

Rate of Reaction

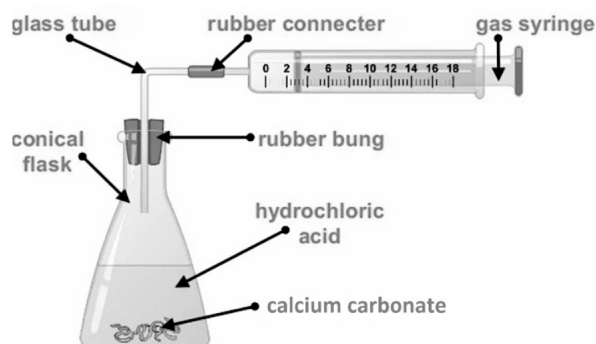
Definition: Rate of Reaction

- The **rate (or speed)** of a chemical reaction tells us **how quickly or slowly** a reaction takes place.
- It is defined as the **change in amount of product formed or reactant used per unit time**.

Methods to Determine Rate of Reaction

Method 1: Measuring Volume of Gas Produced Per Unit Time

When 0.50 g of CaCO_3 solid is added to excess hydrochloric acid, carbon dioxide is formed. The carbon dioxide gas can be collected in the gas syringe using the following apparatus. To determine the rate of the reaction, the volume of gas collected in the gas syringe can be measured over a period of time. A graph of volume of gas collected against time is plotted.



**note: you need to know how to diagram the experimental set-up.*

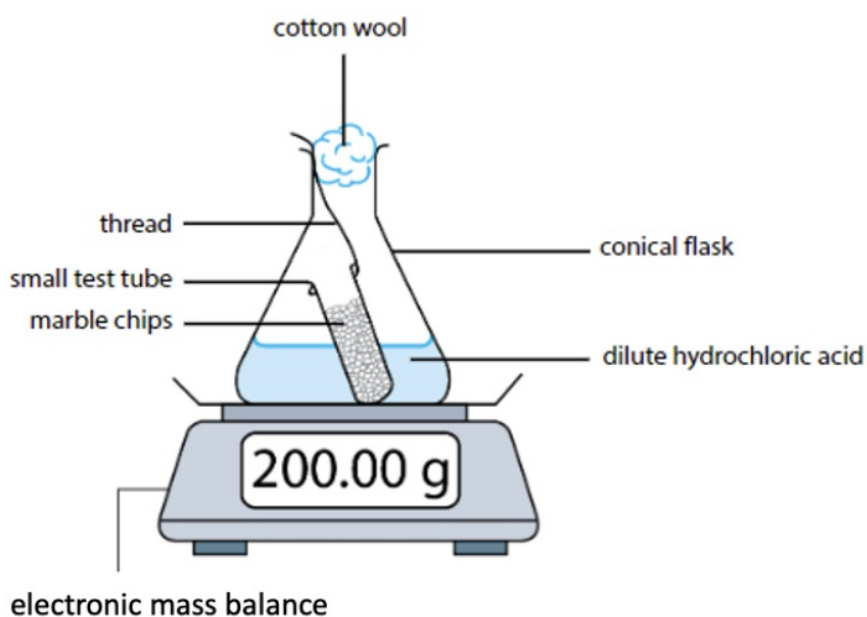
Method 2: Measuring loss in mass of reaction per unit time

To determine the rate of reaction between 0.50 g of CaCO_3 solid and excess hydrochloric acid.

Apparatus

- Conical flask
- Electronic mass balance
- Stopwatch

Experimental Setup

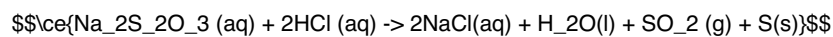


Procedure

1. Set up the apparatus as shown above.
2. Use a cotton wool plug to cover the mouth of the conical flask to prevent acid from splashing out during the reaction (resulting in a loss in reactants).
3. Record the mass of the conical flask and its contents.
4. Release the thread and ensure that the calcium carbonate are added into the acid. Start the stopwatch immediately.
5. Record the mass of the conical flask and its contents at regular time intervals (e.g. every 1 minute) until mass remains constant.
6. Use the data collected to plot a graph of mass of conical flask and contents against time.
7. The rate of reaction can be determined from the gradient of the graph of mass against time.

Method 3: Measuring a changing quantity

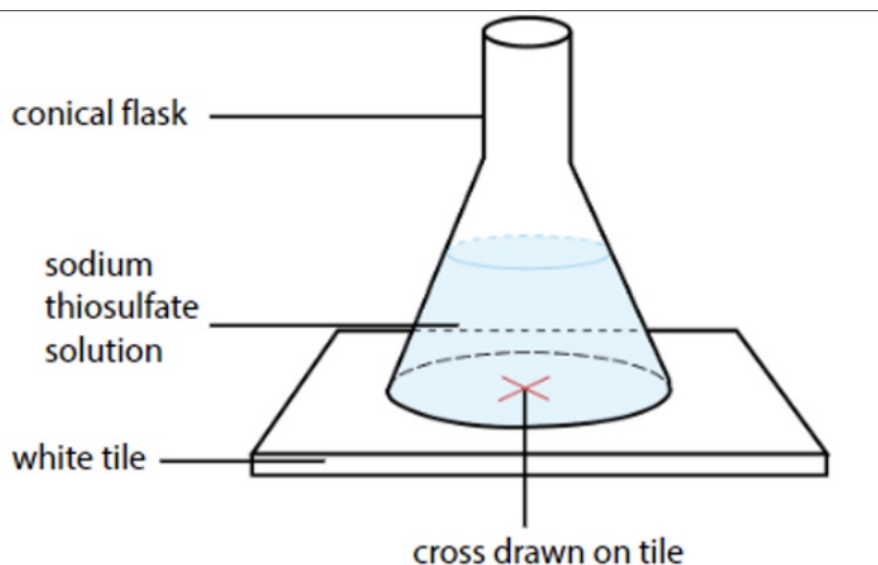
To determine the rate of the reaction between 50 cm³ of 0.200 mol/dm³ sodium thiosulfate solution and hydrochloric acid.



Apparatus

- Conical flask/beaker
- Stopwatch
- White tile with cross or colourimeter

Experimental Setup



Procedure

1. Set up the apparatus as shown above.
2. Using a measuring cylinder, measure 50 cm^3 of 0.2 mol/dm^3 of sodium thiosulfate solution into the conical flask
3. Using a measuring cylinder, add 100 cm^3 of 0.2 mol/dm^3 dilute hydrochloric acid into the conical flask and start the stopwatch immediately.
4. Look at the cross through the mouth of the flask. Stop the stopwatch when the cross is no longer visible. (yellow solid sulfur produced is opaque and blocks the cross from view as more of it is formed.)
5. Repeat the experiment with a changed variable (e.g. temperature, concentration of solution).
6. The change in rate of reaction can be determined by comparing the change in time taken for the cross to be covered.

Alternative Method

1. Use the colourimeter to measure the colour intensity of the reaction mixture in the conical flask at regular time intervals (e.g. every 1 minute).
2. Use the data collected to plot a graph of colour intensity against time.
3. The rate of reaction can hence be determined through the gradient of the graph of colour intensity against time.

Collision Theory

- For a reaction to occur and products to form, **the reactant particles must collide with each other.**
- However, only **effective collisions will result in a reaction.** An effective collision must fulfil the following conditions:
 - The reactant particles must **collide with sufficient energy, equal to or greater than the reaction's activation energy.**
 - The reactant particles must also collide with the **correct orientation**, i.e., "the

right way round".

Factors Affecting Rate of Reaction

There are five factors affecting the rate of reaction. They are:

1. **Concentration:** applicable for reactants in **aqueous** state
2. **Pressure:** applicable for reactants in **gaseous** state
3. **Particle Size:** applicable for reactants in **solid** state.
4. **Temperature**
5. **Presence of Catalyst**

Factor 1: Concentration

Lower Concentration	Higher Concentration
Fewer reactant particles per unit volume	More reactant particles per unit volume
Reactant particles are further from each other.	Reactant particles are closer to each other
Fewer collisions per unit time between particles	More collisions per unit time between particles
Frequency of effective collisions is lower	Frequency of effective collisions is higher
Slower rate of reaction	Faster rate of reaction

Factor 2: Pressure

Lower Pressure	Higher Pressure
Fewer reactant particles per unit volume	More reactant particles per unit volume
Reactant particles are further from each other	Reactant particles are closer to each other
Fewer collisions per unit time between particles	More collisions per unit time between particles
Frequency of effective collisions is lower	Frequency of effective collisions is higher
Slower rate of reaction	Faster rate of reaction

Factor 3: Particle Size

Large Particle Size (Granulated Solid)	Small Particle Size (Powdered Solid)
Bigger pieces of solid reactants	Smaller pieces of solid reactants

Large Particle Size (Granulated Solid)	Small Particle Size (Powdered Solid)
Smaller surface area of solid reactants (given same mass, less number of particles)	Bigger surface area of solid reactants (given same mass, more number of particles)
Fewer collisions per unit time between particles	More collisions per unit time between particles
Frequency of effective collisions is lower	Frequency of effective collisions is higher
Slower rate of reaction	Faster rate of reaction

Factor 4: Temperature

Lower Temperature	Higher Temperature
Particles have less kinetic energy and move slower	Particles have more kinetic energy and move faster
Less frequent collisions between particles	More frequent collisions between particles
Less particles collide with energy greater or equal to activation energy, E_a	More particles collide with energy greater or equal to activation energy, E_a
Frequency of effective collisions is lower	Frequency of effective collisions is higher
Slower rate of reaction	Faster rate of reaction

Catalysts

A catalyst increases the rate of chemical reactions without being chemically changed at the end of the reaction.

