Sulphuric Acid

Sulphuric acid is the compound involves combination of hydrogen, oxygen and sulphur (H_2SO_4).

Manufacture of Sulphuric Acid

Sulphuric acid is produced in large scale by using contact process. Contact process involves four major stages

Stages Involves In Contact Process

i. Production of sulphur dioxide
ii. Purification of sulphur dioxide
iii. Catalytic conversion of sulphur dioxide to sulphur trioxide
iv. Conversion of sulphur trioxide to sulphuric acid

Production of Sulphur Dioxide

i. The sulphur dioxide may be obtained by Burning sulphur in air.

$S(s) + 0_2(g) \quad S0_2(g)$

NB: This is the most convenient method of producing sulphur dioxide.

ii. Burning sulphide ores such as iron pyrite (FeS2) and zinc blende (ZnS)

$4FeS_{2(s)} + 110_{2(g)} \quad 8SO_{2(g)} + 2Fe_2O_{3(s)}$

$2ZnS(s) + 3O_2(g) \quad 2S0_2(g) + 2ZnO(s)$

Sulphur dioxide gas is produced as a by-product (secondary product)

Purification of Sulphur Dioxide

Sulphur dioxide obtained in the first stage is purified by remove impurities, such as dust and arsenic (III) oxide, which may poison catalyst in the next stage. Also is passed through conc. H_2SO_4 for drying sulphur dioxide.

Catalytic Conversion of Sulphur Dioxide to Sulphur Trioxide

Dry sulphur dioxide is heated at 4500C at normal atmospheric pressure in presence of catalyst, may be **vanadium pentaoxide** or **platinum**. Since vanadium pentaoxide is cheapest it preferred. Reaction is reversible

$2SO_2(g) + 0_2(g)\ 2 \quad 2SO_3(s) + \text{heat}$

Conversion of Sulphur Trioxide to Sulphuric Acid

The sulphur trioxide from the conversion chamber is passed through a heat exchanger to remove excess heat. It is then taken to an absorption tower where it is dissolved in concentrated sulphuric acid to form **oleum** or **fuming sulphuric acid**.

$H_2S0_4(l) + S0_3(g) \quad H_2S_20_7(l)$

Oleum is then diluted to give concentrated sulphuric acid.

$H_2S_20_7(l) + H_20(l) \quad 2H_2S0_4(l)$

Nb: Sulphur trioxide cannot be dissolved directly in water to form sulphuric acid, the reaction is highly exothermic and the heat produced vaporizes the sulphuric acid formed. This makes it difficult to collect the acid

Physical Properties of Conc. H_2s0_4

i. It is a dense oily liquid. It is sometimes referred to as oil of vitriol.
ii. It is a colourless liquid with a specific gravity of 1.84 g/cm3.
iii. It has a boiling point of 333°C. It decomposes at this temperature to give sulphur dioxide gas and water.

Conc. $H_2S0_4(l) \quad S0_3(g) + H_20(g)$

iv. Concentrated sulphuric acid has no effect on litmus paper
v. It does not conduct electricity
vi. It does not give hydrogen when reacted with metals. This is because the acid is a covalent compound and it is not ionized.

Nb: The properties of concentrated and dilute sulphuric acid are not the same

Chemical Properties of Dil. H_2s0_4

i. **Reaction with metals**

Dilute sulphuric acid reacts with common metals such as magnesium, zinc and iron to form a sulphate of the metal and hydrogen gas

$Mg(s) + H_2S0_{4(aq)} \quad MgSO_4(aq) + H_{2(g)}$

$Cu(s) + H_2S0_{4(aq)} \quad CuSO_4(aq) + H_{2(g)}$

$Fe(s) + H_2S0_{4(aq)} \quad FeSO_4(aq) + H_{2(g)}$

ii. **Reaction with metal hydroxides**

The reactions between dilute sulphuric acid and metal hydroxides are neutralization reactions.

$2NaOH\ (aq) + H_2S0_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2H_20(l)$

iii. **R eaction with metal carbonates**

Dilute sulphuric acid reacts with metal carbonates to give a metal sulphate, carbon dioxide and water.

