

## CARBON AND ITS COMPOUNDS

Carbon is in period 2 and group 4 of the periodic table. (It has an atomic number of 6)  
Its electronic structure is 2:4

### **Occurrence:**

Carbon occurs as an element in both natural and synthetic forms

| <b>Natural Forms</b> | <b>Synthetic Forms</b> |
|----------------------|------------------------|
| Diamond              | Coke                   |
| Graphite             | Charcoal               |
| Coal                 | Carbon fibre           |

### Allotropy of carbon

*Allotropy is the existence of an element in two or more forms without change of state.*

Allotropes are different forms of an element in the same physical state.

An allotrope is one of the different forms of an element in the same physical state.

Different allotropes of the same element have different crystalline structures and different physical properties but the same chemical properties.

The following elements exhibit allotropy

| <b>Element</b> | <b>Allotropes</b>                             |
|----------------|---|
| Carbon         | <i>Diamond, graphite and amorphous carbon</i> |
| Sulphur        | <i>Rhombic and monoclinic</i>                 |
| Tin            | <i>Grey and white tin</i>                     |
| Phosphorus     | <i>Red and white phosphorus</i>               |
| Oxygen         | <i>Diatomic and triatomic</i>                 |

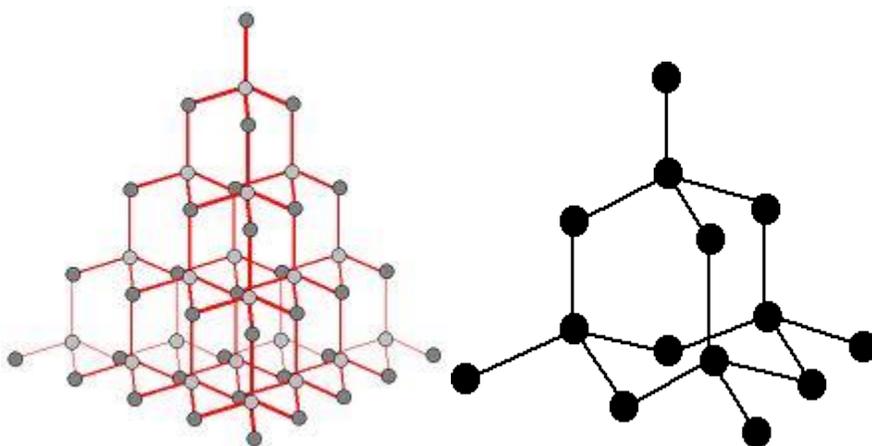
**Polymorphism** is the existence of a substance in more than one crystalline form. (Graphite and diamond)

Carbon has 3 allotropes i.e.

Graphite, diamond and amorphous carbon

### Diamond:

- It has a tetrahedral structure in which each carbon atom is joined to four other carbon atoms by covalent bonds.



- The valency electrons are all used up in forming covalent bonds leaving no free mobile electrons (*this explains why diamond does not conduct electricity*)
- This structure is extended throughout the substance in three dimensions.

#### *Physical properties of diamond*

- It is a colorless, transparent with a very high refractive index and sparkles.
- It is the hardest substance known
- Has a high melting point of  $3930^{\circ}\text{C}$
- Has a high density of  $3.5\text{gcm}^3$
- It is a non-conductor of electricity

#### *Uses of diamond*

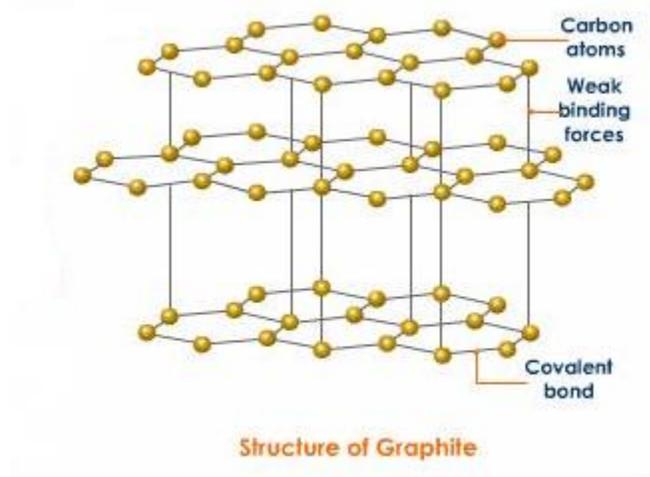
- Used in making jewelry due to its sparkling appearance
- Used for making glass cutters and drill tips for drilling rocks since it is a very hard substance.
- Used as diamond windows in X-ray machines since it is resistant to heat

#### **Graphite:**

- Graphite is a two-dimensional layered structure.
- The carbon atoms within the layers are arranged in hexagonal rings and each carbon is covalently bonded to three other carbon atoms.
- Therefore of the four valence electrons, 3 electrons are used in forming covalent bonds.
- The remaining electron is delocalized or free or mobile over the whole layer. (*This is why graphite conducts electricity and heat*).

- The layers lie on top of one another and are joined by weak van der waal's forces of attraction. The weak forces can easily break causing the layers to slide over one another (*this explains why graphite is slippery*)

It is manufactured artificially by heating coke to very high temperatures in the electric furnace.



