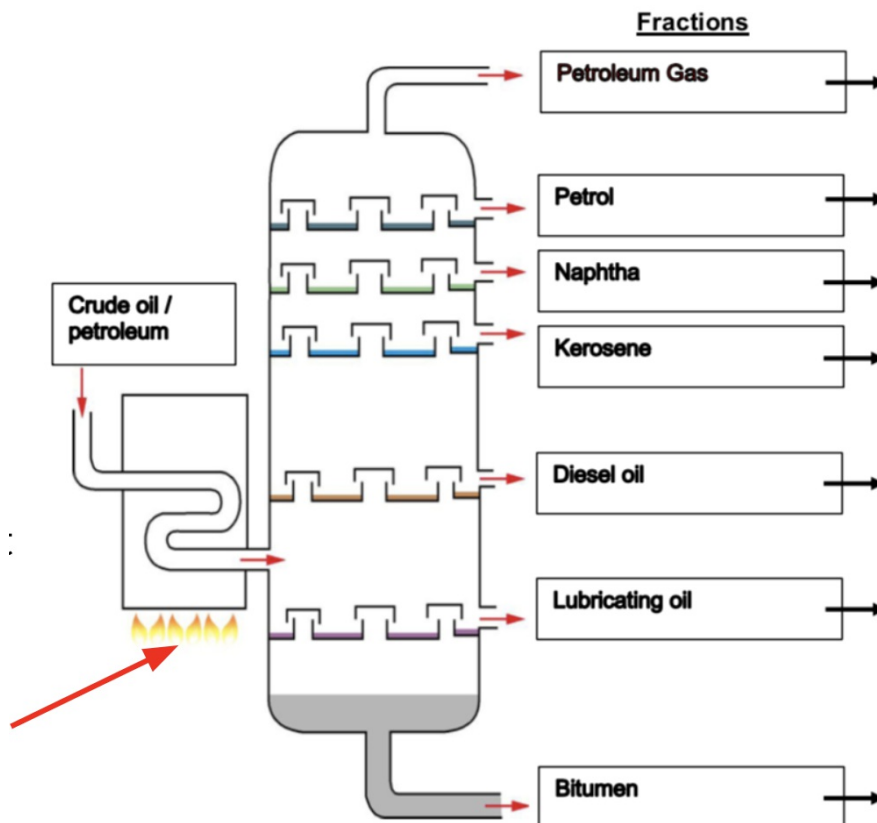


Fuels and Crude Oil

Editor's note: wtf is this topic

Fractionating Column



Mixture of crude oil is heated so that **it enters the fractionating column as a *gaseous mixture*** → ALL the fractions exist as vapour at the same time and then **separate out by selectively condensing**

Note: During fractional distillation the liquids **DO NOT 'boil' or 'evaporate' one at a time**

State the relationship between the position of the fractions collected and their respective boiling points.

Carbon-Intensive Sectors

- Electricity generation
- Transport industry

- Heavy industry
 - Types of businesses that carry a high capital cost, high barriers to entry and low transportability, e.g. petrochemical industry, oil refining, aerospace and aviation.

Introduction to Organic Chemistry

Organic Compounds

Elements that most organic compounds contain:

- All organic compounds contain **carbon**.
- Most also contain hydrogen
- Others may contain elements such as oxygen, nitrogen or a halogen.

Homologous Series

- A homologous series is a family of compounds with the **same general formula and similar chemical properties**.
- Each homologous series has a general formula.
 - Subsequent members of organic compounds in the same homologous series differ by a chemical formula of $\text{C}_n\text{H}_{2n+2}$

Functional Group

- A functional group is a group of atoms that are bonded together in a way that is unique to that particular homologous series. The functional group is responsible for a compound's characteristic **chemical properties**.
- As they possess the same functional group, members of the same homologous series will **react in a similar way to one another**.

Alkanes and Alkenes

Properties

	Alkane	Alkene
Saturated or Unsaturated?	Saturated hydrocarbon. An alkane contains only single covalent bonds between carbon atoms, C-C .	Unsaturated hydrocarbon. An alkene contains a double covalent bond between carbon atoms, C=C
General Formula	$\text{C}_n\text{H}_{2n+2}$	C_nH_{2n}

Using Prefixes to Name Organic Compounds

Prefix	Number of Carbon	Name of Alkane	Name of Alkene
meth-	1	methane	-
eth-	2	ethane	ethene

Prefix	Number of Carbon	Name of Alkane	Name of Alkene
prop-	3	propane	propene
but-	4	butane	butene
pent-	5	pentane	pentene

Properties

- E.g: Boiling point, Melting point, Electrical conductivity (physical properties).
- For Organic Chemistry: **Viscosity (physical)** and **Flammability (Chemical)**

Viscosity

- Measure of "thickness" and stickiness of a liquid
- Ability of fluid to flow

Flammability

- **Chemical property:** The ability of a chemical to burn or ignite, causing fire or combustion.
- Can compare flashpoints: GENERALLY - **higher flashpoint, lower flammability**.

Trends: MP and BP (physical)

Trend and Explanation

The melting points and boiling points of the alkanes generally increases as *relative molecular mass* increases.

- Larger molecules have stronger intermolecular forces of attraction.
- More energy is required to overcome the stronger intermolecular forces of attraction.
- higher MP and BP.

Trends: Flammability (chemical)

Trend and Explanation

The flashpoints of the alkanes and alkenes generally increases as *relative molecular mass* increases → *flammability* decreases.

- larger molecules have stronger intermolecular forces of attraction, thus less volatile (likelihood of existing as a gas)
- larger molecules also have higher percentage of carbon → requires more oxygen to burn → less flammable.

Trends: Viscosity (Physical)

Trend and Explanation

As relative molecular mass increases, the liquid alkanes and alkenes have *higher viscosity*.

- larger molecules have stronger intermolecular forces of attraction.
- harder for liquid containing larger molecules to flow.

Summary

The gradual change in properties can be observed in **all** homologous series as a result of the *increase in molecular mass of molecules down the homologous series*.

Isomerism

Definition

Isomerism is the phenomenon where **compounds (isomers) have the same molecular formula but different structural formulae**.

- Do not have to be of the same homologous series.
- Different homologous series can have the **same molecular formula**.