$Na_2CO_3\ (aq) + H_2S0_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + C0_2\ (g) + H_20\ (1)$

$PbCO_3(s) + H_2S0_4\ (aq) \longrightarrow PbSO_4(s) + C0_2\ (g) + H_20\ (l)$

Chemical Properties of Conc. H_2S0_4

In chemical reactions, concentrated sulphuric acid can act as a dehydrating agent, a drying agent or an oxidizing agent.

i. **Sulphuric acid as a dehydrating agent**

As a dehydrating agent, sulphuric acid removes the elements of water (oxygen and hydrogen) from a compound to form a new compound.

$C_{12}H_{22}0_{11}(s) + H_2S0_4 \longrightarrow 12C(s) + 11H_20(g)$
(Sugar) (Black)

The reaction is highly exothermic.

$CuSO_4.5H_20(s) + H_2S0_4 \longrightarrow CuSO_4(s) + 5H_20(l)$
(Blue) (White)

ii. **Sulphuric acid as a drying agent**

As a drying agent, concentrated sulphuric acid absorbs traces of water from substances. It is especially used as a drying agent during the laboratory preparation of gases, with the exception of ammonia and carbon dioxide

iii. **S ulphuric acid as an oxidizing agent**

Hot concentrated sulphuric acid is a strong oxidizing agent. It oxidizes both metals and non-metals while it is reduced to sulphur dioxide.

$C(s) + 2H_2S0_4(l) \longrightarrow C0_2(g) + 2S0_2(g) + 2H_20(1)$

$S(s) + 2H_2S0_4(l) \longrightarrow 3S0_2(g)\ t\ 2H_20(l)$

NB: The orange colour of the dichromate (VI) paper changes to green. This confirms the presence of sulphur dioxide gas

Uses of Sulphuric Acid

i. **Extraction of metals**

Sulphuric acid is used in large quantities in the iron and steel-making industry to remove rust and scale from rolled iron sheets

ii. **Manufacture of fertilizers**

Ammonium sulphate, an important nitrogenous as fertilizer, is commonly produced as a by-product in the production of coke from coal. The ammonia produced in the thermal decomposition of coal is reacted with waste sulphuric acid to produce ammonium sulphate

iii. **M anufacture of aluminium sulphate**

Sulphuric acid is used in the manufacture of aluminium sulphate, which is used in water treatment plants to filter impurities and to improve the taste of the water. Aluminium sulphate is made by reacting 'bauxite with sulphuric acid

iv. **Refining crude oil**

A large quantity of sulphuric acid is used in refining petroleum. The acid is used as a catalyst for the reaction of isobutane with isobutylene to give isooctane, a compound that raises the octane rating of petrol

v. Sulphuric acid is used in lead-acid (car) batteries

vi. Sulphuric acid is used as a dehydrating agent in its concentrated form

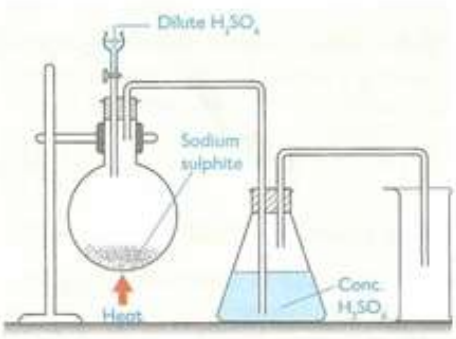
vii. Sulphuric acid is used in the manufacture of a wide range of pigments

Sulphur Dioxide

By Defn: Sulphur dioxide is a binary compound of sulphur with oxygen.

Preparation of Sulphur Dioxide

1. When sulphur burns in air (oxygen) it gives sulphur dioxide.
   $S(s) + 0_2(g) \longrightarrow S0_2\ (g)$
2. In the laboratory, sulphur dioxide is prepared by reacting a sulphite or a hydrogen sulphite with an acid.



$Na_2SO_3(s) + H_2S0_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + H_20\ (l) + S02\ (g)$

Nb:

i. Conc. Sulphuric acid in the conical flask for dry chlorine gas
ii. It collected down ward delivery because it is denser than air

3. In the laboratory, sulphur dioxide is prepared by the reaction between copper turnings and concentrated sulphuric acid can be represented by the following equation.
   $Cu(s) + 2H_2S0_4\ (l) \longrightarrow CuSO_4\ (s) + S0_2\ (g) + 2H_20(l)$

Physical Properties of Sulphur Dioxide

i. It is a colourless gas with an irritating chocking smell.
ii. It is poisonous and should therefore be prepared in the fume chamber.
iii. It is two and a half times denser than air.
iv. It is readily liquefied. Liquid sulphur dioxide boils at about -10°C.