**Question:** Describe the structure of:-

- Graphite
- Diamond – tetrahedral

**Physical properties of graphite**

- It is a black, opaque and slippery substance
- It is soft
- It is a good conductor of electricity
- Has a low density of  $2.3\text{gcm}^3$  compared to diamond
- Melts at about  $2300^\circ\text{C}$

**Uses of graphite**

- Used to make pencil lids since it is black and soft
- It is used as a lubricant since it is soft and slippery
- It is used as an electrode since it is a good conductor of electricity
- It is used as a moderator in atomic reactors

**Physical properties of diamond and graphite**

| Property | Diamond | Graphite |
|----------|---------|----------|
|----------|---------|----------|



### **Observation**

A black solid turns reddish-brown.

A colorless gas that turns lime water milky

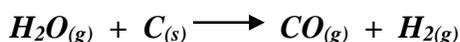
### **Example 2**



### **Observation:**

- Orange hot solid turns grey
- A colorless gas that turns lime water milky.

**Water gas**, which is a mixture of carbon monoxide and hydrogen, is industrially produced by the reaction between *steam and red hot coke*.



**Producer gas** consists of about  $\frac{1}{3}$  of carbon monoxide and nitrogen gas. It is manufactured by passing dry air over heated coke

### **Uses of Carbon:**

#### **1. Diamond:**

- It is used as drill tip and glass cutter because of its hardness.
- It is used for making jewelry because of its sparkling appearance.

#### **2. Graphite:**

- It is used as a lubricant.
- it is used to make lids of pencils.
- used as electrodes because it has mobile electrons.

### **CARBON DIOXIDE:**

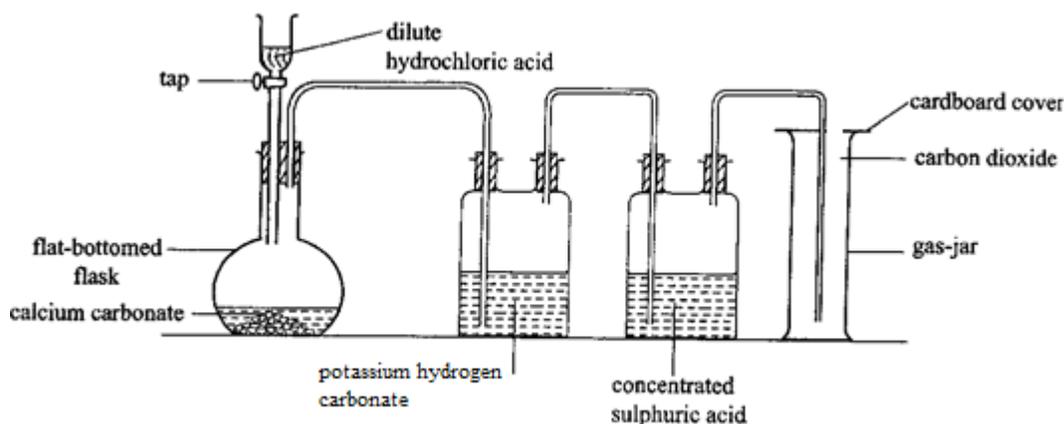
#### **Laboratory Preparation:**

- Calcium carbonate is put in a flask because it is a solid.
- The flask is fitted with a cork, a thistle funnel and a delivery tube
- Dilute hydrochloric acid is added to the flask from the funnel because it is a liquid

- An effervescence of a colorless gas is given off.



- The gaseous product is passed through potassium hydrogen carbonate solution to remove the acid sprays hydrogen chloride gas
- The gas is passed through concentrated sulphuric acid which dries it concentrated sulphuric acid being hygroscopic and does not react with the gas
- The gas is collected by downward delivery because it is denser than air



**NOTE:** 1. Sulphuric acid is not suitable for preparation of carbon dioxide since an insoluble layer of calcium sulphate is formed on the marble chips (calcium carbonate) preventing further reaction. Low yield of carbon dioxide are obtained

2. Carbon dioxide is also produced in a number of other reactions for example, combustion of hydrocarbons, action of heat on carbonates, fermentation and respiration.

### **Test for carbon dioxide**

If carbon dioxide is bubbled through lime water (calcium hydroxide) the lime water turns milky and then clear.

### **Physical Properties:**

1. Carbon dioxide is a colorless odorless gas.
2. It is denser than air.
3. Under pressure, carbon dioxide is quite soluble in water
4. Solid carbon dioxide is white and known as “dry ice”. Dry ice is used for cooling. When heated it does not melt but sublime forming the gas.



### Question:

When carbon dioxide was bubbled through a solution of calcium hydroxide (lime water) a white precipitate was formed which then dissolved to give a clear solution. Explain the observation.

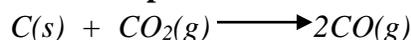
### Uses of Carbon dioxide:

1. It is used as a preservative in drinks e.g. sodas
2. It is used as a refrigerating agent for perishable goods.
3. It is used in fire extinguishers.
4. It is used in the solvay process to make sodium carbonate
5. It is used in bakeries

### CARBON MONOXIDE (CO).

Carbon monoxide is formed when carbon or carbon substances or compounds are heated in limited supply of air.