Characteristics of Catalysts

- Increased rate of reaction but not the yield
- Not used up in the reaction as it can be regenerated at the end of the reaction
- Selective in action (i.e. each catalyst only catalyses certain reactions)
- Can be poisoned by impurities
- Physical appearance may change but remain chemically the same
- Provide an alternative pathway with a lower activation energy, E_a , for the reaction

Factor 5: Presence of Catalyst

Uncatalysed reaction	Catalysed reaction
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Uncatalysed reaction	Catalysed reaction
No catalysts to provide alternative pathway of lower activation energy for the reaction to proceed.	Catalysts increase the speed of reaction by providing an alternative pathway of lower activation energy for the reaction to succeed.
Less particles collide with energy greater or equal to activation energy, E_a	More particles collide with energy greater or equal to activation energy, E_a
Frequency of effective collisions is lower	Frequency of effective collision is higher
Slower speed of reaction	Faster rate of reaction

Industrial Catalyst

- Industrial catalysts are usually transition metals and their compounds due to the ability to have variable oxidation states
- By using catalysts, less extreme physical conditions (heat and pressure) are required, saving energy.

Catalyst	Use
Iron	Manufacture of ammonia in the Haber process
Aluminium or Silicon Dioxide	Cracking of hydrocarbons
Nickel	Manufacture of margarine from vegetable oils
Platinum, Palladium, and rhodium	Catalytic converters

Biological Catalysts (Enzymes)

- Enzymes are biological catalysts, which catalyse biological processes such as digestion, energy production, detoxification etc.
- Enzymes are made of proteins
- Sensitive to temperature - most enzymes operate best at body temperature (between 35 degrees Celsius and 40 degrees Celsius). Enzymes are dormant/inactive at low temperatures and **denatured** at high temperatures. Denaturation is when the specific active site changes and enzyme cannot catalyse the specific reaction anymore.
- Sensitive to pH - operates best at optimal pH range. Different enzymes have different optimal pH range.

Specific Catalytic Action

- The active site of each enzyme is specific, and can only fit certain reactant molecules in a specific orientation.
- Applications - Manufacture of food such as cheese, yoghurt and bread, in laundry detergent.

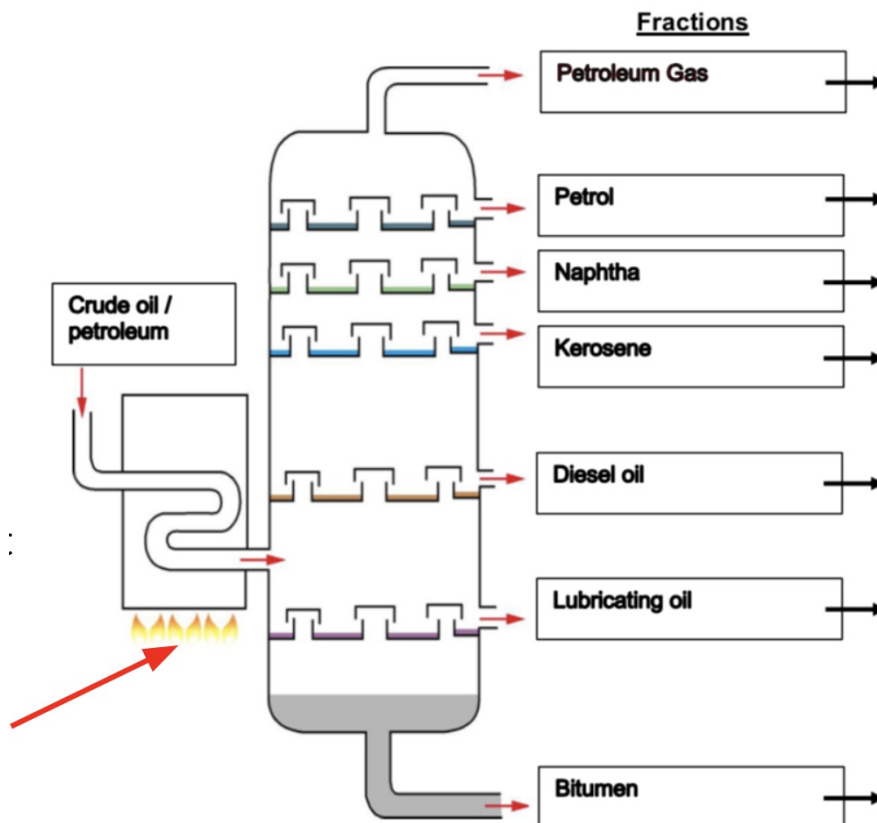
Graphical Analysis

- Graphs are often used to describe information about a chemical reaction.
- Gradient of graph indicates the speed of the reaction. The steeper the gradient, the faster the speed of reaction. The gradient can change as reaction proceeds as reactants are used up.
- The final volume of gas / mass of the reaction indicates the yield of the reaction. The yield of the products depends on the amount of limiting reagents.

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.

Difference between bioresources and biofuels

Bioresources are raw materials from human or animal activity that are renewable and biodegradable, however, biofuels are fuels that may be derived from bioresources, to be used as alternatives to conventional fossil fuels.

Government's Plan to Support Domestic Power Generation

It's plan is to blend the use of biomethane in larger quantities with natural gas in the piped networks to support domestic power generation.

Three Factors for Consideration before Extensive Implementation of Use of Biofuels in Singapore.

- Cost effectiveness in emissions reduction
- Existing infrastructure for fuels and feedstocks
- Suitability of applications for deployment by 2050.

Suggest one negative impact of the use of biofuel and bioresources on the biodiversity, land use and carbon cycle.

- Biodiversity : With the need for biofuel, there might be a tendency to increase the production of certain types of crops. This will lead to the loss of biodiversity and possible disruption in the natural eco-systems.
- Land use: With the demand for biofuel, there might be demand for land use to grow such crops to meet the demand. There will be competing uses for the land such as for agricultural needs to support food compared to the fuel industry.
- Carbon cycle: Deforestation and land conversion programmes to clear and prepare the lands required to produce biofuel, may release significant amount of CO₂. This can offset the environmental sustainability factor of biofuels of being carbon-neutral.

Possible Contribution of Singapore in Future of Biofuel

Singapore could possibly be a leader in the importing of biofuels and refining them for export.

Biofuels

- Biofuels are alternative renewable energy sources to crude oil and natural gas.
 - Biofuels are a renewable energy source that is derived from plant, algal or animal biomass.

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.
 - $\text{CH}_3\text{COOH} + \text{CH}_3\text{OH} \xrightarrow[\text{CH}_3\text{COOCH}_3 + \text{H}_2\text{O}]{\text{[Conc. H}_2\text{SO}_4 \text{ Catalyst}]}$
- Hydrolysis
 - Water reacted with a molecule, causing the reactant molecule to break down into smaller molecules
 - $\text{CH}_3\text{COOCH}_3 + \text{H}_2\text{O} \xrightarrow{\text{[dilute acid/alkali catalyst]}} \text{CH}_3\text{COOH} + \text{CH}_3\text{OH}$

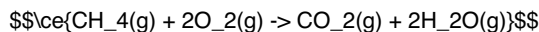
Alkanes

Unreactivity of Alkanes

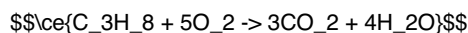
- Strong C-C and C-H covalent bonds → Hard to break
- Undergo **combustion**, **substitution** and **cracking**.

Combustion

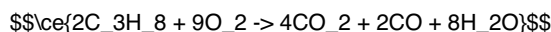
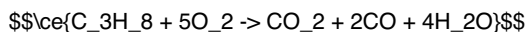
The complete combustion of an alkane produces carbon dioxide and water. The reaction is very important because it is highly exothermic.



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.

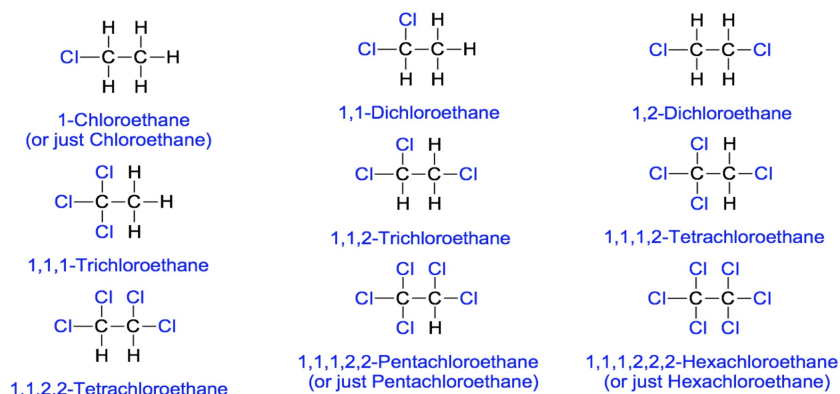
Limewater/aqueous calcium hydroxide can be used to test for carbon dioxide. A white precipitate will be formed when carbon dioxide is bubbled through limewater.

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 chloro atom.
- The products belong to a homologous series - *halogenoalkanes*.
- The reaction produces a complex mixture of products as multiple substitutions are also possible.

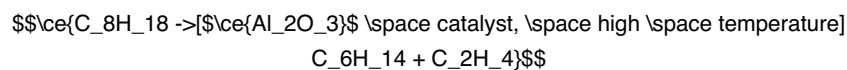


All possible reaction products



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** .