Chemical Properties of Sulphur Dioxide

i. Sulphur dioxide is an acidic gas. Dissolve in water to form acidic solution of sulphuric acid which is commonly known as **sulphurous acid**.
   $S0_2(g) + H_20(l) \longrightarrow H_2S0_3(aq)$

ii. **Solubility of sulphur dioxide**

Sulphur dioxide has a high solubility in water. It dissolves to form sulphurous acid. 1 cm_3 of water dissolves about 700 cm_3 of sulphur dioxide

iii. **Reducing property of sulphur dioxide**

Sulphur dioxide is a strong reducing agent. It reduces potassium permanganate and potassium dichromate (VI) solutions to manganese sulphate and chromic sulphate respectively. In moist conditions, sulphur dioxide is a **bleaching agent**.

$2KMnO_4\ (aq) + 5S0_2\ (g) + 2H_2O\ (l) \longrightarrow K_2SO_4\ (aq) + 2MnSO_4\ (aq) + 2H_2S0_4\ (aq)$

Where:

$2KMnO_4$ = purple
$2MnSO_4$ = colourless

$K_2Cr_2O_7\ (aq) + H_2S0_4\ (aq) + 3S0_2\ (g) \longrightarrow K_2S0_4\ (aq) + Cr_2(S0_4)_3(aq) + H_20(l)$

Where:

$K_2Cr_2O_7$ = orange
$Cr_2(S0_4)_3$ = green

iv. **Oxidizing property of sulphur dioxide**

In the presence of moisture Sulphur dioxide is also an oxidizing agent. Consider the chemical reaction below.

$FeS(s) + 2HCl(aq) \longrightarrow HS(g) + FeCl_2(aq)$
$2H_2S(g) + S0_2(g) \longrightarrow 3S(s) + 2H_20(l)$

v. **Reaction of sulphur dioxide with oxygen**

Sulphur dioxide does not burn. However, it combines with oxygen in the presence of a heated catalyst (platinized asbestos) to give sulphur trioxide.

$2S0_2\ (g) + 0_2(g) \longrightarrow 2S0_3\ (g)$

Test for Sulphur Dioxide

i. The presence of sulphur dioxide can be tested using a filter paper that has been soaked in acidified potassium dichromate (VI). If sulphur dioxide is present, the colour of the paper changes from orange to green due to the reduction of dichromate (VI) to chromate (III).
ii. Sulphur dioxide also decolorizes acidified potassium permanganate solution.

Pollution Effects of Sulphur Dioxide

i. Soil erosion because the major sources of sulphur dioxide in the air are power plants that use fossil fuels such as coal and diesel, industrial boilers, and exhaust emissions from motor vehicles.
ii. It causes acidic rain.
iii. It can cause impairment of respiratory function and heart diseases

Uses of Sulphur Dioxide

i. The main use of sulphur dioxide is in the manufacture of sulphuric acid through the contact process.
ii. It is used as a bleaching agent for fibres, mainly of animal origin, for example wool, silk, straw and sponges.
iii. It is used in the manufacture of calcium and sodium hydrogensulphites. Calcium hydrogensulphite ($CaHSO_3$) is used for bleaching wood-pulp in the manufacture of paper and artificial silk. Sodium hydrogensulphite, $NaHSO_3$, is used in the manufacture of sodium sulphinate, which is a reducing agent used in dyeing.
iv. It is used for fumigating houses and clothing to kill micro-organisms. It is also used to kill insects such as white ants because it is poisonous.
v. Sulphur dioxide is used in small doses as a preservative of some liquids such as lemon and orange juices. This is because even a small concentration of the gas prevents fermentation of the liquids as it reacts with oxygen. This prevents oxidation of the liquids.
vi. Liquid sulphur dioxide is used in refrigerators because it liquefies at three atmospheres at room temperature.