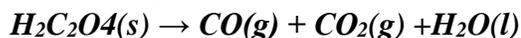
#### **For Example:**



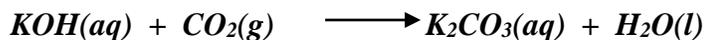
### Laboratory preparation of carbon monoxide:

Carbon dioxide is prepared by dehydrating oxalic acid using concentrated sulphuric acid.

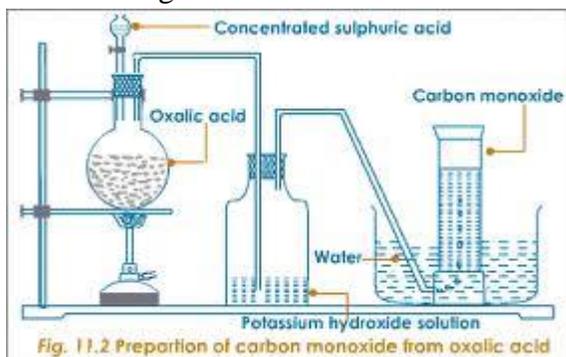
- Oxalic acid crystals are put in a flask because they are solid
- The flask is fitted with a cork, delivery tube and a thistle funnel
- Concentrated sulphuric acid is added to the flask from the funnel because it a liquid. The mixture is gently heated



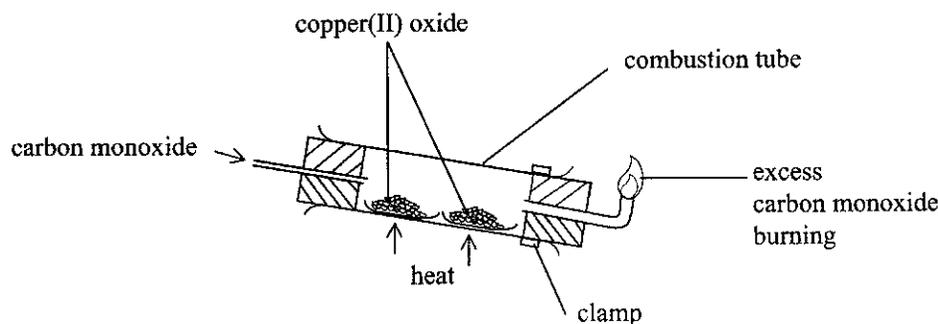
- The gaseous product is then passed through potassium hydroxide solution which absorbs carbon dioxide.



- The gas is collected over water because it slightly soluble in water



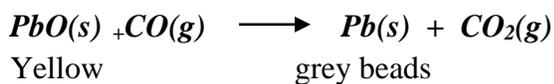




*Fig 1.12 Reducing action of carbon monoxide*

**Observation:**

- The black solid turns reddish brown.
- A colorless gas that turns lime water milky.



**Observation:**

- The orange hot solid turns to grey beads.
- A colorless gas turns lime water milky.

**CARBONATES**

These are salts derived from carbonic acid.

- They are generally insoluble in water except  $\text{NH}_4\text{CO}_3$ ,  $\text{Na}_2\text{CO}_3$  and  $\text{K}_2\text{CO}_3$
- The carbonates of K and Na do not decompose on heating because they are very stable
- Other carbonates on heating produce carbon dioxide gas and their metal oxide
- Ammonium carbonate produces carbon dioxide, ammonia gas and water on heating

**Laboratory preparation of sodium carbonate**

It is prepared in the laboratory by bubbling excess carbon dioxide gas in concentrated sodium hydroxide solution

- ✓ Excess carbon dioxide gas is bubbled into concentrated sodium hydroxide solution in a test tube placed in an ice-water mixture until a white precipitate is formed
- ✓ The precipitate if filtered, the solid residue collected on a filter paper washed with cold distilled water at least 2times and transferred to a hard glass test tube
- ✓ It is strongly heated to a constant mass
- ✓ The white powder obtained is sodium carbonate (soda ash)

**Explanation**

Carbondioxide reacts with concentrated sodium hydroxide solution to form sodium carbonate a soluble salt



When excess carbon dioxide is bubbled, sodium hydrogen carbonate a slightly insoluble salt is formed

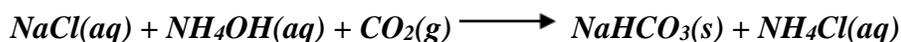


On heating the solid residue, sodium carbonate powder is formed



### Preparation of sodium carbonate by the Solvay process (large scale)

- Concentrated sodium chloride solution (*brine*) is saturated with ammonia gas in a tower to form ammoniacal brine
- The ammoniacal brine is run down the tower to react with the excess carbon dioxide  
*Note. The towers are fitted with perforated (pierced) mushroom shaped baffles at intervals to delay the liquid for the reaction to take place*
- Sodium hydrogen carbonate is formed and not being very soluble in water precipitates out



- The sodium hydrogen carbonate is filtered from the white sludge, washed and then heated to convert it to sodium carbonate



#### **Note**

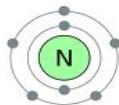
- The carbon dioxide produced by heating the hydrogen carbonate is recycled back into the tower*
- The ammonium chloride formed is reacted with calcium hydroxide to produce ammonia gas which is also recycled*
- The residual product in this process is therefore **calcium chloride solution***

## NITROGEN AND ITS COMPOUNDS

Nitrogen is group V period 2 element of the periodic table. It has an atomic structure shown.