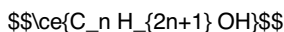
Alcohols and Carboxylic Acids

Alcohols

Functional Group:



General Formula:



Name	Molecular Formula	Structural Formula
Methanol	CH_3OH	$\text{CH}_3 - \text{O} - \text{H}$
Ethanol	$\text{C}_2\text{H}_5\text{OH}$	$\text{CH}_3 - \text{CH}_2 - \text{O} - \text{H}$
Propanol	$\text{C}_3\text{H}_7\text{OH}$	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{O} - \text{H}$
Butanol	$\text{C}_4\text{H}_9\text{OH}$	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{O} - \text{H}$

How are Alcohols Synthesised?

Method A: Fermentation of Glucose

- Fermentation: chemical process in which **microorganisms** such as yeast act on glucose to produce ethanol and carbon dioxide.
- Yeast contains **enzymes** that catalyse the breakdown of glucose to form ethanol and carbon dioxide.

Balanced Chemical Equation



Essential Conditions

- Temperature is kept at 37 degrees Celsius
- Absence of oxygen
- Enzymes in yeast

Method B: Hydration of Alkanes

Essential Conditions

- High temperature and pressure (300 degrees Celsius, 60 atm),
- H_3PO_4 catalyst

Physical Properties of Alcohols

- Soluble in water
- Most are liquids at room temperature
- Alcohols with more carbon atoms (12 and above) are solids at room temperature

The ability to form hydrogen bonds (a type of intermolecular force) allows alcohols to interact with water molecules allowing it dissolves in water. More energy is also required to overcome the hydrogen bonds resulting in higher melting and boiling points of alcohol

Solubility of Alcohols

- Alkanes
 - All alkanes are insoluble in water.
 - Alcohols: solubility decreases as number of carbons in chain increases
- The alkyl group (carbon chain) is non-polar and cannot form hydrogen bonds with water.
- The longer the carbon chain, the more insoluble the alcohol.

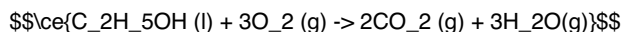
How do Alcohols React?

Alcohols are generally more reactive than alkanes.

- The C-O and O-H bonds in alcohols are more reactive than the C-C and C-H bonds in alkanes

Combustion of Alcohol

- Like all organic compounds, alcohols undergo combustion in the presence of oxygen and heat to produce **carbon dioxide and steam**.



Oxidation



What is the proof that the alcohol has been oxidised?

- Oxygen atom has been added to propanol to form propanoic acid.

What happens if wine is left exposed to air?

- Ethanol in wine is oxidised by aerobic bacteria, forming ethanoic acid.
- This process only happens to ethanol

Dehydration



As a solvent

- Used in paints, varnishes and perfumes.
- Why is alcohol a good solvent?

The structure of alcohol enables interaction with the organic substances which would not have dissolved in water.

The presence of the hydroxyl group allows to form hydrogen bonds readily with other non-organic substances.

As a fuel

- Main constituent in methylated spirit (used in spirit lamps and burners)
- In some countries, it is added to petrol and sold as a "blend".

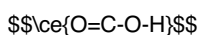
Summary: Chemical Reactions that Yield Ethanol

1. Fermentation of sugar with yeast
2. Addition of steam to alkenes (Hydration)

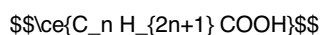
Carboxylic Acids

Structure of Carboxylic Acids

Functional Group:



General Formula:



Name	Molecular Formula	Structural Formula
Methanoic Acid	HCOOH	$\text{H}-\text{COOH}$
Ethanoic Acid	CH_3COOH	CH_3-COOH
Propanoic Acid	$\text{C}_2\text{H}_5\text{COOH}$	$\text{CH}_3-\text{CH}_2-\text{COOH}$
Butanoic Acid	$\text{C}_3\text{H}_7\text{COOH}$	$\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{COOH}$

How are Carboxylic Acids Made?

Oxidation of Alcohols

1. Heat alcohol with acidified KMnO_4 solution



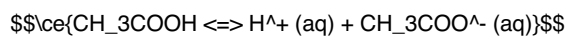
Physical Properties of Carboxylic Acids

- Soluble in water
- Most are liquids at room temperature

How Do Carboxylic Acids React?

As acids

Ionic Equation of Ethanoic Acid



General Reactions as Acids

- Being weak acids, carboxylic acids react with:
 - Metals
 - Bases
 - Carbonates

Write balanced chemical equations for reactions involving carboxylic acids

- $2\text{CH}_3\text{COOH} (\text{aq}) + \text{Na}_2\text{CO}_3 (\text{aq}) \rightarrow 2\text{CH}_3\text{COONa} (\text{aq}) + \text{H}_2\text{O} (\text{l}) + \text{CO}_2 (\text{g})$
- $2\text{C}_3\text{H}_7\text{COOH} (\text{aq}) + \text{CaO} (\text{s}) \rightarrow (\text{C}_3\text{H}_7\text{COO})_2\text{Ca} (\text{aq}) + \text{H}_2\text{O} (\text{l})$
- $\text{HCOOH} (\text{aq}) + \text{KOH} (\text{aq}) \rightarrow \text{HCOOK} (\text{aq}) + \text{H}_2\text{O} (\text{l})$

Esterification - Introduction to Esters

What are esters?

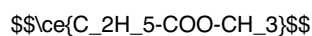
- Sweet aroma
- Used in food and perfume industry

Homologous Series - Esters

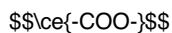
Ethyl Ethanoate



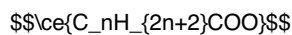
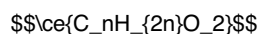
Methyl Propanoate



Functional Group



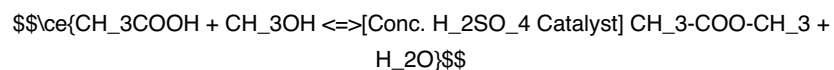
General Formula



Name

-yl -anoate

Synthesis of Esters



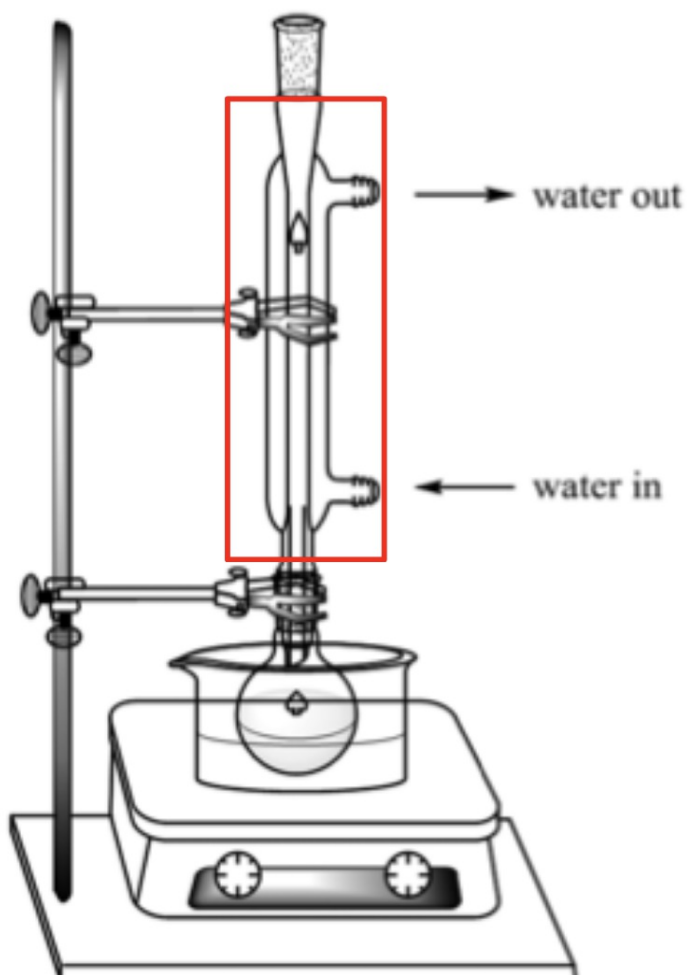
Carboxylic acid + alcohol \rightleftharpoons ester + water

Experimental Set-up

- The carboxylic acid and alcohol are heated under **reflux** in the presence of a **catalyst** (a few drops of concentrated sulfuric acid).

Which apparatus is the condenser? purpose?

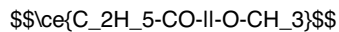
- To prevent volatile organic reagents from escaping.
- Any vapour will condense on the cool surface of the condenser and flow back into the flask.



Hydrolysis of Esters

- Esters can be **hydrolysed** (broken down by water) to form original carboxylic acid and alcohol
- Warm with dilute acid

1. Look out for the C-O bond in the functional group of ester



- methyl propanoate

2. Separate out the molecule



3. Add H to the C=O , to form an acid. Add H to the remaining part of the compound to form an alcohol.



- propanoic acid + methanol

Note

Why does the use of dilute acid favour the reverse reaction?

Hint: What product is formed in the forward reaction?

The presence of water shifts the equilibrium to the left, hence a dilute acid solution would favour the reverse reaction.