Alkanes and Alkenes

Chemical Reactions

- Substitution
 - One atom or a group of atoms in a molecule is replaced by another atom or group of atoms from another substance.
 - $\text{C}_2\text{H}_5\text{Br} + \text{KOH (aq)} \rightarrow \text{C}_2\text{H}_5\text{OH} + \text{KBr}$
- Addition
 - Two or more molecules react to form a single product where atoms of one reactant are added to adjacent atoms across a carbon-carbon multiple bond.
 - $\text{C}_4\text{H}_8 + \text{H}_2 \rightarrow \text{C}_4\text{H}_{10}$
- Elimination
 - The removal of atoms attached to adjacent carbon atoms in an organic compound to form an unsaturated product and a small molecule.
 - $\text{C}_4\text{H}_9\text{OH} \rightarrow \text{C}_4\text{H}_8 + \text{H}_2\text{O}$
- Condensation
 - Two molecules combine with the removal of water or some other small molecule.
 - $\text{CH}_3\text{COOH} + \text{CH}_3\text{OH} \xrightarrow[\text{Catalyst}]{\text{Conc. H}_2\text{SO}_4} \text{CH}_3\text{COOCH}_3 + \text{H}_2\text{O}$
- Hydrolysis
 - Water reacted with a molecule, causing the reactant molecule to break down into smaller molecules
 - $\text{CH}_3\text{COOCH}_3 + \text{H}_2\text{O} \xrightarrow{\text{dilute acid/alkali catalyst}} \text{CH}_3\text{COOH} + \text{CH}_3\text{OH}$

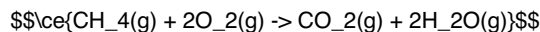
Alkanes

Unreactivity of Alkanes

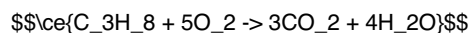
- Strong C-C and C-H covalent bonds → Hard to break
- Undergo **combustion**, **substitution** and **cracking**.

Combustion

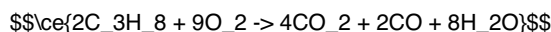
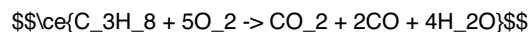
The complete combustion of an alkane produces carbon dioxide and water. The reaction is very important because it is highly exothermic.



Complete combustion of propane:



Incomplete combustion of propane (insufficient oxygen):



Substance X can be used to detect the presence of water. Identify substance X and describe any changes you would observe if water was present in the test-tube containing X.

Anhydrous copper(II) sulfate can be used to test for water. The white anhydrous copper(II) sulfate will change into blue hydrated copper(II) sulfate upon the addition of water.

Substance Y can be used to detect the presence of carbon dioxide. Identify substance Y and describe any change you would observe if carbon dioxide was present in the test-tube containing Y.

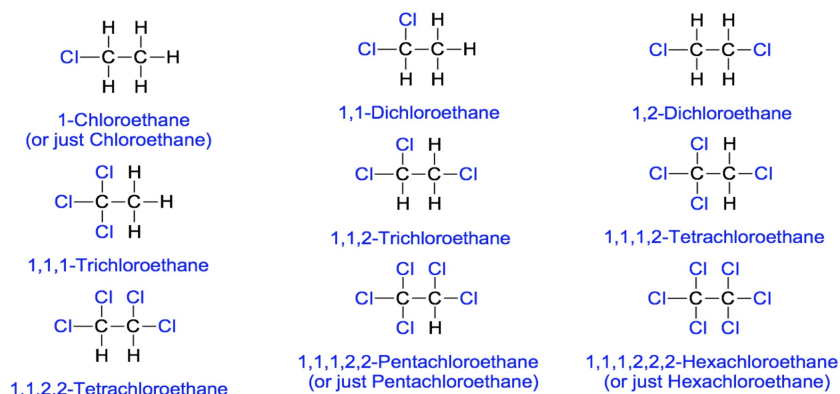
Limewater/aqueous calcium hydroxide can be used to test for carbon dioxide. A white precipitate will be formed when carbon dioxide is bubbled through limewater.

Substitution

- Alkanes react with chlorine in the presence of ultra-violet light.
- In this reaction, UV light is used to break the covalent bond in the chlorine molecule to produce chlorine atoms.
- A hydrogen atom in methane is replaced by a chlorine atom.
- The products belong to a homologous series - *halogenoalkanes*.
- The reaction produces a complex mixture of products as multiple substitutions are also possible.

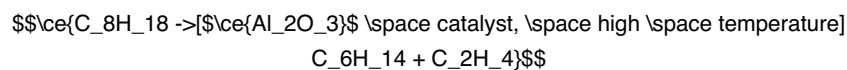


All possible reaction products



Cracking

Under **high temperatures** and in the presence of Al_2O_3 catalyst, long chain alkanes can be broken down to smaller alkenes + alkanes/hydrogen gas



For:

- production of **smaller molecules with higher demand** (e.g. ethene, propene, petrol)
- Produce **hydrogen gas for fuel or for the Haber process** .

7.1 Redox Reactions

Nikola

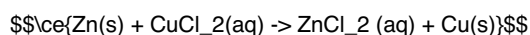
Definitions of Redox

Type	Definition of Oxidation	Definition of Reduction
Oxygen	Gain of oxygen	Loss of oxygen
Hydrogen	Loss of hydrogen	Gain of hydrogen
Electrons	Loss of electrons	Gain of electrons
Oxidation States	Increase in Oxidation State	Decrease in Oxidation State

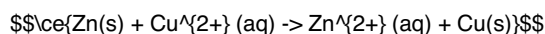
Transfer of Electrons - Definition

In the following redox reaction, **zinc metal is oxidised while copper(II) ions are reduced.**

Consider the following reaction:



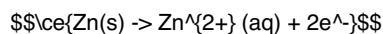
Ionic equation:



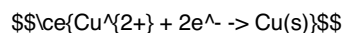
An ionic half equation shows which reactant particles gain or lose electrons to form products. To construct ionic half equations, follow these steps:

1. **Isolate reactant and product of the same element** into a single ionic half equation.
2. **Add coefficients** to balance the number of elements on the reactant and product sides of the ionic half equation
3. **Add electrons** to the ionic half equations to balance the charges on the reactant and product sides. Note that electrons are negatively charged.
4. **Write state symbols** for all reactant and product particles.

Oxidation ionic half equation involving zinc:



Reduction ionic half equation involving copper:



Reactant Particle	Is reactant particle oxidised/reduced?	Are electrons on reactant or product side?	Did reactant particle undergo gain or loss of electrons?
Zn	Oxidised	product	loss

Reactant Particle	Is reactant particle oxidised/reduced?	Are electrons on reactant or product side?	Did reactant particle undergo gain or loss of electrons?
Cu^{2+}	Reduced	reactant	gain

Using ideas about electron transfer, explain why the reaction between zinc metal and aqueous copper(II) chloride is a redox reaction.