7: Nitrogen

2,5



2:5

### Existence of nitrogen

Nitrogen exists as a diatomic molecule in the atmosphere constituting 78% of air. *This high percentage enables nitrogen gas to dilute the reactive oxygen gas in the atmosphere slowing down chemical processes like rusting, combustion and respiration, nitrogen being an unreactive gas*

Nitrogen also exists in compounds such as nitrates and other ammonium compounds

### Formation of a nitrogen molecule

Two nitrogen atoms covalently bond by sharing 3 pairs of electrons, each contributing 3 electrons to the bond pairs

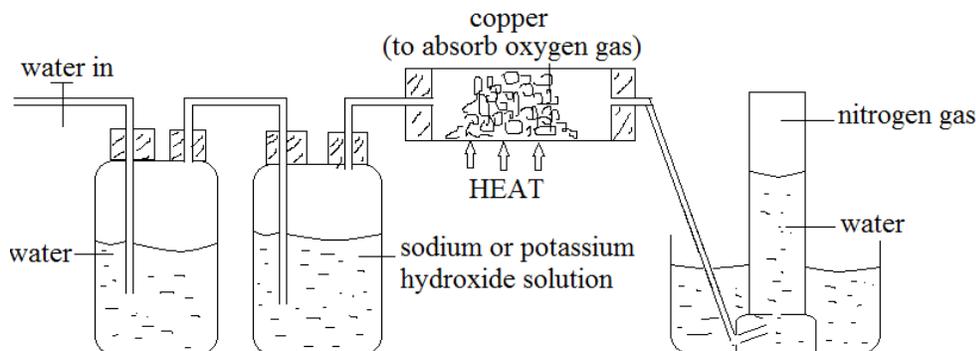
This is why a nitrogen molecule has a triple bond in its structure

Note; A *covalent bond* ( — ) is a bond pair of electrons (*two electrons shared between two atoms*)

### *Illustration*

## PREPARATION OF NITROGEN.

### (a) Isolation from air Apparatus



Note; Avoid sharp edges on your delivery tubes  
Endeavor to draw your diagram neatly and without any leakages

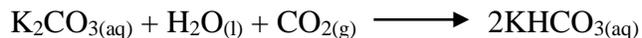
### Procedure.

- Air is driven out of the aspirator by passing water from a tap into the aspirator
- The air is then passed through a wash bottle containing concentrated potassium hydroxide solution to *remove carbon dioxide from air.*

### Equations:



Then

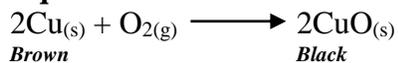


- The carbon dioxide-free air is then passed into a combustion tube with heated copper metal to *remove oxygen from the air.*

### Note:

In this reaction the brown copper metal is oxidized to black copper(II) oxide.

### Equation:



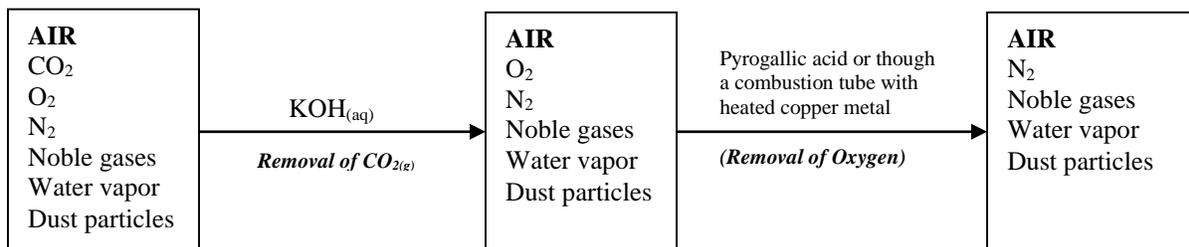
### Note:

- Alternatively oxygen can be removed by passing the carbon dioxide-free air through pyrogallous acid.
- *The remaining part of air is mainly nitrogen and is collected over water.*

### Note:

- Nitrogen obtained by this method contains noble gases like xenon, argon etc as impurities.

## Summary.



### (b) Removal from liquid air.

- Liquid air is primarily a mixture of nitrogen and oxygen with small amounts of noble gases.
- This method involves manufacture of liquid air and consequent fractional distillation.

### The chemical process.

#### Step 1: removal of dust particles.

- Dust particles are first removed by either of the two processes:

- Electrostatic precipitation
- Filtration.

#### i) Electrostatic precipitation:

- Air is passed through oppositely charged plates hence an electric field.
- Dust particles (charged) are consequently attracted to plates of opposite charges.

#### ii) Filtration:

- The air is passed through a series of filters which traps dust particles as the air is forced through.

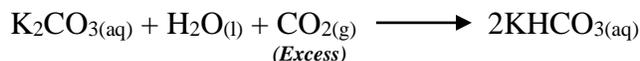
#### Step 2: removal of carbon dioxide.

- The dust-free air is passed through a solution of potassium hydroxide; to remove carbon dioxide.

#### Equations:



Then:



- Alternatively, sodium hydroxide may be used in place of potassium hydroxide.

#### Step 3: Removal of water vapour.

- The dustless, carbon dioxide-free air is next passed into a chamber with **concentrated sulphuric acid** or **anhydrous calcium chloride** in which water vapour is separated and removed.