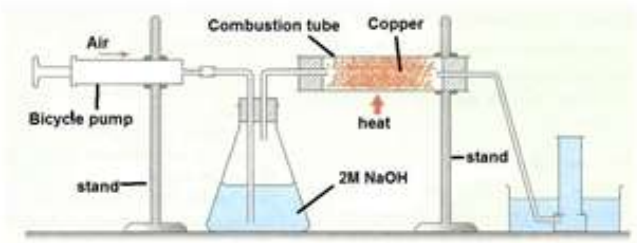
Nitrogen

Nitrogen exists in gaseous state as a diatomic molecule (N_2). Nitrogen gas is odourless and colourless. Nitrogen also occurs in combined state in the form of nitrates and oxides. It is also found in plants and animals as a constituent of proteins.

Preparation of Nitrogen Gas

Nitrogen prepared in the laboratory by isolation from atmospheric air.

Diagram:



When air is passed through a solution of sodium hydroxide, carbon dioxide from the air dissolves in the alkali which removes carbon dioxide from the air.

$NaOH\ (aq) + C0_2\ (g) \longrightarrow NaHCO_3\ (aq)$

In the combustion tube, the hot copper turnings (brown) react with oxygen to form copper (II) oxide (black) where oxygen is removed from the air

$2Cu(s) + 0_2(g) \longrightarrow 2CuO(s)$

Nb:

i. The nitrogen obtained is impure which contains inert gases such as **neon** and **argon** makes the gas impure
ii. Dry nitrogen gas can be obtained by passing it through concentrated sulphuric acid instead of collecting it over water

Physical Properties of Nitrogen

i. Nitrogen is a colourless, odourless and tasteless gas.
ii. It is almost insoluble in water.
iii. It has a boiling point of -196°C.
iv. It can be liquefied to form liquid nitrogen.
v. Nitrogen prepared from air is denser than pure nitrogen since it contains the noble gases
vi. It occupies about 78% by volume in the atmosphere

Chemical Properties of Nitrogen

Nitrogen is relatively Unreactive

Nitrogen gas is stable below 3000°C. It only takes part in reactions at very high temperatures.

Reaction with oxygen

Nitrogen does not burn nor does it support combustion. When heated, the gas combines with oxygen to form nitrogen monoxide gas.

N_2 (g) + 0_2(g) ⟶ 2N0 (g)

Reaction with hydrogen

Nitrogen reacts with hydrogen when heated to form ammonia.

N_2 (g) + $3H_2$ (g) ⟶ $2NH_3$ (g)

Reaction with metals

When heated together with metals, nitrogen forms **metal nitrides**. For example, with magnesium it forms magnesium nitride.

3Mg(s) + N_2 (g) heat ⟶ Mg_3N_2 (g)

Uses of Nitrogen

The following are some uses of nitrogen gas.

1. Manufacture of fertilizers

Nitrogen is used to manufacture nitrogenous fertilizers. These include Diammonium phosphate (DAP), calcium ammonium nitrate (CAN), ammonium superphosphate (ASP), ammonium nitrate (AN), ammonium phosphate sulphate (APS), ammonium sulphate nitrate (ASN), and ammonium sulphate (AS) and urea.

2. Refrigeration

Nitrogen gas is used as a refrigerant because of its low boiling point (inert atmosphere 196°C).

3. Processing reactive substances

Because of its low reactivity; nitrogen is used to provide an inert atmosphere for storing and processing reactive substances.

4. Plastic industries

Nitrogen is used in the manufacture of synthetic fibres such as polyamides. Polyamides are commonly known as nylons. Nylons are chemically inert and are stronger than natural fibres. They are used in making fishing nets, clothes and many other items.

5. Manufacture of ammonia

Nitrogen is used in the manufacture of ammonia through the Haber process. Reaction is reversible

N_2 (g) + $3H_2$ (g) ⟶ $2NH_3$ (g) + heat

6. Manufacture of nitric acid

The ammonia gas manufactured in the Haber process is used in the manufacture of nitric acid by catalytic oxidation.

This Process Involves Three Main Stages.

i. Catalytic oxidation of ammonia. platinum gauze used as catalyst at 700°C

$4NH_{3(g)}$ +$50_{2(g)}$ ⟶ $4N0_{(g)}$ + $6H_20_{(g)}$ + Heat

ii. Oxidation of nitrogen monoxide to nitrogen dioxide at 30°C. Reaction is reversible

2N0 (g) + $0_{2(g)}$ ⟶ $2N0_2$ (g)

iii. Reaction of nitrogen dioxide with water to form nitric acid.