- Zinc atom loses electrons to form Zn^{2+} ions and copper(II) ions gained electrons to form copper atoms. Hence, zinc atom is oxidised and copper(II) ions are reduced simultaneously.

Based on the definitions of oxidation and reduction written in (iii), deduce if oxidation and reduction can take place independently of one another. Explain your answer.

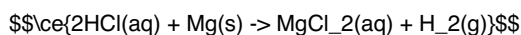
- Oxidation and reduction cannot take place independently of one another as electrons that are lost by a particle is gained by another particle.

From your answer in (v), briefly define a redox reaction.

- A redox reaction is a chemical reaction in which reduction and oxidation occur at the same time.

Examples

By constructing ionic and ionic half equations for the following reactions, use ideas about electron transfer to explain why they are redox reactions.



Ionic equation: $2\text{H}^+(\text{aq}) + \text{Mg}(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2(\text{g})$

Ionic half equations:

Oxidation: $\text{Mg}(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^-$

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

Using ideas about electron transfer, explain why the reaction is a redox reaction.

Magnesium atom loses electrons and is oxidised to form Mg^{2+} ions, while hydrogen ions gains electrons and is reduced to H_2 molecules simultaneously.

Oxidation State Rules

- Elements that are uncombined with other elements are assigned an oxidation number of 0.
- In simple ions, the oxidation number is simply the charge on the ion.

3. In complex ions, the sum of the oxidation numbers is equivalent to the net charge for polyatomic ions.
4. The oxidation number of hydrogen in all of its compounds is +1, except in metal hydrides where its oxidation number is -1.
5. The oxidation number of oxygen in all its compounds is -2, except in peroxides where it is -1.
6. The oxidation number of Group 1 elements (e.g. sodium) in their compounds is +1, for Group 2 elements (e.g. magnesium) in their compounds is +2, and for aluminium in its compounds is +3.
7. There are many oxidation numbers for Group 17 elements in their compounds but the usual one is -1.
8. The sum of the oxidation numbers of all the elements in a compound is zero.

Explaining Redox Using O.S

The oxidation state of chlorine decreased from 0 in Cl_2 to -1 in KCl , hence chlorine was reduced. The oxidation state of bromine increased from -1 in KBr to 0 in Br_2 , hence bromine was oxidised simultaneously.

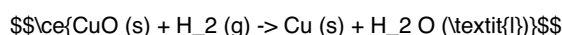
Oxidising and Reducing Agents

In a redox reaction, both oxidation and reduction take place.

- An **oxidising agent** (or oxidant) oxidises another substance - and is itself reduced.
- An **reducing agent** (or reductant) reduces another substance - and is itself oxidised.

Example

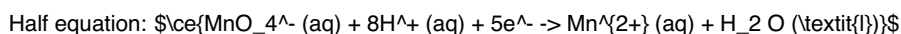
Identify the oxidising agent and reducing agent in the following reaction:



The oxidation state of **hydrogen** increased from 0 in H_2 to +1 in H_2O . Hence, hydrogen is oxidised. The oxidation state of **copper** decreased from +2 in CuO to 0 in Cu . Therefore, hydrogen is the reducing agent and CuO is the oxidising agent.

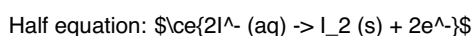
Testing of Oxidising Agents and Reducing Agents

Acidified potassium manganate(VII) solution



Acidified KMnO_4 solution is a strong oxidising agent and it turns from purple to colourless when coming into contact with a reducing agent.

Potassium iodide solution



KI solution is a reducing agent and it turns from colourless to brown when coming into contact with an oxidising agent.

7.2 Electrochemistry

Nikola

Introduction to Electrochemistry

Electrolyte

- Ionic compound in molten or aqueous state
- Conducts electricity due to free mobile ions
- Examples: mobile dilute H_2SO_4 , molten NaCl and CuSO_4 solution

Electrodes

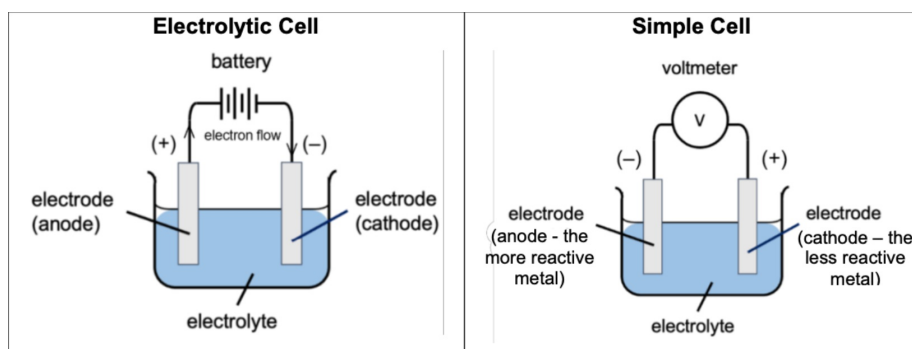
- A rod or plate which conducts electricity through the electrolyte
- **ANODE - OXIDATION** takes place at this electrode
- **CATHODE - REDUCTION** takes place at this electrode
- Types of Electrodes:
 - **Inert Electrode** - An electrode that does not take part in chemical reactions during electrolysis, e.g. graphite/carbon and platinum.
 - It does not react with the product of electrolysis or the electrolytes, and not oxidised in the process.
 - **Reactive Electrode** - An electrode that takes part in chemical reactions during electrolysis, e.g. metal plates such as copper.

Electrolytes

Electrolytes conduct electricity differently from electrodes such as metals and graphite.

	Electrolytes (ionic compound)	Metals	Graphite
Method of Conduction	Mobile ions	Flow of mobile electrons from one end to another	Flow of mobile valence electrons along the layers
Effect of Conduction	Decomposed to form new substances (chemical change)	Remains chemically unchanged.	Remains chemically unchanged

Electrolytic Cells vs Simple Cells



	Electrolytic Cell	Simple Cell
Source of Energy	Electrical energy is supplied by an external source (e.g. battery)	Electrical energy is produced through chemical reactions
Energy Conversion	Electrical -> Chemical	Chemical -> Electrical
Electron Movement	Electrons move from negative terminal of battery to cathode, through electrolyte into anode.	Electrons move from anode to external circuit (voltmeter/bulb e.t.c) into cathode.