#### Note:

To remove water vapour, air may be alternatively passed into a **freezing chamber** where it is condensed at  $-25^\circ\text{C}$ .

- The water vapour **solidifies** and is then absorbed by **silica gel** and separated out.

Note; Air is freed from carbon dioxide, water vapour and dust particles (before compression) so as to prevent **blockage** of the pipes caused by solid materials at liquefaction temperatures i.e. carbon dioxide and water vapour form solids which may block the collection pipes.

**Step 4: Liquefaction of air.**

- The air free from dust, carbon dioxide and water vapour is then compressed at about 200 atmospheres, cooled and allowed to expand through fine jet.
- This sudden expansion causes further cooling and the gases eventually liquefy.
- The liquid is tapped off through a valve while gas which has escaped liquefaction returns to the compressor.
- Liquid air is a transparent **pale blue liquid**.
- This liquid is then fractionally distilled.

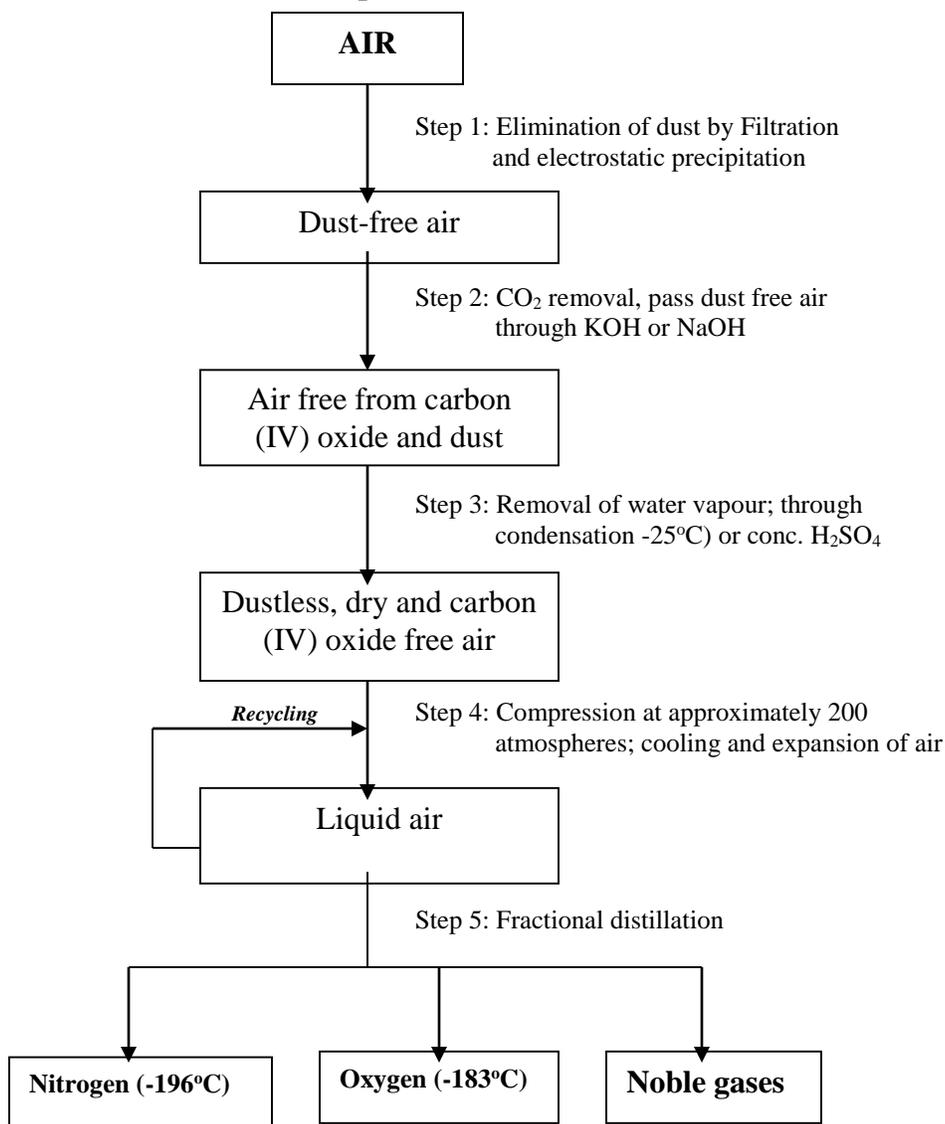
**Step 5: Fractional distillation of liquid air.**

- The boiling point of nitrogen is  $-196^{\circ}\text{C}$  (77K) and that of oxygen is  $-183^{\circ}\text{C}$  (90K).
- Consequently when liquid air is allowed to warm up, nitrogen boils off first and the remaining liquid becomes richer in oxygen.
- The top of the fractionating column is a few degrees cooler than the bottom.
- Oxygen, the liquid with the higher boiling point ( $-183^{\circ}\text{C}$ ) collects at the bottom as the liquid.
- The gas at the top of the column is nitrogen which has a lower boiling point.
- The more easily vapourised nitrogen is taken off.
- This way about 99.57% nitrogen is obtained.