Polymers

Macromolecules

- Macromolecules are made from smaller subunits, such as:
 - Polymers
 - Smaller subunits are the same type of molecules
 - Non-Polymers

Polymers and Repeating Units

- Polymers are made from many **smaller molecules** called **monomers**
- Each small unit within the polymer that repeats itself is a **repeating unit**

Maintaining Air Quality

Composition of Air

Composition (by volume) of Air

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{
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      "Oxygen",
      "Argon",
      "Carbon dioxide"
    ],
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          0.96,
          0.44
        ],
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          "#cba6f7",
          "#89b4fa",
          "#fab387"
        ]
      }
    ]
  }
}
```

- **Nitrogen:** 78%
- **Oxygen:** 21%
- **Argon:** 0.96%
- **Carbon Dioxide:** 0.04%

Humidity

Humidity is the measurement of water vapour content in air and the water vapour content in air varies.

Suggest a possible reason why the water vapour content in air varies

Humidity is affected by **temperature** and **pressure**.

Location can also affect the humidity. In the desert, it is likely that there will be less water vapour present as compared to the coastal areas.

Separation of Air

How can air be separated to obtain the various constituent gases?

Type of Gas	Composition by Volume	Boiling Point / °C
Nitrogen	78%	-196.0
Oxygen	21%	-183.0
Argon	0.93%	-185.0
Carbon dioxide	0.04%	-78.5

Based on the table, each gas has its own boiling point and **through fractional distillation of liquefied air**, it is possible to obtain each individual gas.

- In the Haber Process, the nitrogen used in the manufacture of ammonia is obtained through distillation of liquefied air.
- However, this is not encouraged for gases aside from nitrogen.
 - That is due to the fact that attempting to obtain any gas aside from nitrogen from liquefied air gives a low yield as the composition of gases in air shows that majority of air is made up of nitrogen.

Types of Air Pollutants

Air pollutants, their sources, and effects

Carbon Monoxide, CO

Source of Pollutant

- **Incomplete combustion** of carbon-containing fuels
- Examples of sources:
 - Vehicles and industrial plants
 - Forest fire

Effects of Pollutant

- Colourless, odourless **toxic** gas that **reduces the ability of haemoglobin in blood to transport oxygen to the rest of the body**. (due to formation of carboxyhaemoglobin)
- Leads to **breathing difficulties, headaches, fatigue before death**.

Ways to Reduce Pollutant

- **Catalytic converters** to convert carbon monoxide to carbon dioxide
- **Reduce** the use of carbon-containing fuels

Relevant Chemical Equations

Chemical equations in catalytic converters:

- $\text{2NO (g)} + \text{2CO (g)} \rightarrow \text{2CO}_2 \text{ (g)} + \text{N}_2 \text{ (g)}$
- $\text{2CO (g)} + \text{O}_2 \text{ (g)} \rightarrow \text{2CO}_2 \text{ (g)}$

Oxides of Nitrogen - NO , NO_2

Source of Pollutant

- N_2 and O_2 in the air reacts at **high temperatures** to form NO .
- NO then further reacts with O_2 to form NO_2 .

Examples:

- Combustion engine of vehicles
- Forest fire
- Lightning

Effects of Pollutant

Impact on Health

- Irritates eyes and lungs, causing breathing difficulties

Effects on Environment

- Forms **acid rain** which **corrodes buildings and harms plants and aquatic life** (Give the impacts!)

Ways to Reduce Pollutant

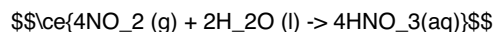
- **Catalytic Converters** to convert nitrogen monoxide to nitrogen

Relevant Chemical Equations

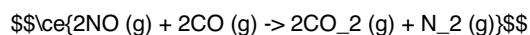
How are oxides of nitrogen produced?

1. $\text{N}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightarrow \text{2NO (g)}$
2. $\text{2NO (g)} + \text{O}_2 \text{ (g)} \rightarrow \text{2NO}_2 \text{ (g)}$

How is acid rain formed?



How does catalytic converter reduce oxides of nitrogen?



Sulfur Dioxide - SO_2

Source of Pollutant

- Combustion of fossil fuel such as coal, petroleum

Examples of Sources:

- Power stations
- Volcanic eruption

Effects of Pollutant

Impact on Health

- Irritates eyes and lungs, causing breathing difficulties

Impact on Environment

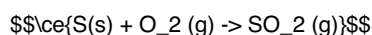
- Forms **acid rain** which corrodes buildings and harms plants and aquatic life.

Ways to Reduce Pollutant

- Expensive method:
 - Remove fossil fuels before they are burnt
- Cheaper method:
 - Flue gas desulfurisation

Relevant Chemical Equations

How is sulfur dioxide produced?



How is acid rain formed?

1. $\text{2SO}_2 \text{ (g)} + 2\text{H}_2\text{O (l)} \rightarrow 2\text{H}_2\text{SO}_4 \text{ (aq)}$
2. **or** $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$
3. $\text{2H}_2\text{SO}_3 + \text{O}_2 \rightarrow 2\text{H}_2\text{SO}_4$

How does flue gas desulfurisation work?

- $\text{CaCO}_3 \text{ (s)} \rightarrow \text{CaO(s)} + \text{CO}_2 \text{ (g)}$
- $\text{CaO(s)} + \text{SO}_2 \text{ (g)} \rightarrow \text{CaSO}_3 \text{ (s)}$
- $2\text{CaSO}_3 \text{ (s)} + \text{O}_2 \text{ (g)} \rightarrow 2\text{CaSO}_4 \text{ (s)}$

Unburnt Hydrocarbons (e.g. octane)

Source of Pollutant

- Incomplete combustion of fuels in vehicle engine

Effects of Pollutant

Impact on Health

- Cancer causing (carcinogenic)
- Reacts with sunlight and other pollutants to form ozone.

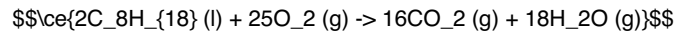
Ways to Reduce Pollutant

- **Catalytic converters** to convert unburnt hydrocarbons to carbon dioxide and water

- Use of cleaner fuels such as hydrogen

Relevant Chemical Equations

How do catalytic converters reduce unburnt hydrocarbons?



Effects of Acid Rain

Ground level ozone, O_3

Source of Pollutant

- Forms when **unburnt hydrocarbons and nitrogen oxides** react in the presence of sunlight

Effects of Pollutant

Impact on health

- Creates **photochemical smog** (a mixture of ozone, unburnt hydrocarbons, and nitrogen oxides) which **irritates eyes and lungs, causing breathing difficulties**

Impact on environment

- damages crops

Ways to Reduce Pollutant

see methods of reducing unburnt hydrocarbons and nitrogen oxides as they react to produce ground level ozone.

Ozone Layer

Ozone and its Importance

- Ozone is an important part of the stratosphere
- Acts as shield, **filtering out some of the harmful ultraviolet radiation**, *reducing the chances of getting skin cancer, genetic mutations and eye damage.*

How is Ozone Similar to Oxygen?

- Ozone is an allotrope of oxygen, formed out of 3 oxygen atoms.

Important Fact

- In small concentrations, ozone is non-toxic
- However, in high concentrations (100 ppm), **ozone becomes toxic**
- Too much ozone can damage the tissues of the respiratory tract.

Chlorofluorocarbons

- Chlorofluorocarbons are examples of ozone-depleting substances (ODS)
