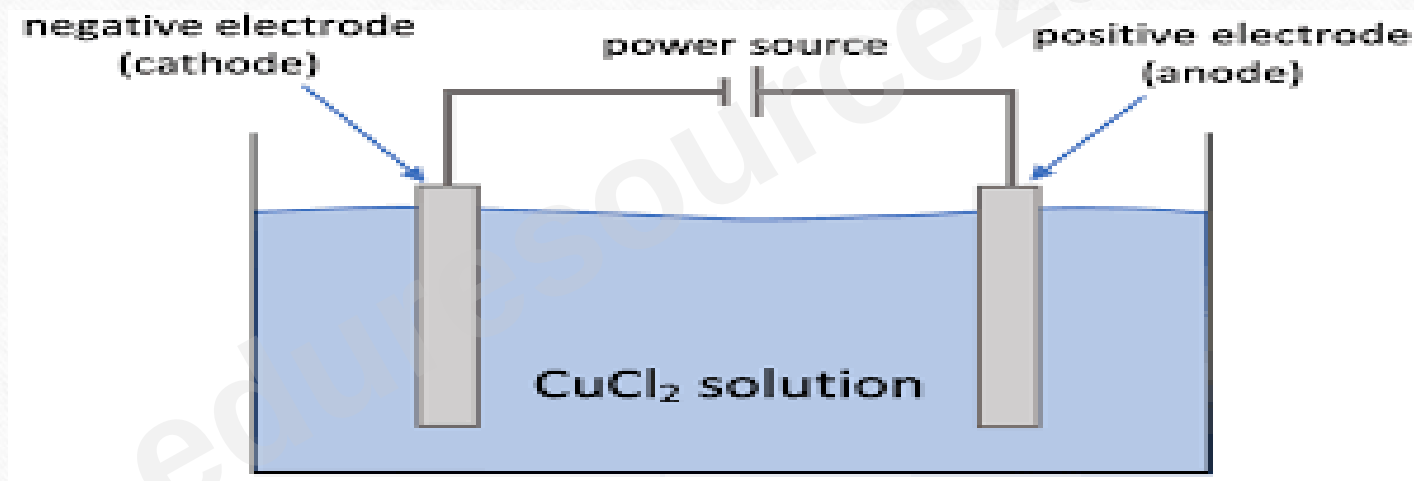
Zinc atoms have a stronger reducing ability than copper atoms.

Observation:

- Zinc electrode gets eaten away.
- Electrode will experience a decrease in mass.

THE ELECTROLYTIC CELL

The electrolytic cell is the electrochemical cell that **converts electrical energy to chemical energy**.