**Note:**

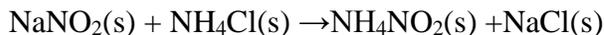
- The separation of nitrogen and oxygen from air is a proof that air is a mixture and not a compound.

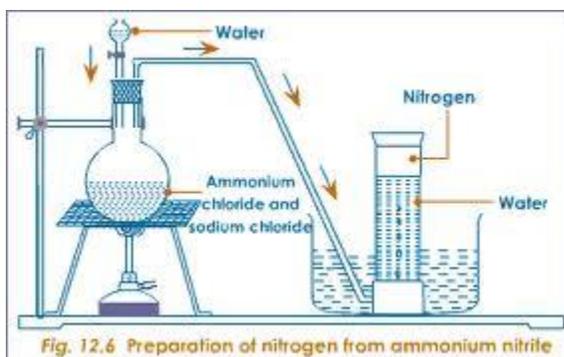
### Summary: Fractional distillation of liquid air.



### c) Laboratory preparation from ammonium nitrite

- ✓ *Nitrogen can also be prepared by heating ammonium nitrite.*
- ✓ However ammonium nitrite is not a stable compound and cannot be kept in the laboratory.
- ✓ A mixture of sodium nitrite and ammonium chloride is therefore heated to form ammonium nitrite which is then decomposed by heat liberating nitrogen gas





### Properties of nitrogen.

#### (a) Physical properties.

1. It is a colourless, odourless and tasteless gas; almost completely insoluble in water.
2. Slightly lighter than air.

#### (b) Chemical properties.

3. It is inert (unreactive)

#### Reason:

- The inert nature of nitrogen is due to the *strong triple covalent bonds* between the two nitrogen atoms in the molecule; N<sub>2</sub>.

- In air, it neither burns nor supports combustion and acts mainly as a diluent for the reactive oxygen; slowing down the rate of burning.

#### Chemical test for nitrogen.

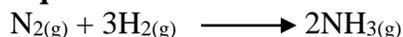
- A gas is identified to be nitrogen by elimination of other gases: -

- ✓ It extinguishes a lighted splint and does not burn; hence it is not oxygen, hydrogen or carbon monoxide.
- ✓ It has neither **smell** nor **colour**; and therefore is not chlorine, ammonia, sulphur dioxide or hydrogen chloride.
- ✓ It does not form a **white precipitate** in lime water, and so it is not carbon dioxide.
- ✓ It is **neutral** to litmus and therefore cannot be carbon dioxide, hydrogen sulphide, ammonia, hydrogen chloride

#### 4. Reaction with hydrogen.

- Under special conditions (i.e. high pressure, low temperatures and presence of iron catalyst), nitrogen combines with hydrogen to produce ammonia.

#### Equation:



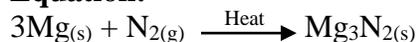
- This reaction forms the basis of **Haber process** used in the manufacture of ammonia.

#### 5. Reaction with burning magnesium.

- A piece of burning magnesium ribbon is introduced into a gas jar full of nitrogen.

**Observations:**

- The magnesium ribbon continues burning and a **white solid**; magnesium nitride is formed.

**Equation:**

## Explanation

*Burning magnesium produces a lot of heat that decomposes the strong triple bond in a nitrogen molecule*

**Note:**

- When magnesium nitride is treated with water the characteristic choking smell of ammonia is detected.

**Equations:****In water**

Note; other metals like **calcium** and **lithium** like magnesium react with nitrogen in the same way

**Uses of nitrogen**

1. Used in the **Haber process** in the manufacture of **ammonia**.
2. Due to its **inert nature**, it is mixed with argon to fill electric bulbs (to avoid soot formation).
3. In liquid state it is used as an inert **refrigerant** e.g. storage of semen for artificial insemination.
4. Due to its **inert nature**, it is used in food preservation particularly for canned products i.e. it prevents combination of oxygen and oil which tends to enhance food turning rancid (*become putrid*).
5. It is used in oil field operation called enhanced oil recovery where it helps to force oil from subterranean deposits.
6. It's used in filling petrol and oil tanks to prevent explosion since nitrogen is inert.
7. Nitrogen is also used in filling sodium storage tanks during the extraction of sodium metal.

## OXIDES OF NITROGEN

- The three main oxides of nitrogen are:

- Nitrogen (I) oxide, N<sub>2</sub>O
- Nitrogen (II) oxide, NO
- Nitrogen (IV) oxide, NO<sub>2</sub>

### Nitrogen (I) oxide (Dinitrogen oxide)

#### Preparation of nitrogen (I) oxide, N<sub>2</sub>O

##### Procedure:

- Ammonium nitrate is gently heated in a boiling tube and gas produced collected *over warm water*.
- Heating is stopped while excess ammonium nitrate still remains *to avoid chances of an explosion*.

##### Observations:

- The solid (ammonium nitrate) **melts** and gives off nitrogen (I) oxide which is collected over **warm water** because *Nitrogen (I) oxide is slightly soluble in cold water*.

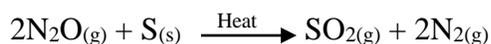
##### Equation:



#### Properties

1. It is a colourless gas, denser than air, fairly soluble in cold water and neutral to litmus.
2. It supports combustion by **oxidizing** elements like sulphur, magnesium and phosphorus under strong heat.