$4N0_2$ (g) + $2H_2O$ (l) + O_2 (g) ⟶ $4HN0_3$ (aq)

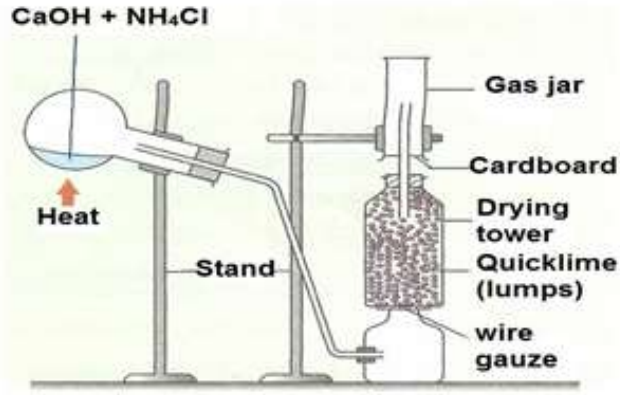
Ammonia

Ammonia is a compound of hydrogen and nitrogen. Its chemical formula is NH_3. It exists in nitrogenous organic materials such as hoofs and horns of animals.

Preparation of Ammonia

Ammonia can be prepared in the laboratory by heating any ammonium salt together with an alkali. The most commonly used alkalis are potassium hydroxide with ammonium chloride.

Diagram:



Calcium hydroxide reacts with ammonium chloride to produce ammonia gas, calcium chloride and water

$2NH_4Cl(s) + Ca(OH)_2(s) \longrightarrow CaCl_2(aq) + 2H_2O(l) + 2NH_3(g)$

Ammonia gas is dried by passing it over quicklime

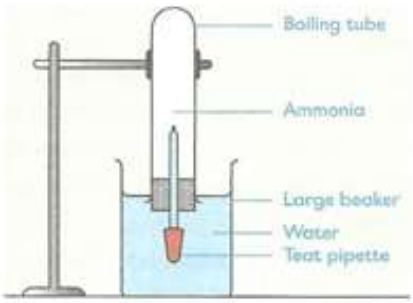
Nb:

i. The round-bottomed flask is tilted to prevent any condensed water from running back into the hot flask, which would make the flask crack.
ii. The common drying agents such as concentrated sulphuric acid and calcium chloride are not used because they react with the gas. Reactions is reversible
$2NH_3 (g) + H_2S0_4 (l) \longrightarrow (NH_4)_2S0_4 (l)$
$8NH_3 (g) + CaCl_2(s) \longrightarrow CaCl_2.8NH_3(s)$
iii. A mmonia is collected by upward delivery or downward displacement of air. This is because it is less dense than air.
iv. Ammonia is an alkaline gas and turns wet red litmus paper blue.
v. When sodium hydroxide or potassium hydroxide is used, they are used in solution form because they are very reactive in solid form

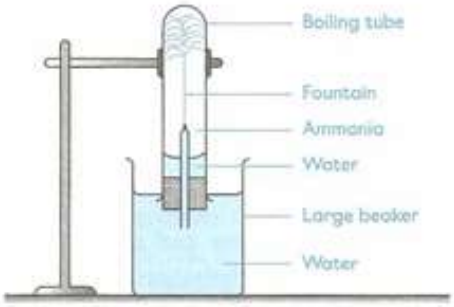
Physical Properties of Ammonia

i. Ammonia is a colourless gas with a pungent choking smell
ii. It is less dense than air
iii. A mmonia is highly soluble in water

Fountain Experiment



When a drop of water is released into the boiling tube containing ammonia, the water dissolves most of the ammonia gas, thus leaving a partial vacuum. This lowers the pressure inside the boiling tube. The rubber remains pressed inside because of atmospheric pressure.



When the teat is removed, the water in the beaker rushes into the boiling tube, thus dissolving the remaining ammonia gas. The water forms a fountain, thus the name of the experiment.

Nb:

i. The solution of ammonia which is a base is called aqueous ammonia.
ii. In aqueous ammonia, the molecules of ammonia continuously interact with water molecules to produce ammonium ions ($NH4^+$) and hydroxyl ions (OH^-). Reaction is reversible
$NH_3 (g) + H_20 (l) \longrightarrow NH_4^+ (aq) + OH^-(aq)$
iii. A mmonium hydroxide does not exist as a molecule. Instead it exists as NH_4^+ and $0H^-$ ions.
iv. The presence of the hydroxyl ions in aqueous ammonia, OH^-, makes the solution alkaline.