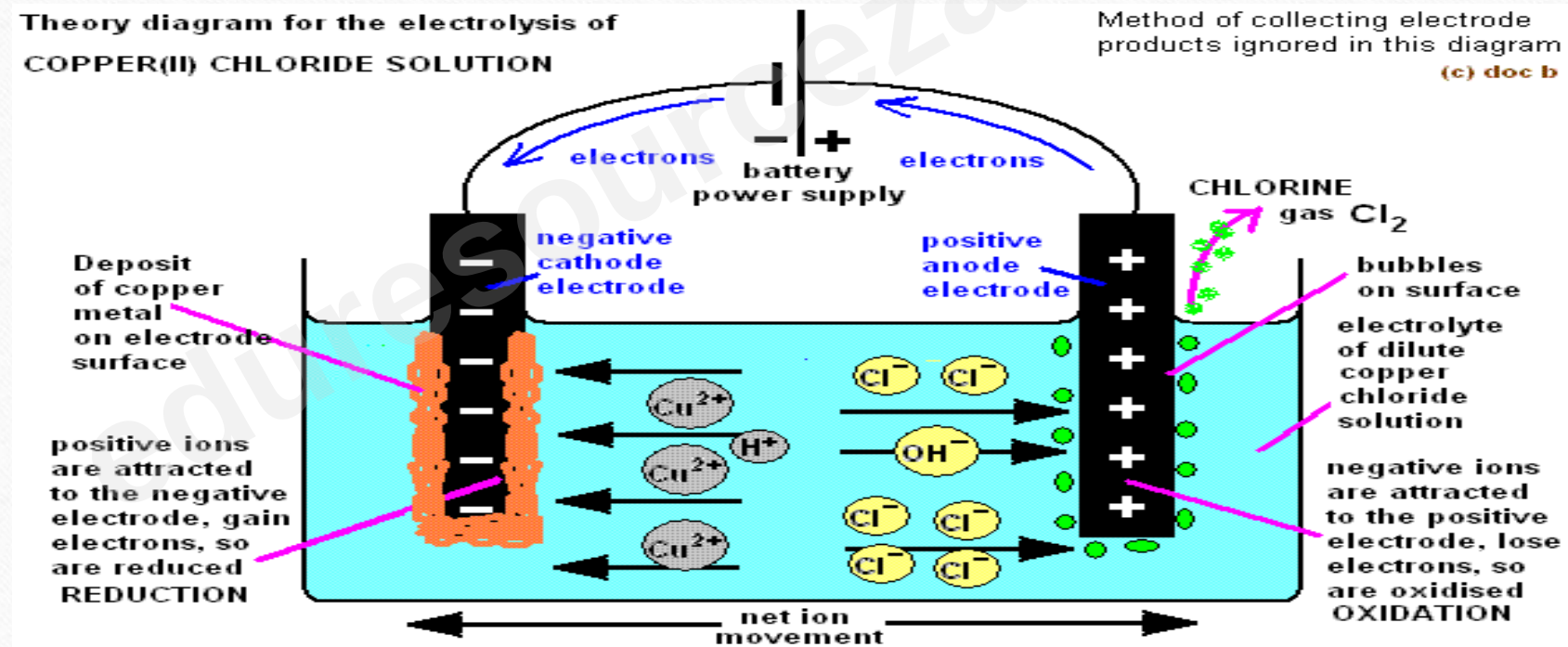


Electrolytic Cell Conti.....

- The REDOX reaction is not spontaneous and electrical energy has to be supplied to initiate the reaction.
- Both electrodes are placed in the same container with the solution of molten electrolyte.
- The anode is the positive electrode.
- The cathode is the negative electrode.
- The half reaction at the anode is oxidation (an-ox)
- The half reaction at the cathode is reduction (cat-red)
- The external battery supplies the electrons. They enter through the cathode and come out through the anode.

ELECTROLYSIS

Using a copper chloride solution, a power pack (or batteries) and carbon electrodes allow a current to pass through the solution.



Quantitative Aspects of Electrolysis

Electrolysis is a process where an Electric current drives chemical reactions across the electrodes

- Electrical current is defined as the rate of flow for any charged species. Most relevant to this discussion is the flow of electrons. Current is measured in a composite unit called an ampere, defined as one coulomb per second ($A = 1 \text{ C/s}$). The charge transferred, Q , by passage of a constant current, I , over a specified time interval, t , is then given by the simple mathematical product
- $Q = It$
- When electrons are transferred during a redox process, the stoichiometry of the reaction may be used to derive the total amount of (electronic) charge involved. For example, the generic reduction process
- $M^{n+}(aq) + ne^{-} \longrightarrow M(s)$
- involves the transfer of n mole of electrons. The charge transferred is, therefore,

$$Q = n_{e^{-}} \times q_e \quad n = N/N_A$$

Example 1

In one process used for electroplating silver, a current of 10.23 A was passed through an electrolytic cell for exactly 1 hour.

1.1 How many moles of electrons passed through the cell?

1.2 What mass of silver was deposited at the cathode from the silver nitrate solution?

Solution:

1.1 $n = 0.3817 \text{ mol e}^-$

1.2 $\text{Ag} = 41.19 \text{ g Ag}$

Example 2

Aluminum metal can be made from aluminum(III) ions by electrolysis.

2.1 What is the half-reaction at the cathode?

2.2 What mass of aluminum metal would be recovered if a current of 25.0 A passed through the solution for 15.0 minutes? (**573,61g**)

THANK YOU