Similarities

- Both have positive and negative electrodes
- Both have electrolyte
- Both form a closed circuit with the electrons flowing.
- Electrons flow from anode to cathode.

Differences

Electrolytic Cell

- Electrodes connected by battery (source of energy)
- Anode is positive, cathode is negative

Simple Cell

- Electrodes connected by voltmeter
- Electrodes are different metals (difference in reactivity as the source of chemical energy)
- Anode is negative, cathode is positive.

Why are there these similarities and differences?

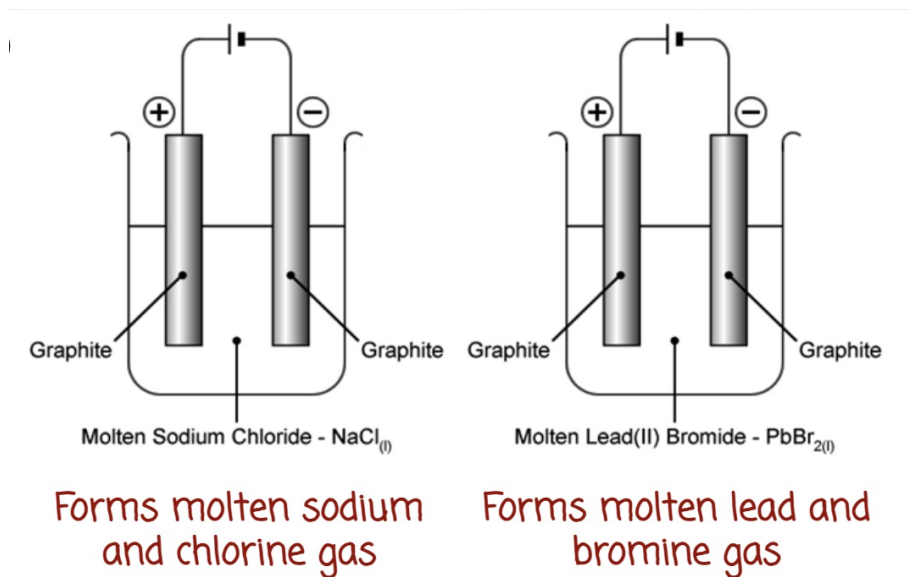
- Opposite polarity for anode and cathode due to difference in energy conversion
- Oxidation always occurs at anode, reduction always occurs at cathode.

Factors Affecting Products of Electrolysis

The polarity of a terminal in an electrolytic cell is **determined by the polarity of the battery it is connected to**.

Definition of Electrolysis

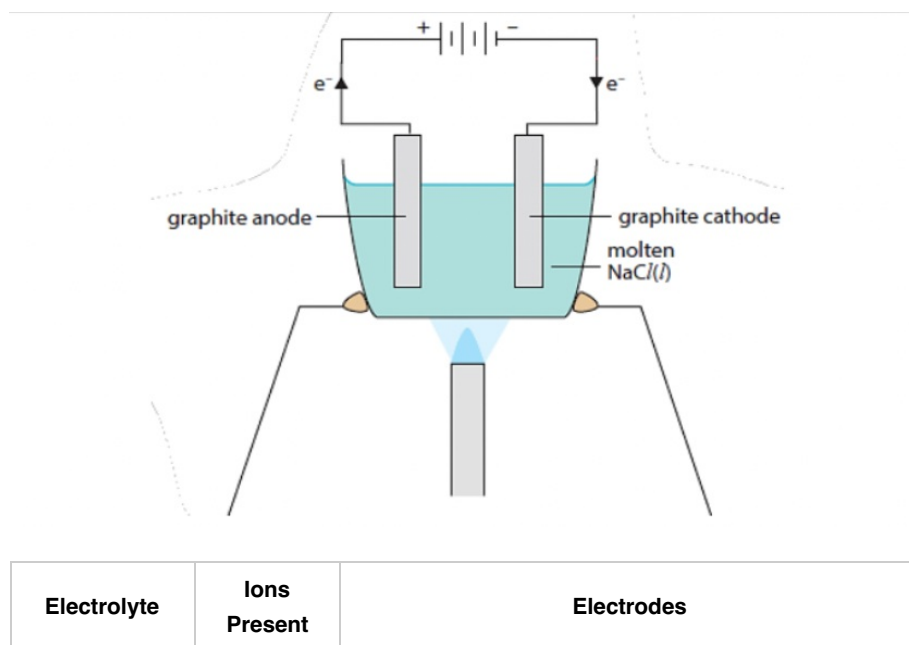
Electrolysis is the decomposition of a compound by electricity.



When the ionic compound is **molten** or when it is **dissolved in water**, the positive and negative **ions** become **mobile** and are free to move towards the electrode of opposite charge, thus **conducting electricity**.

Electrolysis of Molten Sodium Chloride

Set up



Electrolyte	Ions Present	Electrodes
molten sodium chloride, NaCl(l)	$\text{Na}^+(\text{l})$, $\text{Cl}^-(\text{l})$	Graphite electrode (inert) - may have to be replaced periodically due to graphite reacting with oxygen gas under high temperature.

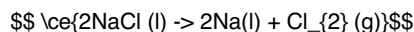
Anode (+) (Oxidation)

- Cl^- attracted to anode
- Cl^- lose electrons to form $\text{Cl}_2(\text{g})$, Cl^- is oxidised
- Cl^- discharged (*meaning ions leave the electrode*)
- Half equation: $2\text{Cl}^-(\text{l}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$

Cathode (-) (Reduction)

- Na^+ attracted to cathode
- Na^+ gain electrons form Na(l) , Na^+ is reduced
- Na^+ is discharged
- Half equation: $\text{Na}^+(\text{l}) + \text{e}^- \rightarrow \text{Na(l)}$ (reduction)

Overall Equation

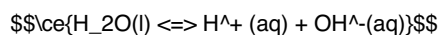


Observations

- Effervescence of greenish-yellow gas appears around at the anode.
- Greyish (or silvery) liquid is formed at the cathode.



Selective Discharge of Cations and Anions (Aqueous Electrolytes)

- Consider the difference in ions present in molten sodium chloride and aqueous sodium chloride - molten sodium chloride contains only Na^+ and Cl^- ions, while aqueous sodium chloride contains Na^+ , Cl^- , H^+ , OH^- ions.
- In an aqueous solution, water undergoes partial ionisation to form H^+ and OH^- ions:




- In an aqueous electrolyte, there will always be more than one type of cation and anion present.
- However, only one type of cation and anion will be selectively (preferentially) discharged during electrolysis.