Chemical Properties of Ammonia

Reaction with concentrated hydrochloric acid

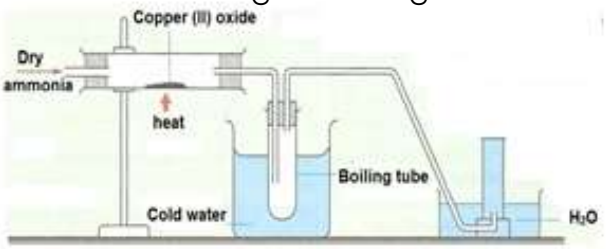
When a glass rod dipped in concentrated hydrochloric acid is placed at the mouth of a gas jar containing ammonia, dense white fumes of ammonium chloride are formed.

NH_3 (g) + HCl (g) ⟶ NH_4Cl(s)

This is used as a test for ammonia gas.

Oxidation of ammonia

Oxidizing agents, such as copper (II) oxide, oxidize ammonia gas to nitrogen.



When ammonia gas is passed over heated copper (II) oxide, the gas is oxidized to nitrogen while the copper (II) oxide is reduced to copper.

$3CuO(s) + 2NH_3(g) \longrightarrow 3Cu(s) + 3H_20(l) + N_2(g)$

The liquid collected in the test tube is water. The liquid changes white anhydrous copper (II) Sulphate to blue or blue cobalt (II) chloride paper to pink.

Uses of Ammonia

The following are some of the uses of ammonia

1. Manufacture of fertilizers

Ammonia is used in the manufacture of nitrogenous fertilizers such as ammonium sulphate nitrate (ASN), ammonium sulphate (AS), di-ammonium phosphate (DAP), calcium ammonium nitrate (CAN) and ammonium nitrate (AN).

2. Cleaning

Ammonia softens water and neutralizes acid stains caused by perspiration, thus making washing easier.

3. Refrigeration

Liquid ammonia can be used as a refrigerant because it is highly volatile.

4. Manufacture of nitric acid

Nitric acid is manufactured by the catalytic oxidation of ammonia

Carbon

Carbon is a Group IV and Period 2 element. It usually forms covalent bonds when combining with other elements. Carbon is found in a variety of natural substances. These include shells, coal, diamond, and graphite. It is also found in compounds of carbonates such as chalk ($CaCO_3$).

Allotropes of Carbon

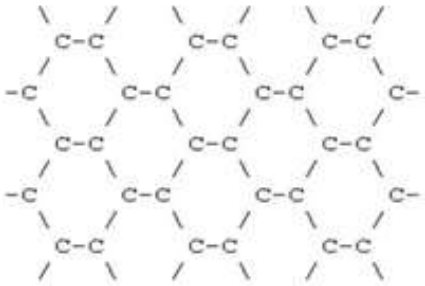
Carbon exists in three main forms include

i. Graphite Carbon
ii. Diamond Carbon
iii. A morphous Carbon

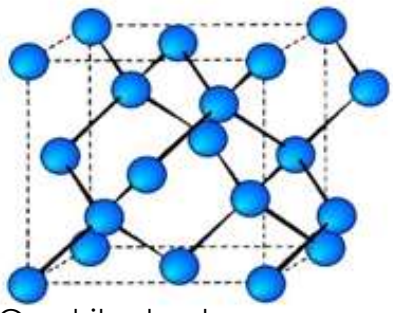
Graphite Carbon

Each carbon atom is bonded to three other carbon atoms, giving it a **trigonal structure**. Three out of the four valence electrons of each carbon atom are used in bond formation. The fourth electron is referred to as a delocalized electron, and is free to move in the graphite structure. The trigonal units come together to form a hexagonal ring. These rings form flat parallel layers, one over the other. The layers can slide over one another. This makes graphite soft and slippery.

Diagram:



Bond formation in a graphite carbon



Graphite structure

Physical Properties of Graphite Carbon

i. It has a low density
ii. It is soft and greasy.
iii. G ood conduct of heat and electricity due to the delocalized of electrons.

iv. It has relatively low melting and boiling points (compared to diamond).