- Also known as CFCs, they are compounds that contain the elements carbon, fluorine, and chlorine.
- **CFCs are generally used in the manufacture of aerosol sprays, blowing agents for foams and packing materials, as solvents, and as refrigerants.**

Role of CFCs

CFCs in the atmosphere are decomposed by sunlight to produce chlorine radicals. These chlorine radicals then react with ozone molecules and destroy them by converting them into oxygen molecules.

1. $\text{CFCl}_3 \xrightarrow{\text{UV light}} \text{CFCl}_2 + \text{Cl}^\bullet$
2. $\text{Cl}^\bullet + \text{O}_3 \rightarrow \text{ClO}^\bullet + \text{O}_2$
3. $\text{ClO}^\bullet + \text{O}^\bullet \rightarrow \text{O}_2 + \text{Cl}^\bullet$

Carbon Cycle

- Describes the process in which carbon atoms continually travel from the atmosphere to the Earth and then back into the atmosphere.
- This cycle is important for maintaining a stable climate and carbon balance on Earth.
- Carbon dioxide is **0.03%** of atmosphere content by volume
- This represents a huge reservoir of carbon, which is constantly being removed from and returned to the atmosphere by a variety of processes.

Process that takes in carbon from the atmosphere	Process which gives out carbon to the atmosphere
Photosynthesis by plants	Respiration by animals
Ocean uptake	Respiration by plants
	Combustion of carbon-containing fuels
	Bacterial decomposition of organic matter

Salts

What are Salts?

A salt is an ionic compound that consists of a cation (positive ion) and an anion (negative ion).

Reactions that produce salts

- acid + reactive metal → salt + hydrogen
- acid + carbonate → salt + water + carbon dioxide
- acid + base → salt + water
- base + ammonium salt → salt + water + ammonia gas

Solubility of Salts

Compound Containing	Solubility	Exceptions
Sodium, Potassium, Ammonium ions	All are soluble	-
Nitrate ions	All are soluble	-
Chloride ions	All are soluble	lead, silver insoluble
Sulfate ions	All are soluble	barium, calcium, lead insoluble. Silver sparingly soluble
Carbonate ions	All are insoluble	sodium, potassium, ammonium, Group 1 are soluble
Hydroxide ions	All are insoluble	sodium, potassium, Group 1, barium soluble. Calcium sparingly soluble.
Halides	All are soluble	silver, lead
Oxides	All are insoluble	-

Preparation of Salts

Uses of Salts

- Ammonium Phosphate
 - $(\text{NH}_4)_3\text{PO}_4$

- A fertiliser used to provide plants with the elements nitrogen and phosphorus, essential for the plants' healthy growth.
- Monosodium Glutamate
 - $\text{NaC}_5\text{H}_8\text{NO}_4$
 - A sodium salt of glutamic acid which is used as a flavour enhancer in the food industry, and is especially used to enhance the taste of savoury foods.
- Sodium Fluoride
 - NaF
 - A salt used to provide the fluoride ions in toothpaste. The fluoride ions prevent cavities and tooth decay.

Criteria for Preparation of Salts

- Solubility of the salt product
- Solubility of the reactants
- Method to achieve minimal contamination
- Ease of obtaining pure products
- Safety of procedure

Method 1: Reaction of acid with an excess insoluble substance, metal / carbonate / base

- To prepare a soluble salt from an acid
- Cation of the salt is provided by the insoluble substance, metal / carbonate / base
- Anion of the salt is provided by the acid.

Considerations

- Very reactive metals such as Group 1 metals are not reacted with acids due to safety as these metals react explosively with acids.
- Unreactive metals such as copper, silver, and gold do not react with acids.

Steps

1. Using a measuring cylinder, transfer 50cm^3 of a suitable acid into a beaker.
2. Add **excess** suitable insoluble metal / carbonate / base to the acid. Stir the mixture continuously until no more insoluble metal / carbonate / base can dissolve.
 - Excess of the insoluble substance can be seen at the bottom of the beaker.
 - **Rationale**
 - Insoluble substance is added in excess to ensure that all the acid has reacted. If not, the mixture at the end will contain excess acid and contaminate the salt produced.
3. Filter to remove excess insoluble substance (unreacted) as residue. Collect the filtrate, which is the desired salt solution
 - To remove excess insoluble substance
4. Heat the filtrate in an evaporating dish until saturated.
 - To obtain a saturated solution for crystallisation.
5. Cool the saturated solution to allow salt crystals to form.
 - Solubility decreases as solution cools.
6. Filter to collect the crystals. Wash the crystals with a little cold distilled water. Dry the

crystals between sheets of filter paper.

- Use cold distilled water to minimise dissolving of the soluble salt crystals.

Method 2: Titration

- To prepare a soluble salt containing group 1 or ammonium cation
- Cation of the salt is provided by a soluble alkali or group 1 metal carbonate
- Anion of the salt is provided by the acid.

Considerations

- Titration is used to prepare group 1 salts as the metals of these salts are very reactive metals and group 1 carbonates and bases are soluble. Hence, method 1 was unsuitable.

Steps

1. Fill a burette with suitable dilute acid. Note the initial burette reading, $V_1 \text{ cm}^3$
2. Pipette 25.0 cm^3 of suitable aqueous alkali or aqueous carbonate into a conical flask.
3. Add a few drops of a suitable indicator to the solution in the conical flask.
4. Add the dilute acid from the burette slowly, swirling the conical flask, until the indicator changes colour permanently. Record the final burette reading, $V_2 \text{ cm}^3$
5. Find the volume of dilute acid added for complete reaction, $V_2 - V_1 \text{ cm}^3$
 - The colour change at the end point indicates that the reactant has been fully neutralised. The end point indicates the volume of alkali required to react completely with the acid.
6. Repeat the titration without adding the indicator. Add $V_2 - V_1 \text{ cm}^3$ of the dilute acid to the aqueous alkali or aqueous carbonate into a conical flask, to obtain the desired salt solution.
 - Titration repeated without indicator so that the final salt produced is not contaminated by the indicator.
7. Heat the salt solution in an evaporating dish until saturated.
 - To obtain a saturated solution for crystallisation.
8. Cool the saturated solution to allow salt crystals to form.
 - Solubility decreases as solution cools
9. Filter to collect the crystals. Wash crystals with a little cold distilled water. Dry the crystals between sheets of filter paper.
 - Use cold distilled water to minimise dissolving of the soluble salt crystal.

Method 3: Ionic Precipitation

- To prepare an insoluble salt using 2 soluble solutions.
- Cation of the salt is provided by an aqueous salt solution, usually a nitrate of the desired cation as well as all nitrate salts are soluble.
- Anion of the salt is provided by another aqueous salt solution or acid, usually a sodium salt of the desired anion as all sodium salts are soluble.