##### Equations:



3. Magnesium **decomposes** the gas and continues to burn in it.

##### Equation:



4. When exposed over red-hot finely divided copper it is reduced to **nitrogen**.

##### Equation:



5. Chemical test.

- It **relights** a *brightly glowing splint*.

*Dinitrogen oxide is decomposed by the heat from a brightly glowing splint releasing oxygen gas that relights the splint*

**Note:**

- It can be distinguished from oxygen by the following tests:
- It has a **sweet sickly smell**; oxygen is odourless.
- It will not give **brown fumes** (NO<sub>2</sub>) with nitrogen (II) oxide; oxygen does.
- It is **fairly soluble** in cold water; oxygen is insoluble.
- It extinguishes feebly burning **sulphur**; oxygen does not.

**Uses of nitrogen (I) oxide.**

- It was formerly used in hospitals as an **anesthetic** for **dental surgery** but has since been discontinued due to availability of more efficient anesthetics.

**Note:**

- Nitrogen (I) oxide is also called **laughing gas**; because patients regaining consciousness from its effects may laugh hysterically.

**Nitrogen(II) oxide, NO (nitrogen monoxide)**

**Laboratory Preparation**

**Procedure:**

- Action of heat on 50% concentrated nitric acid on copper turnings.
- Not any heat is required.

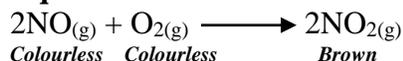
**Equation:**



**Observations:**

- An **effervescence** occurs in the flask; with **brown fumes** because the nitrogen(II) oxide produced reacts with oxygen of the air in the flask to form a brown gas, nitrogen(IV) oxide.

**Equation:**



- The brown fumes eventually disappear and the gas collected over water.
- The NO<sub>2</sub> fumes dissolve in the water in the trough, resulting into an acidic solution of nitrous acid.
- The residue in the flask is a **green** solution of **copper (II) nitrate**.
- Industrially, the gas is obtained when ammonia reacts with oxygen in the presence of platinum catalyst.
- This is the first stage in the production of nitric acid.

### Properties.

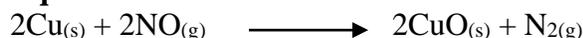
1. It is a colourless, **insoluble** and **neutral** to litmus. It is also slightly denser than air.
2. Readily combines with oxygen in air and forms brown fumes of nitrogen (IV) oxide.
3. Does not support combustion except in the case of **strongly burning magnesium** and **phosphorus**; which continues to burn in it, thereby reducing it i.e. it is an oxidizing agent.

### Example:



4. When passed over red-hot finely divided copper, it is reduced to nitrogen gas.

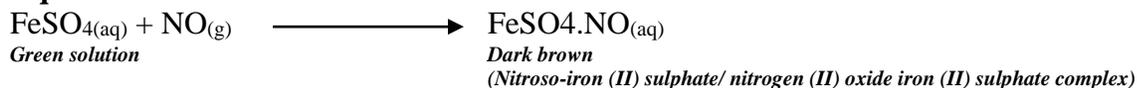
### Equation:



### 5. Reaction with iron(II) sulphate.

- When iron(II) sulphate solution (freshly prepared) is poured into a gas jar of nitrogen(II) oxide, a dark brown colouration of **Nitroso-iron(II) sulphate is obtained**.

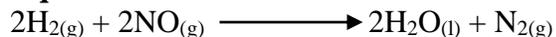
### Equation:



### 7. Reaction with hydrogen.

- When electrically sparked with hydrogen, NO is reduced to **nitrogen**.

### Equation:



### Chemical test:

- When exposed to air, nitrogen(II) oxide forms brown fumes of nitrogen(IV) oxide.

### Uses of Nitrogen(II) oxide.

**Note:** -It is not easy to handle owing to its ease of **oxidation**.

1. It is an intermediate material in the manufacture of nitric acid

## Nitrogen (IV) oxide (nitrogen dioxide)

### Laboratory Preparation:

#### Procedure:

- Action of concentrated nitric acid on copper metal.

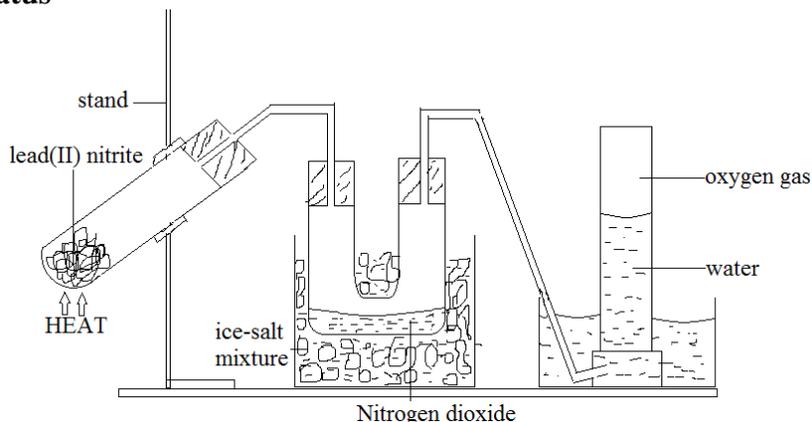
#### Equation:



#### Note:

- $\text{NO}_2$  is also prepared by the action of heat on **nitrates of heavy metals** like lead nitrate.
- $\text{NO}_2$  is given off together with **oxygen** when nitrates of heavy metals are heated.
- It is best prepared by heating **lead (II) nitrate** in a hard glass test tube.
  - ✓ Lead (II) nitrate is the most suitable because it crystallizes without **water of crystallization** (like other nitrates) which would interfere with preparation of nitrogen(IV) oxide that is **soluble** in water.
- The gas evolved passes into a U-tube immersed in an ice-salt mixture.