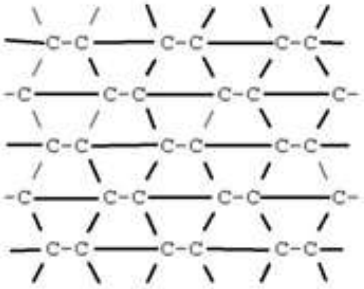
Uses of Graphite Carbon

i. It is used as a lubricant in high temperature processes because of its slippery nature and high boiling point.
ii. It is used to make electrodes due to its good electrical conductivity
iii. It is mixed with clay to make the lead in pencils.

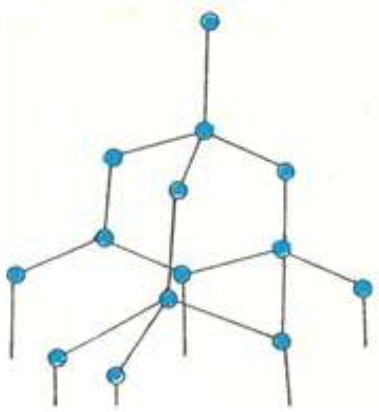
Diamond Carbon

In diamond, all the four valence electrons of a carbon atom are used in bonding, forming four strong covalent bonds in each carbon atom. The electrons form a tetrahedral shape.

Diagram:



Bond formation in a diamond carbon



Diamond structure

Physical Properties of Diamond Carbon

i. It has a high density compared to graphite.
ii. It is the hardest natural substance known.
iii. It is a bad conductor of electricity and heat.
iv. It has a high melting point of 3,550°C and a high boiling point of 4,289°C.
v. It has a high refractive index of 2.45. The high refractive index results in high dispersion of light, making it suitable for use in jewellery

Uses of Diamond Carbon

i. It is used in making jewellery
ii. It is used to make glass cutters and rock borers because of its hardness.

Amorphous Carbon

Amorphous carbon is carbon that does not have any clear shape, form or crystalline structure. Amorphous carbon is made up of tiny bits of graphite with varying amounts of other elements considered as impurities. It is formed when a material containing carbon is burned in limited supply of oxygen, resulting in incomplete combustion.

Example of Amorphous Carbon

i. Sugar charcoal -made by dehydrating sugar
ii. Wood charcoal - made by heating wood in a limited supply of air
iii. C oal

Uses of Amorphous Carbon

i. it used to make ink, paint and rubber products.
ii. It is also pressed into shapes and used to form cores of dry batteries.
iii. U sed as fuel for cooking example wood charcoal and coal

Chemical Properties of Carbon

The following are some chemical properties of carbon.

1. **Carbon burns in excess oxygen to form carbon dioxide gas.**
   2. $C(s) + 0_2(g) \longrightarrow C0_2\ (g)$
   In insufficient oxygen, carbon monoxide gas is formed.
   $2C(s) + 0_2(g) \longrightarrow 2C0\ (g)$

2. **Carbon has a high affinity for oxygen and thus acts as a reducing agent.**
   Carbon reduces oxides of less reactive metals to their respective metals.
   $ZnO(s) + C(s) \longrightarrow Zn(s) + CO\ (g)$
   $PbO(s) + C(s) \longrightarrow Pb(s) + CO\ (g)$

   Carbon reduces hot concentrated nitric acid and concentrated sulphuric acid.
   $4HN0_3\ (l) + C(s) \longrightarrow 4N0_2\ (g) + C0_2\ (g) + 2H_20(l)$

$2H_2S0_4$ (l) + C(s) ⟶ $2S0_2$ (g) + $C0_2$ (g) + $2H_20$(l)

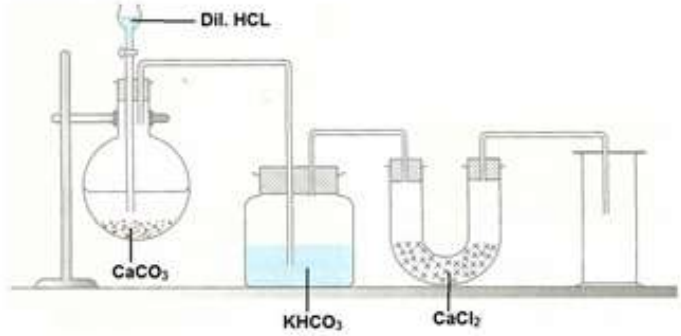
Carbon Dioxide

Carbon dioxide is the covalent compound formed between carbon and oxygen

Preparation of Carbon Dioxide

Carbon dioxide can be prepared in the laboratory by the action of dilute hydrochloric acid on marble chips (calcium carbonate).