Selective discharge of cations (reactivity series)

Increasing reactivity 	potassium (K)	K^+	 Increasing ease of discharge of cation
	sodium (Na)	Na^+	
	calcium (Ca)	Ca^{2+}	
	magnesium (Mg)	Mg^{2+}	
	zinc (Zn)	Zn^{2+}	
	iron (Fe)	Fe^{2+}	
	lead (Pb)	Pb^{2+}	
	hydrogen (H)	H^+	
	copper (Cu)	Cu^{2+}	
	silver (Ag)	Ag^+	

Selective discharge of anions (electrochemical series)

sulfate	SO_4^{2-}	 Increasing ease of discharge of anion
nitrate	NO_3^-	
chloride	Cl^-	
bromide	Br^-	
iodide	I^-	
hydroxide	OH^-	

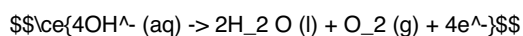
Comparing Dilute Sodium Chloride vs Dilute Copper(II) Nitrate

	Electrolysis of Dilute Sodium Chloride solution	Electrolysis of dilute Copper(II) nitrate solution
Set-Up		
Electrolyte	aqueous sodium chloride, NaCl(aq)	aqueous copper(II) nitrate, $\text{Cu(NO}_3)_2\text{(aq)}$
Ions present	$\text{Na}^+\text{(aq)}$, $\text{Cl}^-\text{(aq)}$, $\text{H}^+\text{(aq)}$, $\text{OH}^-\text{(aq)}$	$\text{Cu}^{2+}\text{(aq)}$, $\text{NO}_3^-\text{(aq)}$, $\text{H}^+\text{(aq)}$, $\text{OH}^-\text{(aq)}$
Electrodes	graphite electrode (inert)	graphite electrode (inert)

Anode (+) (Oxidation)

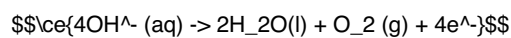
Dilute Sodium Chloride

- Cl^- and OH^- attracted to anode
- OH^- is lower than Cl^- in electrochemical series, hence OH^- is selectively discharged.
- Half equation:



Dilute Copper(II) Nitrate

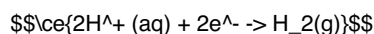
- NO_3^- and OH^- attracted to anode.
- OH^- is lower than NO_3^- in electrochemical series, hence OH^- is selectively discharged.
- Half equation:



Cathode (-) (Reduction)

Dilute Sodium Chloride

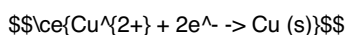
- Na^+ and H^+ attracted to cathode
- Hydrogen is less reactive than sodium, hence H^+ is selectively discharged
- Half equation:



Dilute Copper(II) Nitrate

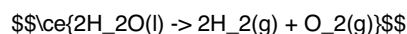
- Cu^{2+} and H^+ attracted to cathode
- Copper is less reactive than hydrogen, hence copper is selectively discharged.

- Half equation:

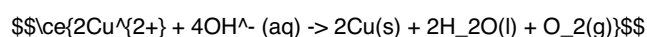


Overall Equation

Dilute Sodium Chloride



Dilute Copper(II) Nitrate



Observations

Dilute Sodium Chloride

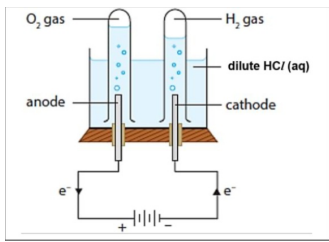
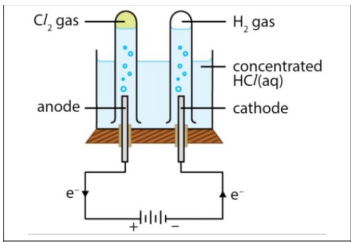
1. Solution remains neutral as both H^{+} and OH^{-} ions are discharged.
2. Water levels drops
3. Effervescence of colourless gas at both electrodes. Volume of hydrogen gas collected at the cathode is twice the volume of oxygen gas collected at the anode. For every 4 moles of electrons transferred, 2 moles of hydrogen gas and 1 mole of oxygen is produced.

Dilute Copper(II) Chloride

1. Solution becomes more acidic as OH^{-} ions are selective discharged while H^{+} ions remain in solution.
2. Reddish-brown solid formed on cathode.
3. Effervescence of colourless gas at anode.
4. Blue solution fades/turns colourless

Concentration of Halide Ions (For aqueous electrolyte)

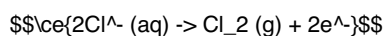
- In a concentrated solution, the ease of discharge of Cl^{-} , Br^{-} , or I^{-} anions increases.

	Electrolysis of dilute HCl(aq) acid	Electrolysis of concentrated HCl(aq)
Set Up		
Electrolyte	dilute hydrochloric acid, HCl(aq)	concentrated hydrochloric acid, HCl(aq)

	Electrolysis of dilute HCl(aq) acid	Electrolysis of concentrated HCl(aq)
Ions present	$\text{H}^+(\text{aq}), \text{OH}^-(\text{aq}), \text{Cl}^-(\text{aq})$	$\text{H}^+(\text{aq}), \text{OH}^-(\text{aq}), \text{Cl}^-(\text{aq})$
Electrodes	graphite electrode (inert)	graphite electrode (inert)

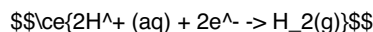
Anode (+) (Oxidation - Concentrated)

- Cl^- and OH^- are attracted to anode.
- Concentration effect applies since concentration of Cl^- is greater than OH^- , hence Cl^- is selectively discharged.
- Half equation:

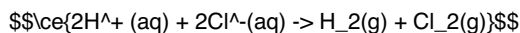


Cathode (-) (Reduction - Concentrated)

- H^+ attracted to cathode
- Since H^+ is the only cation, H^+ selectively discharged
- Half equation:



Overall Equation - Concentrated



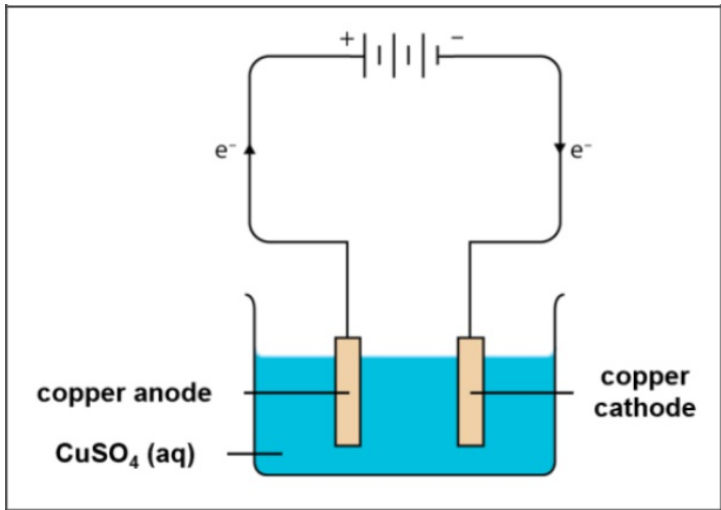
Observations - Concentrated

- Solution becomes less acidic as there is a net discharge of H^+ while OH^- ions remain in solution.
- Effervescence of colourless gas at cathode and greenish-yellow gas at anode. Volume of hydrogen gas collected at the cathode is same as the volume of chlorine gas collected at the anode. For every 2 mole of electrons transferred, 1 mole of hydrogen gas and 1 mole of chlorine is produced.

Types of Electrodes

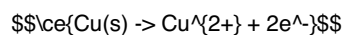
- If the electrode used is not graphite or platinum but another metal, the electrode is reactive and is oxidised as the anode during electrolysis.

	Electrolysis of dilute copper(II) sulfate solution with reactive electrode
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	Electrolysis of dilute copper(II) sulfate solution with reactive electrode
Set Up	
Electrolyte	dilute copper(II) sulfate solution
Ions present	$\text{Cu}^{2+}(\text{aq}), \text{SO}_4^{2-}(\text{aq}), \text{H}^+(\text{aq}), \text{OH}^-(\text{aq})$
Electrodes	copper electrode (inert)

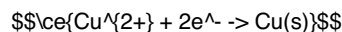
Anode (+) (Oxidation)

- SO_4^{2-} and OH^- attracted to anode
- However, copper electrode is a reactive electrode, copper electrode is oxidised instead of anions.
- Half equation:



Cathode (-) (Reduction)

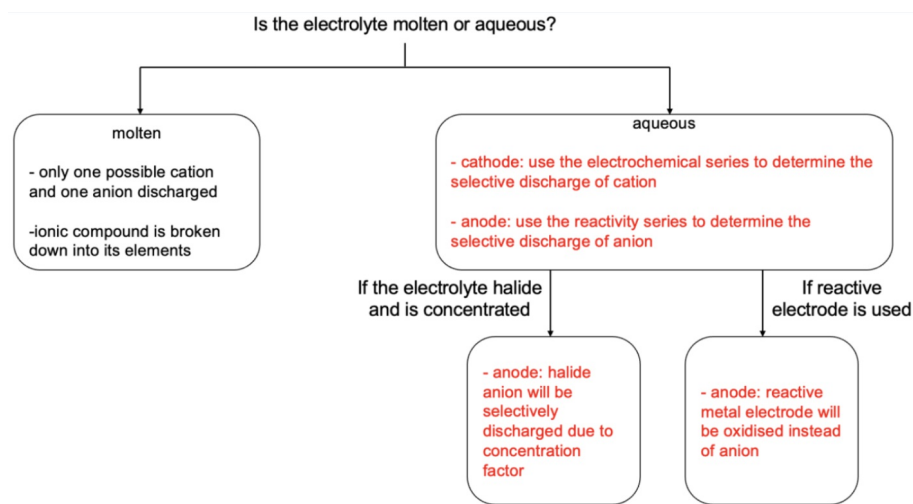
- Cu^{2+} and H^+ attracted to cathode
- Copper is less reactive than hydrogen, hence Cu^{2+} is selectively discharged.
- Half equation:



Observations

1. Solutions remains neutral as no H^+ and OH^- ions are discharged.
2. Intensity of blue solution remains as concentration of Cu^{2+} remains unchanged.
3. Size of anode decreases, size of cathode increases.

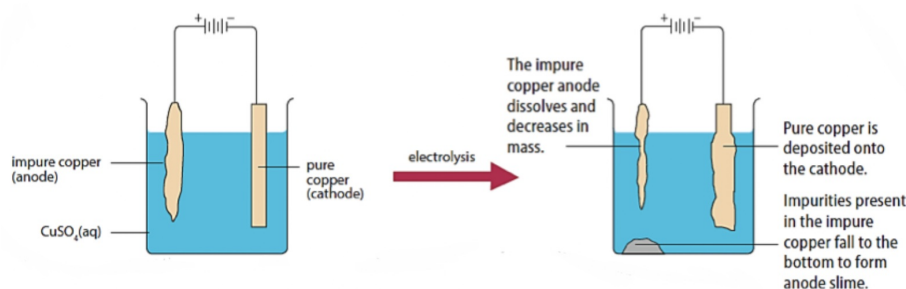
Summary



Applications of Electrolysis

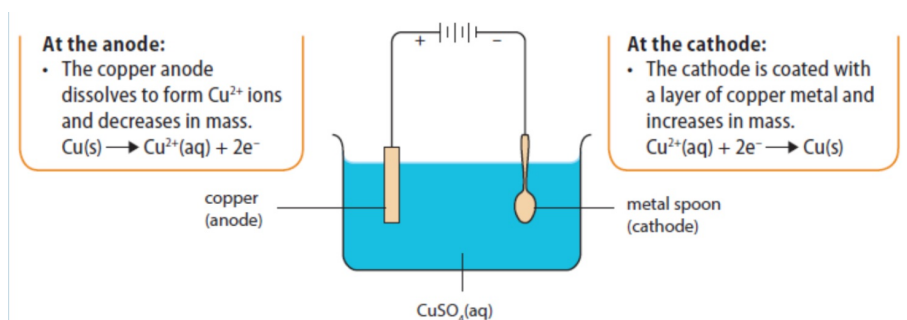
Metal Purification

- Impure copper can be purified using electrolysis
- The piece of impure copper is made the anode (+) and the piece of pure copper is made the cathode(-).
- Electrolyte is an aqueous solution of copper(II) ions.



Electroplating

- Electroplating is coating a thin layer of metal onto an object.
- This is done to protect the object from corrosion and give the object a more attractive appearance.
- The plating metal is made the anode (+), and the object to be plated is made the cathode (-)
- **Electrolyte is an aqueous solution of a salt of the plating metal.**



Hydrogen Fuel Cells

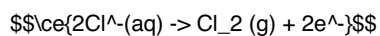
Suggest why scientists want to make hydrogen from sea water.