- A precipitate is formed when two clear solutions react together to form an opaque solid product
- All solutions are clear but may be coloured, e.g. "blue solution of copper(II) sulfate".
- All precipitates are opaque and may be coloured, e.g. "yellow precipitate of lead(II) iodide".

- Note: There is no such thing as a clear precipitate!

Steps

1. Using a measuring cylinder, transfer 50cm^3 of a suitable salt solution AB into a beaker
2. Add the other suitable salt solution CD into the beaker and stir. Continue to add excess solution CD until no more precipitate AD forms.
 - Add excess CD so that all AB will be reacted.
3. Filter the mixture to obtain the insoluble salt, AD as the residue
 - The filtrate obtained is salt CB and excess CD
4. Wash the residue AD with distilled water. Dry the residue between sheets of filter paper.
 - The residue is washed to remove any excess soluble CD and soluble CB.

Qualitative Analysis

Solubility of Ionic Compounds

Compound	Soluble / Insoluble	Description
Copper(II) Sulfate	Soluble	A: white powdery solid, B: blue crystalline solid
Sodium Hydroxide	Soluble	White powdery solid
Ammonium carbonate	Soluble	White crystalline solid
Silver chloride	Insoluble	White powdery solid
Barium Sulfate	Insoluble	White powdery solid
Iron(II) Sulfate	Soluble	Green crystalline solid

Cation Test

First Method: Addition of Sodium Hydroxide

Experimental procedure:

1. Add 2-3 drops of sodium hydroxide to the given solution and shake gently. Note observations.
2. Continue adding sodium hydroxide to the mixture and shake gently. Note down your observations upon addition of excess sodium hydroxide to the mixture.

Observation

- Precipitate formed?
- Colour of precipitate?

If precipitate is formed, proceed to the next step.

- Shake and observe carefully to see if the precipitate

Test for Cations - Table

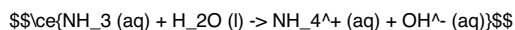
Experimental Procedure: Add 2-3 drops of aqueous sodium hydroxide to the solution **Type of Reaction:** Precipitation

Sample Solution	Cations Present	Chemical Formulae of Products
Aluminium nitrate	Al^{3+} ion	$\text{Al}(\text{OH})_3 (\text{s}) + \text{NaNO}_3 (\text{aq})$
Calcium nitrate	Ca^{2+} ion	$\text{Ca}(\text{OH})_2 (\text{s}) + \text{NaNO}_3 (\text{aq})$
Zinc nitrate	Zn^{2+} ion	$\text{Zn}(\text{OH})_2 (\text{s}) + \text{NaNO}_3 (\text{aq})$
Copper(II) nitrate	Cu^{2+} ion	$\text{Cu}(\text{OH})_2 (\text{s}) + \text{NaNO}_3 (\text{aq})$
Iron(II) nitrate	Fe^{2+} ion	$\text{Fe}(\text{OH})_2 (\text{s}) + \text{NaNO}_3 (\text{aq})$
Iron(III) nitrate	Fe^{3+} ion	$\text{Fe}(\text{OH})_3 (\text{s}) + \text{NaNO}_3 (\text{aq})$

Experimental Procedure: Add equal volumes of sodium hydroxide and test sample in the test tube and warm. **Type of Reaction:** Base + Ammonium salt

Sample Solution	Cations Present	Chemical Formulae of Products
Ammonium nitrate	NH_4^+ ion	$\text{NH}_3 (\text{g}) + \text{H}_2\text{O}(\text{l}) + \text{NaNO}_3 (\text{aq})$

Dissociation Equation for Aqueous Ammonia



As aqueous ammonia dissociates to form ammonium cations and hydroxide anions, the hydroxide anions are able to react with the metal cations to form the insoluble hydroxides which will appear in the form of precipitates.

Second Method: Addition of Aqueous Ammonia

Experimental Procedure

1. Add 2-3 drops of aqueous ammonia to the given solution and shake gently. Note down the observations.
2. Continue adding aqueous ammonia to the mixture and shake gently. Note down your observations upon addition of excess sodium hydroxide to the mixture.

Why do we need 2 reagents?

What are the observations for addition of aqueous sodium hydroxide and aqueous ammonia to calcium nitrate?

- **Observations for addition of sodium hydroxide:** White precipitate forms upon addition of aqueous sodium hydroxide. When excess aqueous sodium hydroxide is added, the white precipitate does not dissolve.
- **Observations for addition of aqueous ammonia:** No precipitate forms upon addition of aqueous ammonia.

Why is there such a difference in observation for addition of sodium hydroxide and aqueous ammonia?

- The white precipitate formed in the addition of sodium hydroxide is calcium hydroxide, $\text{Ca}(\text{OH})_2$. According to the solubility table, calcium hydroxide is partially soluble in water.
- As sodium hydroxide is a strong alkali while aqueous ammonia is a weak alkali, the concentration of OH^- ions present in the solution differs.
- Hence, the concentration of calcium hydroxide when aqueous ammonia is added is lower than that of when aqueous sodium hydroxide is added.
- At lower concentrations, the solid structure of calcium hydroxide does not become a visible size.
- Therefore, we are not able to see the formation of the precipitate.

	Experimental Procedure	Chemical Formula	Ionic Equation
1)	Add 2-3 drops of aqueous sodium hydroxide to copper(II) nitrate solution	$\text{Cu}(\text{NO}_3)_2 (\text{aq}) + \text{NaOH} (\text{aq})$	$\text{Cu}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Cu}(\text{OH})_2 (\text{s})$
2)	Add 2-3 drops of aqueous sodium hydroxide to iron(III) nitrate solution.	$\text{Fe}(\text{NO}_3)_3 (\text{aq}) + \text{NaOH} (\text{aq})$	$\text{Fe}^{3+} (\text{aq}) + 3\text{OH}^- (\text{aq}) \rightarrow \text{Fe}(\text{OH})_3 (\text{s})$

	Experimental Procedure	Chemical Formula	Ionic Equation
3)	Add 2-3 drops of aqueous sodium hydroxide to zinc nitrate solution	$\text{Zn}(\text{NO}_3)_2 (\text{aq}) + \text{NaOH} (\text{aq})$	$\text{Zn}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Zn}(\text{OH})_2 (\text{s})$
4)	Add 2-3 drops of aqueous ammonia to aluminium nitrate solution	$\text{Al}(\text{NO}_3)_3 (\text{aq}) + \text{NH}_3 (\text{g}) + \text{H}_2\text{O} (\text{l})$	$\text{Al}^{3+} (\text{aq}) + 3\text{OH}^- (\text{aq}) \rightarrow \text{Al}(\text{OH})_3 (\text{s})$
5)	Add 2-3 drops of aqueous ammonia to iron(II) nitrate solution	$\text{Fe}(\text{NO}_3)_2 (\text{aq}) + \text{NH}_3 (\text{g}) + \text{H}_2\text{O} (\text{l})$	$\text{Fe}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Fe}(\text{OH})_2 (\text{s})$

Anion Tests

	Aqueous Anions Present	Sample Solutions Containing Ions	Experimental Procedure	Chemical formula of reagents	Chemical formula of products	Type of Reaction
1)	Chlorides	Aqueous copper (II) chloride (any soluble chloride)	Add dilute nitric acid, then add silver nitrate solution	$\text{CuCl}_2 (\text{aq}) + \text{AgNO}_3 (\text{aq})$	$\text{Cu}(\text{NO}_3)_2 (\text{aq}) + \text{AgCl} (\text{s})$	Precipitation
2)	Iodides	Aqueous sodium iodide	Add dilute nitric acid, then add silver nitrate solution	$\text{NaI} (\text{aq}) + \text{AgNO}_3 (\text{aq})$	$\text{NaNO}_3 (\text{aq}) + \text{AgI} (\text{s})$	Precipitation