Diagram:



Dilute hydrochloric acid reacts with marble chips to give calcium chloride, water and carbon dioxide.

$CaCO_3$(s) + 2HCl (aq) ⟶ $CaCl_2$ (aq) + H_20 (l) + $C0_2$ (g)

The potassium hydrogen carbonate solution absorbs any traces of hydrochloric acid from the carbon dioxide. The gas is then dried by passing it through anhydrous calcium chloride. Carbon dioxide is collected by downward delivery because it is denser than air.

Nb:

Other carbonates, such as magnesium carbonate can also be used to prepare carbon dioxide gas.

Physical Properties of Carbon Dioxide Gas

i. Carbon dioxide is a colourless and odourless gas.
ii. It has a melting point of -199°C and a boiling point of -91.5°C.
iii. The gas is denser than air.
iv. Solid carbon dioxide is referred to as dry ice. Dry ice sublimes means changing directly to gas at atmospheric pressure.

Chemical Properties

Carbon dioxide does not support combustion. The gas is slightly acidic. As a

result, it may not have any observable effect on a litmus paper.

Reaction of carbon dioxide with lime water

Carbon dioxide react with Lime water (calcium hydroxide) a white precipitate of calcium carbonate is formed.

$Ca(OH)_2(aq) + C0_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

When excess carbon dioxide is bubbled into the lime water, the white precipitate dissolves due to the formation of soluble calcium hydrogen carbonate.

$CaCO_3(s) + H_20\ (l) + CO_2\ (g) \rightarrow Ca(HCO_3)_2(aq)$

Reaction of carbon dioxide with barium hydroxide

Carbon dioxide react with Barium hydroxide to forms a precipitate of barium carbonate.

$Ba(OH)(aq) + C0_2(g) \rightarrow BaCO_3(s) + H_20(l)$

Reaction of carbon dioxide with metals

Generally, carbon dioxide does not react with metals, except under special conditions.

When a burning magnesium ribbon is lowered into a gas jar containing carbon dioxide, it continues to burn for a short time with a **spluttering flame**. A white ash of magnesium oxide and black specks of carbon are formed.

$2Mg(s) + C0_2\ (g) \rightarrow 2MgO(s) + C(s)$

Reaction with water

Carbon dioxide reacts with water to form weak carbonic acid. The reaction is reversible.

$H_20(l) + C0_2(g) \rightarrow H_2C0_3(aq)$

Test for Carbon Dioxide

By using lime water

When a little carbon dioxide is bubbled into lime water a white precipitate of calcium carbonate is formed.

By using Barium hydroxide

When a little carbon dioxide is bubbled into Barium hydroxide a white precipitate of barium carbonate is formed.

Uses of Carbon Dioxide

The following are some of the uses of carbon dioxide.

1. Fire extinguisher

Carbon dioxide gas is used as a fire extinguishing agent because it does not support combustion and is denser than air. This means it displaces oxygen from the burning site, thus stopping the fire.

2. Manufacture of fizzy drinks

Carbon dioxide is used in making fizzy drinks and mineral salts. A solution of the gas in water has a pleasant taste.

3. Refrigeration

Carbon dioxide is used as a refrigerant. The gas sublimes at -78°C to form dry ice. Dry ice is a good refrigerant because it leaves no residue after sublimation

4. Manufacture of sodium carbonate

Carbon dioxide is used in the manufacture of anhydrous sodium carbonate in the **Solvay process.** Sodium carbonate is used in the manufacture of glass.

5. Manufacture of baking soda

Carbon dioxide is used in making baking powder (sodium hydrogen carbonate). Baking soda is prepared by passing carbon dioxide into cold concentrated sodium carbonate

6. Photosynthesis

Plants make use of carbon dioxide to manufacture food through the process of photosynthesis

7. Cloud seeding

Carbon dioxide is used in cloud seeding. **Cloud seeding** is a method of increasing the amount of rainfall received in an area by dispersing substances that help in cloud formation into the air