To reduce the usage of methane (a fossil fuel) to produce hydrogen from steam reforming.
To use non-potable sea water which is present in large quantities, to produce hydrogen instead of more precious fresh water.

Explain why electrolysis of water has high energy costs.

Electrical energy is constantly supplied to split water to form hydrogen and oxygen gas.

Write a half equation to show how chlorine is produced from chloride ions.



Reactivity Series

Metals and their Chemical Properties

Reactions between Metals

- $$2\text{K(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{KOH(aq)} + \text{H}_2\text{(g)}$$
- $$2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$$
- $$\text{Mg(s)} + \text{H}_2\text{O(g)} \rightarrow \text{MgO(s)} + \text{H}_2\text{(g)}$$
- $$\text{Ca(s)} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$

Reactivity of Metals

Metal	Reaction with Cold Water and Steam	Reaction with dilute HCl(aq) acid
Potassium (K)	react violently with cold water, explode with steam	explode with dilute hydrochloric acid
Sodium (Na)	react violently with cold water, explode with steam	reacts violently with dilute hydrochloric acid
Calcium (Ca)	reacts readily with cold water, explodes with steam	reacts violently with dilute hydrochloric acid
Magnesium (Mg)	reacts slowly with cold water, react violently with steam	react readily with dilute hydrochloric acid
Zinc (Zn)	no reaction with cold water, reacts readily with steam	reacts readily with dilute hydrochloric acid
Iron (Fe)	no reaction with cold water, reacts slowly with steam	reacts slowly with hydrochloric acid
Lead (Pb)	no reaction with cold water and steam	reacts very slowly with dilute hydrochloric acid
(Hydrogen) (H)		
Copper (Cu)	no reaction with cold water and steam	no reaction with dilute hydrochloric acid
Silver (Ag)	no reaction with cold water and steam	no reaction with dilute hydrochloric acid

Please Stop Calling Me Careless Zebra I Like Hyper Cool Smart Giraffe

Exceptions in the Reactivity Series

Lead metal reacts with dilute nitric acid. However, when lead metal is added to dilute

$\text{Pb} + 2\text{HCl} \rightarrow \text{PbCl}_2 + \text{H}_2$ or dilute H_2SO_4 , there is an initial reaction and a little effervescence before the reaction stops. From the reactivity series, lead is above Hydrogen and hence should react with acids to form salt and hydrogen gas. Explain why the observations with dilute HCl or dilute H_2SO_4 do not corroborate with the trends predicted by the reactivity series.

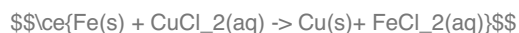
- $\text{Pb} + 2\text{HCl} \rightarrow \text{PbCl}_2 + \text{H}_2$
- $\text{Pb} + \text{H}_2\text{SO}_4 \rightarrow \text{PbSO}_4 + \text{H}_2$

Lead(II) chloride and lead(II) sulfate are both insoluble in water. Upon forming, the soluble salts will coat the lead metal and prevent the metal from further reaction with the acid.

Displacement of Metals

Displacement Reactions

Generally, a displacement reaction involves the reaction between **a more reactive element** and **the compound of a less reactive element**.



The **iron** has displaced **copper** from **copper(II) chloride** solution to form **copper** and **iron(II) chloride** solution.

Practice Question

Add a strip of zinc metal into aqueous copper(II) sulfate Equation: $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$

Observations:

- Reddish-brown solid deposits formed.
- Size of metal strip becomes smaller.
- Blue solution fades (or turns colourless)

Explanations: Zinc is more reactive than copper and would displace copper from copper(II) sulfate. Hence, copper metal is deposited and zinc metal ionises to form a colourless solution of zinc sulfate.

Thermal Stability of Metal Compounds - Decomposition

Relationship between the reactivity of metal and the thermal stability of its compound

A compound is **thermally stable** if it **does not decompose** under the influence of temperature.

Some metal carbonates can be decomposed by heat, producing carbon dioxide gas in the process.

Metal Carbonates	Observations
potassium carbonate	Unaffected by heat
sodium carbonate	
calcium carbonate	decomposes into metal oxide and carbon dioxide upon heating
magnesium carbonate	
zinc carbonate	
Iron(II) carbonate	
lead(II) carbonate	
copper(II) carbonate	
silver carbonate	decomposes into silver and carbon dioxide upon heating

The more reactive the metal is, the more difficult it is to decompose its carbonate by heat.

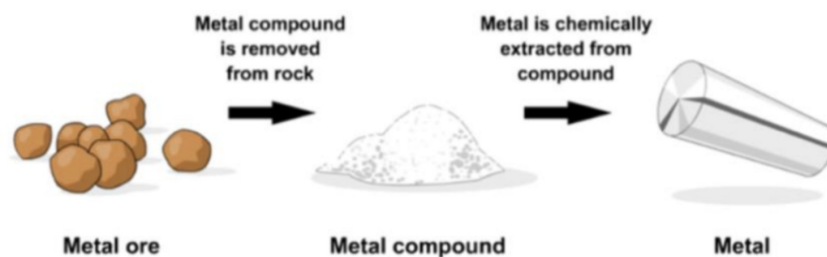
Hence, the more reactive metals form carbonates that are more thermally stable.

Discovering Metals

Extraction of Metals

The process of getting metals from the ores is called the extraction of metals.

The extraction process is as described:



There are two main methods used to chemically extract the given metal from its metal compounds.

1. **Electrolysis**: the melting of metal compounds and the use of electricity to decompose the molten compound to obtain the metal.
2. **Reduction by carbon/coke**: Heating the metal compound with carbon to obtain the pure metal. This is considered as both a redox and a displacement reaction.

The reason why there are two different methods to extract metals is:

1. Metals have different reactivity.

2. Extraction methods have costs.

Only two metals, gold and platinum, are found principally in their native states.

Many other metals naturally exist as metal compounds (such as sulfides, oxides, and hydroxides) in ores.

As most metals are reactive, once they are exposed, they would react with oxygen, water, etc. present in the surrounding to form metal compounds.