3)	Sulfates	Aqueous aluminium sulfate	Add dilute nitric acid, then add barium nitrate solution	$\text{Al}_2(\text{SO}_4)_3 (\text{aq}) + \text{Ba}(\text{NO}_3)_2 (\text{aq})$	$\text{Al}(\text{NO}_3)_3 (\text{aq}) + \text{BaSO}_4 (\text{s})$	Precipitation
4)	Carbonates	Aqueous ammonium carbonate	Add dilute hydrochloric acid	$(\text{NH}_4)_2\text{CO}_3 (\text{aq}) + \text{HCl} (\text{aq})$	$\text{NH}_4\text{Cl} (\text{aq}) + \text{H}_2\text{O} (\text{l}) + \text{CO}_2 (\text{g})$	reaction between acid and carbonate
5)	Nitrates	Aqueous potassium nitrate	Add dilute sodium hydroxide, then add a piece of aluminium foil. Warm carefully.			

Purpose of Adding Nitric Acid

- There may be impurities present: i.e. hydroxide and carbonate ions that can precipitate with the same test reagents and give misleading positive results.
 - Most carbonates and hydroxides are **insoluble**.
- Thus, nitric acid is added to acidify the test sample, so as to **react away any possible**

carbonates and hydroxides in the test sample.

- These carbonates and hydroxides (if present) can also precipitate with the same test reagent and give false positive results.

Testing for Water

- Copper (II) sulfate has two different forms. One is anhydrous and one is hydrated.
- The anhydrous copper(II) sulfate is a white powdery solid while hydrated copper(II) sulfate is a blue crystalline solid.

Summary

Test for anions

Anion	Test	Test Result
Carbonate	add dilute acid	effervescence, carbon dioxide produced.
Chloride	Acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
Iodide	Acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
Nitrate	Add aqueous sodium hydroxide, then aluminium foil; warm carefully	ammonia produced
Sulfate	Acidify with dilute nitric acid, then add aqueous barium sulfate	white ppt.

Test for aqueous cations

Cation	Effect of Aqueous Sodium Hydroxide	Effect of Aqueous Ammonia
Aluminium	White ppt., soluble in excess giving a colourless solution	white ppt., insoluble in excess
Ammonium	Ammonia produced on warming	-
Calcium	White ppt., insoluble in excess	No ppt.
Copper (II)	light blue ppt., insoluble in excess	light blue ppt., soluble in excess giving a dark blue solution
Iron (II)	green ppt., insoluble in excess	green ppt., insoluble in excess
Iron (III)	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
Zinc	White ppt., soluble in excess giving a colourless solution	White ppt., soluble in excess giving a colourless solution

Test for gases

Gas	Test and Test Result
Ammonia	Turns damp red litmus paper blue
Carbon dioxide	Gives white ppt. with limewater (ppt. dissolves with excess CO_2)

Gas	Test and Test Result
Chlorine	Bleaches damp litmus paper
Hydrogen	'pops' with a lighted splint
Oxygen	Relights a glowing splint
Sulfur Dioxide	Turns aqueous acidified potassium manganate (VII) from purple to colourless.

Redox Reactions

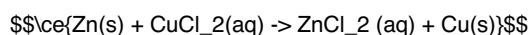
Definitions of Redox

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

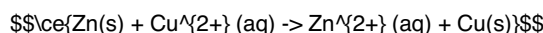
Transfer of Electrons - Definition

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

Consider the following reaction:



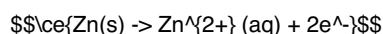
Ionic equation:



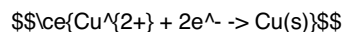
An ionic half equation shows which reactant particles gain or lose electrons to form products. To construct ionic half equations, follow these steps:

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
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.

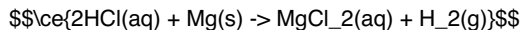
- Oxidation and reduction cannot take place independently of one another as electrons that are lost by a particle is gained by another particle.

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

Ionic half equations:

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

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

1. Elements that are uncombined with other elements are assigned an oxidation number of 0.
2. 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.

- 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.
- There are many oxidation numbers for Group 17 elements in their compounds but the usual one is -1.
- 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.

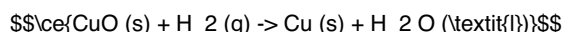
Oxidising and Reducing Agents

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

Half equation: $\text{MnO}_4^- \text{ (aq)} + 8\text{H}^+ \text{ (aq)} + 5\text{e}^- \rightarrow \text{Mn}^{2+} \text{ (aq)} + \text{H}_2\text{O (l)}$

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

Half equation: $2\text{I}^- \text{ (aq)} \rightarrow \text{I}_2 \text{ (s)} + 2\text{e}^-$

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

Electrochemistry

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

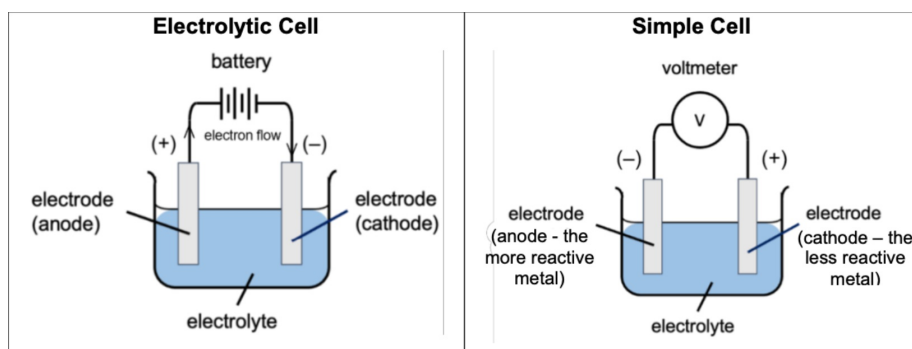
- A rod or plate which conducts electricity through the electrolyte
- **ANODE - OXIDATION** takes place at this electrode
- **CATHODE - REDUCTION** takes place at this electrode
- Types of Electrodes:
 - **Inert Electrode** - An electrode that does not take part in chemical reactions during electrolysis, 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?

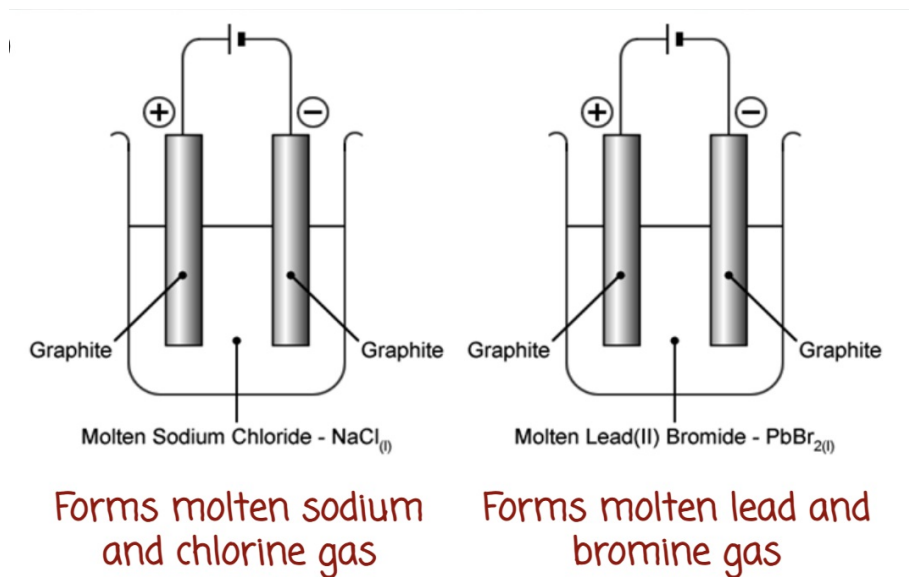
- Opposite polarity for anode and cathode due to difference in energy conversion
- Oxidation always occurs at anode, reduction always occurs at cathode.

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

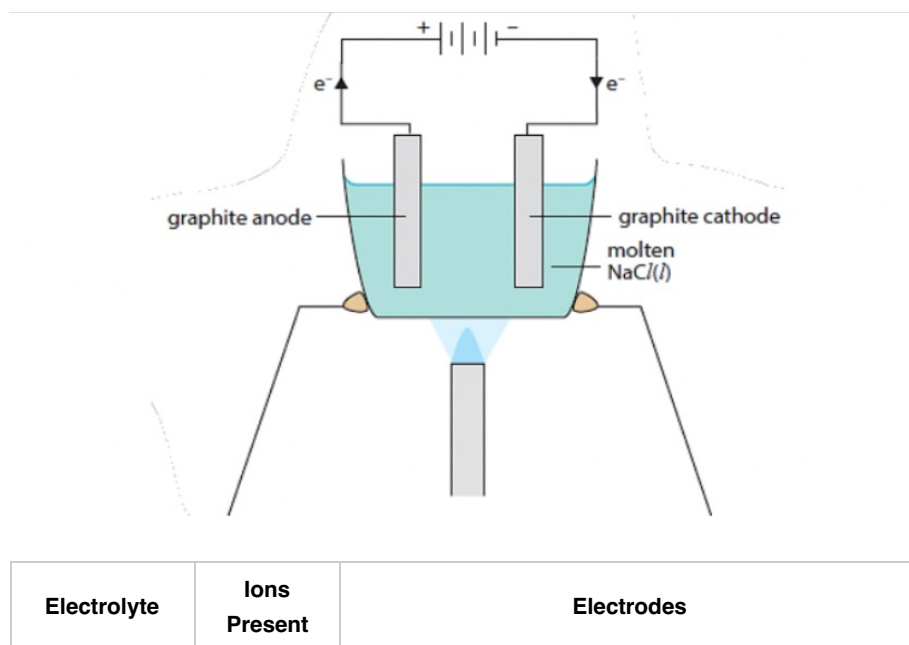
Electrolysis is the decomposition of a compound by electricity.



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.

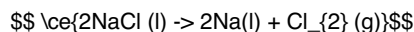
Anode (+) (Oxidation)

- 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

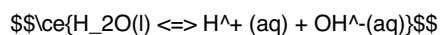


Observations

- 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.

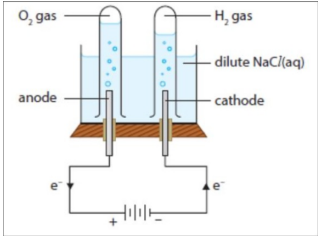
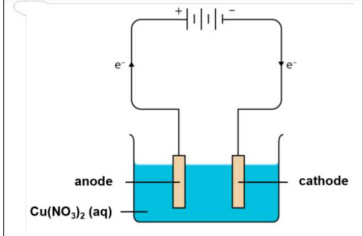
Selective discharge of cations (reactivity series)

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^-	

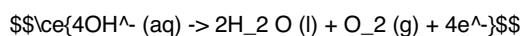
Comparing Dilute Sodium Chloride vs Dilute Copper(II) Nitrate

	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)

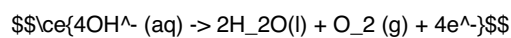
Dilute Sodium Chloride

- 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

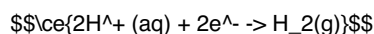
- NO_3^- and OH^- attracted to anode.
- OH^- is lower than NO_3^- in electrochemical series, hence OH^- is selectively discharged.
- Half equation:



Cathode (-) (Reduction)

Dilute Sodium Chloride

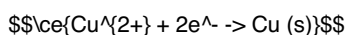
- Na^+ and H^+ attracted to cathode
- Hydrogen is less reactive than sodium, hence H^+ is selectively discharged
- Half equation:



Dilute Copper(II) Nitrate

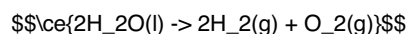
- Cu^{2+} and H^+ attracted to cathode
- Copper is less reactive than hydrogen, hence copper is selectively discharged.

- Half equation:

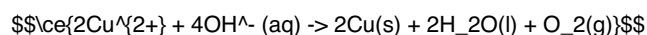


Overall Equation

Dilute Sodium Chloride



Dilute Copper(II) Nitrate



Observations

Dilute Sodium Chloride

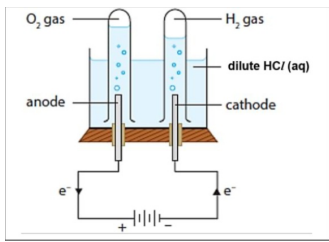
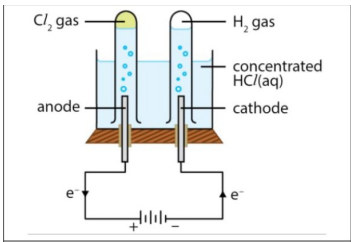
1. Solution remains neutral as both H^{+} and OH^{-} ions are discharged.
2. Water levels drops
3. Effervescence of colourless gas at both electrodes. Volume of hydrogen gas collected at the cathode is twice the volume of oxygen gas collected at the anode. For every 4 moles of electrons transferred,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

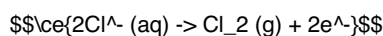
- In a concentrated solution, the ease of discharge of Cl^{-} , Br^{-} , or I^{-} anions increases.

	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)

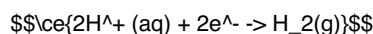
Anode (+) (Oxidation - Concentrated)

- 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:

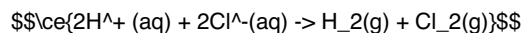


Cathode (-) (Reduction - Concentrated)

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



Overall Equation - Concentrated



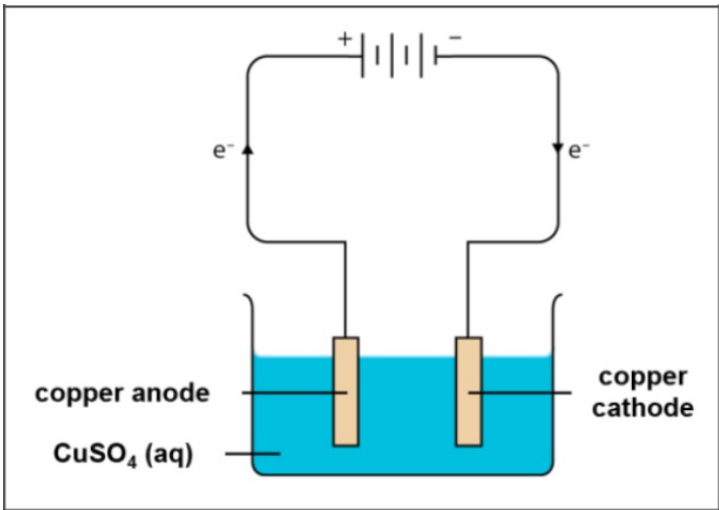
Observations - Concentrated

1. Solution becomes less acidic as there is a net discharge of H^+ while OH^- ions remain in solution.
2. 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

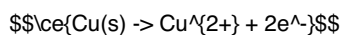
- If the electrode used is not graphite or platinum but another metal, the electrode is reactive and is oxidised as the anode during electrolysis.

	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)

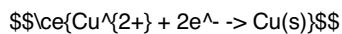
Anode (+) (Oxidation)

- 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)

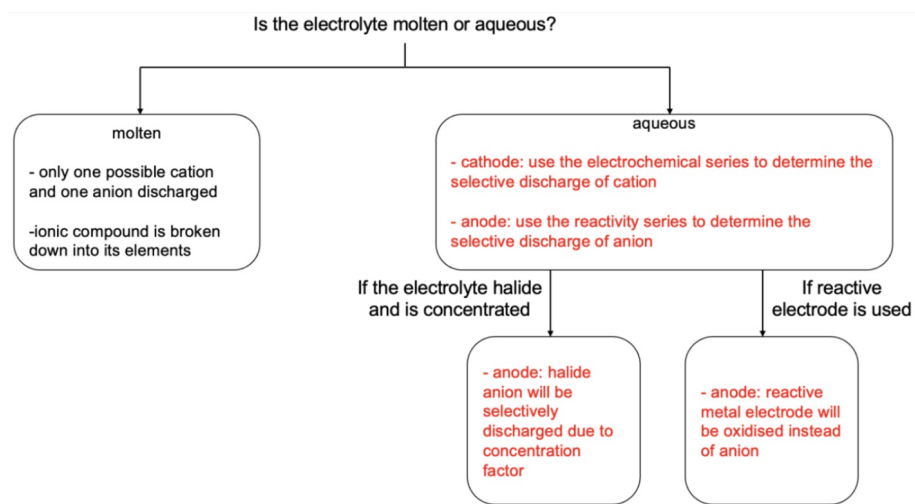
- Cu^{2+} and H^+ attracted to cathode
- Copper is less reactive than hydrogen, hence Cu^{2+} is selectively discharged.
- Half equation:



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.

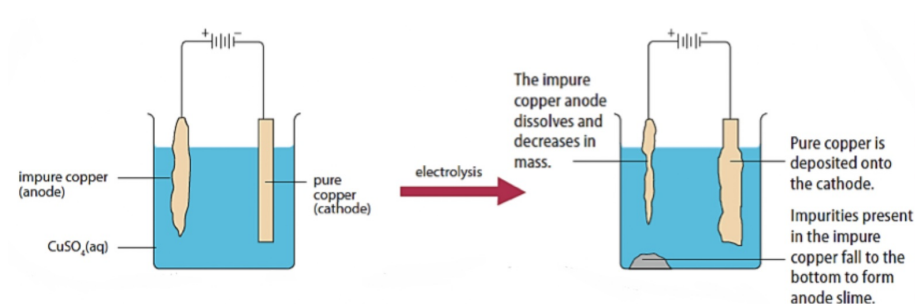
Summary



Applications of Electrolysis

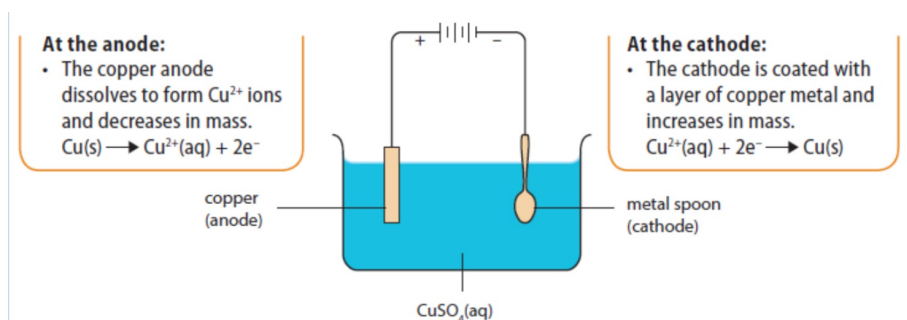
Metal Purification

- 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

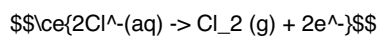
Suggest why scientists want to make hydrogen from sea water.

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.

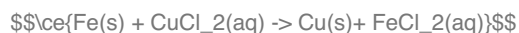
- $\text{Pb} + 2\text{HCl} \rightarrow \text{PbCl}_2 + \text{H}_2$
- $\text{Pb} + \text{H}_2\text{SO}_4 \rightarrow \text{PbSO}_4 + \text{H}_2$

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.

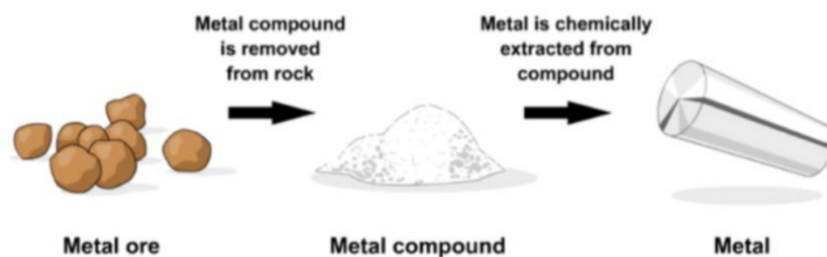
Hence, the more reactive metals form carbonates that are more thermally stable.

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)

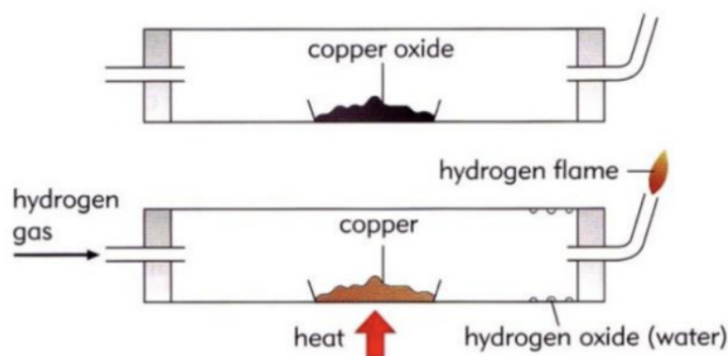
Explain why the less reactive metals (metals below carbon) can be extracted by reduction of their oxides by carbon (displacement reaction).

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

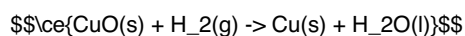
Some metals can be extracted from their metal oxide ores by the reduction with hydrogen gas.



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

Copper < Silver < Gold (easiest)

Chemical Equation:

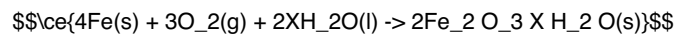


Rusting of Iron

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

Ammonia

Recap on Ammonia

Chemical Formula: NH_3

Weak Alkalis

- A weak alkali partially dissociates in water to form hydroxide ions. $\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$

Conductivity of Ammonia

- Ammonia **cannot conduct electricity in solid, liquid, and gaseous states** due to **absence of mobile ions and mobile electrons** to carry charges.
- It is able to conduct electricity in **aqueous state** due to **the formation of ammonium and hydroxide ions**.

Testing for Ammonia

- Place a **piece of moist/damp red litmus paper** at the mouth of the test paper
- Colorless pungent gas **turns moist/damp red litmus paper blue**.

Uses of Ammonia

Ammonia is a very important chemical, used in the manufacture of:

- Fertilisers
- Nitric acid
- Household cleaning products

Nitrogen is an essential element required for the survival of plants. Even with an abundance of 78% nitrogen in the air, there is a lack of nitrogen for crop growth. Why is this so?

- Though N_2 is present in a large percentage of air, plants are unable to absorb N_2 from the air directly. N_2 has to be dissolved in water before plants can absorb them through the roots.
- (Nitrogen is dissolved naturally when lightning occurs and by certain types of bacteria in soil but these processes do not occur fast enough to support current crop growth.)

Common Artificial Fertilisers

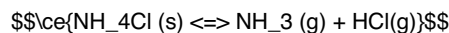
- Ammonium Nitrate, NH_4NO_3
- Ammonium Sulfate, $(\text{NH}_4)_2\text{SO}_4$

Reversible Reactions

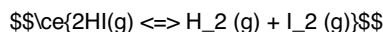
Reversible reactions are chemical reactions that can proceed in both the forward and backward directions.

Examples:

Thermal decomposition of ammonium chloride



Formation and decomposition of hydrogen iodide



In reversible reactions, the reactants can change into products and the products can change back into reactants.

Note

Note the type of arrows to denote reversible chemical reactions.

Production of Ammonia

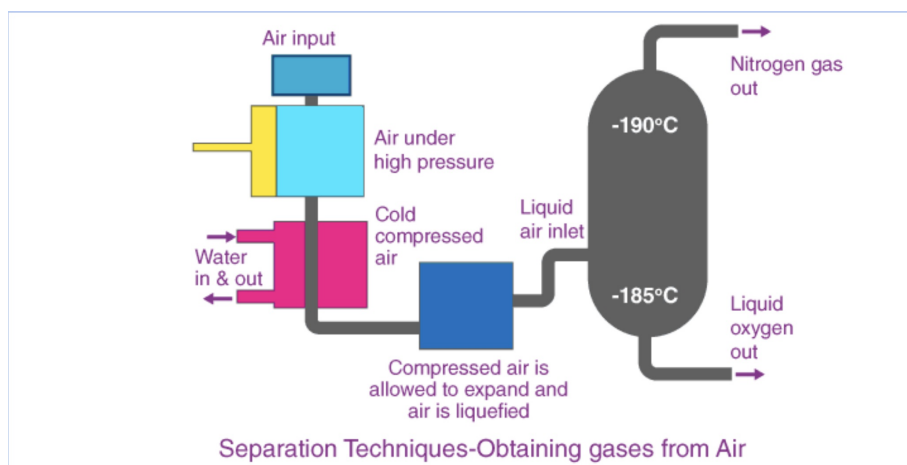
General Word Equation for Reaction Between Ammonium Salt and Base

Alkali + ammonium salt → salt + water + ammonia

Raw materials for Haber Process

- The raw materials in the manufacture of ammonia are **nitrogen** and **hydrogen**.
- The nitrogen is obtained by the **fractional distillation of liquefied air**.
- The hydrogen is obtained from a variety of sources:
 - Reacting together methane and steam
 - methane + steam → carbon monoxide + hydrogen gas
 - $\text{CH}_4 \text{ (g)} + \text{H}_2\text{O (g)} \rightarrow \text{CO (g)} + 3\text{H}_2 \text{ (g)}$

Fractional Distillation of Liquefied Air



- Cracking hydrocarbons found in crude oil:
 - butane → ethene + hydrogen
 - $\text{C}_4\text{H}_{10}(\text{g}) \rightarrow 2\text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g})$

Conditions Suitable for Reversible Reactions

Conditions such as temperature and pressure need to be adjusted for reversible reactions.

Function of the Iron Catalyst

- The iron catalyst is used to increase the rate of reaction
- It increases the speed of reaction by providing an **alternative reaction pathway** with a **lower activation energy**
- The finely divided iron catalyst provides a **larger surface area to volume ratio** for the reaction to take place, further increasing the rate of reaction.

Why is ammonia gas condensed into a liquid and not transported in gaseous state?

- Compressing the gas to liquid state allows more ammonia to be stored per volume
- It also allows for easier storage and transportation of ammonia gas.