Reactivity Series	Method of extraction (Electrolysis of molten ore / reduced by heating with carbon/ found as free metal)	Reason(s)
Potassium	Electrolysis of molten ore	As carbon is less reactive than these metals, it CANNOT displace / reduce the metals from their compound
Sodium	Electrolysis of molten ore	
Calcium	Electrolysis of molten ore	
Magnesium	Electrolysis of molten ore	
(Carbon)	-	
Zinc	Reduced by heating with carbon	As carbon is more reactive than these metals, it CAN displace / reduce the metals from their oxides.
Iron	Reduced by heating with carbon	
Lead	Reduced by heating with carbon	
(Hydrogen)	-	
Copper	Reduced by heating with carbon	As carbon is more reactive than these metals, it CAN displace / reduce the metals from their oxides.
Silver	Reduced by heating with carbon	
Gold	Found as free metal ("native")	Gold does not react readily with other substances.

In a nutshell:

- Metals above carbon require extraction by electrolysis of molten ore
- Metals below carbon can be extracted by heating metal ore with carbon.
- Gold needs not be extracted chemically. It is found free (native) metal in the ground.

Consolidation

The method used to extract a given metal from its compound depends on the reactivity of the metal and the stability of the ore.

Reactivity	Metals	Extracted by Reduction with Carbon	Extracted by Reduction with Hydrogen
High	Potassium Sodium Calcium Magnesium	Cannot be extracted by reduction with carbon or hydrogen. Using electricity to decompose the molten compounds of the metal by electrolysis. E.g. $\text{Na}^+(\text{aq}) + \text{e}^- \rightarrow \text{Na}(\text{s})$ <i>(we will learn more in the chapter of Electrolysis)</i>	
Medium	Carbon		
	Zinc	Reduction with carbon by displacement reaction E.g. $\text{FeO}(\text{s}) + \text{C}(\text{s}) \rightarrow \text{Fe}(\text{s}) + \text{CO}(\text{g})$	
	Iron	See how carbon reduces FeO in this video .	Reduction with hydrogen by displacement reaction <i>*Refer to the next page</i>
	Lead		
Low	Hydrogen		
	Copper Silver	Reduction with carbon by displacement reaction. E.g. $2\text{CuO}(\text{s}) + \text{C}(\text{s}) \rightarrow 2\text{Cu}(\text{s}) + \text{CO}_2(\text{g})$ See how carbon reduces CuO in this video .	Reduction with hydrogen by displacement reaction <i>*Refer to the next page</i>
	Gold	Reduction with carbon or hydrogen, or found in native state	

Explain why the very reactive metals (metals above carbon) can only be extracted by electrolysis and not by the reduction of its oxide by carbon.

These metals are more reactive (or more easily oxidised) than carbon and cannot be extracted from their oxides by reduction of carbon.

These metals are very stable as ions in metal compounds. It is very difficult to extract the metal from its compounds.

Electricity is needed to decompose the molten ionic compound to obtain the metal (electrolysis)

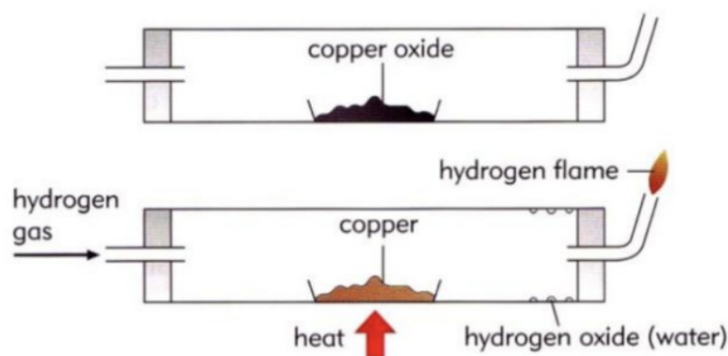
Explain why the less reactive metals (metals below carbon) can be extracted by reduction of their oxides by carbon (displacement reaction).

These metals are less reactive (or less easily oxidised) than carbon and can be displaced from their compounds by reduction with carbon to obtain the metal.

(These metals can be extracted using electrolysis too! But electrolysis is a very expensive process.)

Reduction of Metal Oxides With Hydrogen

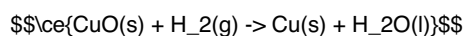
Some metals can be extracted from their metal oxide ores by the reduction with hydrogen gas.



Copper, silver and gold can be extracted from their metal oxides by reduction with hydrogen gas.

Copper < Silver < Gold (easiest)

Chemical Equation:

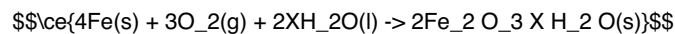


Rusting of Iron

For rusting to occur, the conditions required are:

- Presence of oxygen
- Presence of water

General Chemical Equation



Using Freshly Boiled Water and Oil

- Freshly boiled water has negligible dissolved oxygen as most would have escaped from the solution during the boiling process. This minimises the exposure of the nail to oxygen.
- The oil layer prevents oxygen from the surrounding air from dissolving in the water.

Increasing rate of rusting

- The rusting process involves the transfer of electrons
- Presence of ions in the solution increases the conductivity of the solution, allowing faster transfer of electrons.
- Rate of rust formation is accelerated.

Protection of Iron (against rusting)

Surface Protection

Coating with a layer of:

- paint
- oil/grease
- plastic or **less reactive** metal (tin, silver) by electroplating.

How does this prevent rusting?

- The protective layer prevents the metal from contacting with water and oxygen, thus reducing the rate of rusting
- If the protective layer is scraped off/removed, the iron is exposed to water and oxygen and rusting occurs.

Uses

- machine, motor cars, bridges, ships

Coated with Tin:

- Cans for can food

Coated with Chromium""

- Taps, kettles, bicycles, handlebars

Sacrificial Protection (using sacrificial metal)

Coat with a layer of Zn (galvanising)

Kitchen sinks, 'zinc' roofs

Attaching blocks of Zn or Mg metals to iron or steel

- Note: These blocks of Zn and Mg metal have to be replaced regularly once they have all been reacted.

How does this prevent rusting?

- Both Zn and Mg are more reactive than iron and would react more readily with oxygen and corrode preferentially, preventing the iron from rusting.
- A coating layer of Zn or Mg on the entire iron metal surface is not necessary. As long as Mg or Zn is present, iron will not rust.

Uses

- Underground pipes, ships, columns of steel piers

Organic Chemistry

Fuels and Crude Oil (Biofuels)

Fractions

- Petroleum Gas
- Petrol
- Naphtha
- Kerosene
- Diesel oil
- Lubricating oil
- Bitumen

The lower the fraction collected, the higher the boiling point of the fraction.

- ii. The diagram below shows the fractionating column where crude oil is separated into its various fractions. Fill in the blanks with the names of the fractions.