### Apparatus



Note; draw a neat diagram and avoid sharp edges on your delivery tube. Ensure there are no leakages on your diagram

#### Equation:



#### Observations:

- The heated **white** lead(II) nitrate crystals decompose and **decrepitate** (cracking sound) to yield **orange** lead(II) oxide; which turns **yellow** on cooling.
- A **colourless gas**, oxygen is liberated, followed immediately by **brown fumes** of nitrogen (IV) oxide.



### iii) Sulphur:



#### Note:

- NO<sub>2</sub> reacts with burning substances because the heat decomposes it to NO<sub>2</sub> and O<sub>2</sub>.

#### Equation:



- This is the oxidizing property of nitrogen (IV) oxide.
- The resultant oxygen supports the burning.

### 6. Effects of heat:

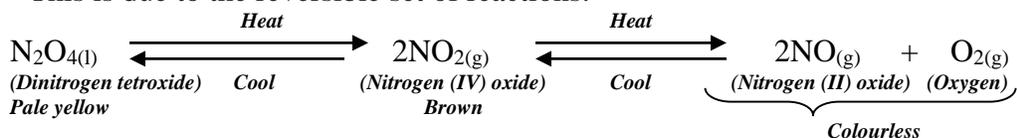
- On heating, nitrogen (IV) oxide dissociates to nitrogen (II) oxide and oxygen and will support a burning splint.

#### Equation:



7. - When liquid nitrogen(IV) oxide or dinitrogen tetroxide is warmed, it produces a **pale brown vapour**.

- This is due to the reversible set of reactions:



- Percentage of each in the equilibrium depends on temperature.
- At low temperatures, percentage of N<sub>2</sub>O<sub>4</sub> is high and the mixture is pale yellow in colour.
- Percentage of nitrogen(IV) oxide increases with increase in temperature and the colour darkens till at 150°C when the gas is entirely NO<sub>2</sub> and is almost black.
- Still at higher temperatures, nitrogen(IV) oxide dissociates into colourless gas (NO and O<sub>2</sub>).

### 8. Reaction with alkalis.

- A solution of aqueous sodium hydroxide is added to a gas jar of nitrogen(IV) oxide and shaken.

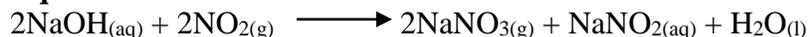
#### Observation:

- The brown fumes disappear.

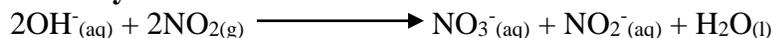
#### Explanation:

- nitrogen dioxide an *acidic gas* reacts with sodium hydroxide an alkali forming sodium nitrate and sodium nitrite.

#### Equation:



#### Ionicallly:



### Uses of nitrogen (IV) oxide.

1. Mainly used in the manufacture of nitric (V) acid.

### Summary on comparison between oxides of nitrogen.

|                                 | Nitrogen (I) oxide   | Nitrogen (II) oxide                              | Nitrogen (IV) oxide  |
|---------------------------------|--|--|--|
| Colour                          | - Colourless gas<br>- Sweet sickly smell                       | - Colourless; turns brown in air;<br>- Odourless | - Red brown gas;<br>- Choking pungent smell;                               |
| 2. Solubility                   | - Fairly soluble in cold water; but less soluble in hot water; | - Almost insoluble in water                      | - Readily soluble in water to form nitric (V) acid and nitrous (III) acid; |
| 3. Action on litmus             | - Neutral to litmus  | - Neutral to litmus                              | - Turns moist blue litmus paper red; i.e. acidic.                          |
| 4. Combustion                   | - Supports combustion; relights a glowing splint;              | - Does not support combustion;                   | - Does not support combustion.   |
| 5. Density                      | - Denser than air  | - Slightly denser than air                       | - Denser than air;   |
| 6. Raw materials and conditions | - Ammonium nitrate and heat;                                   | - Copper and 50% nitric acid;                    | - Copper metal and concentrated nitric acid;                               |

### Action of heat on nitrates.

- All nitrates except ammonium nitrate decompose on heating to produce oxygen gas as one of the products.

- Nitrates can be categorized into 4 categories based on the products formed when they are heated.

- The ease with which nitrates decompose increases down the electrochemical series of metals.

1. Nitrates of metals higher in the electrochemical series like sodium and potassium decompose on heating to give the corresponding metal nitrite and oxygen.

#### Examples:

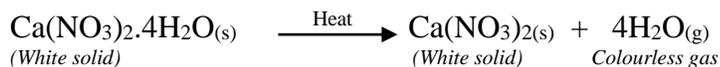


2. Nitrates of most other metals (heavy metals) that are average in the electrochemical series decompose on heating to give the metal oxide; nitrogen (IV) oxide and oxygen gas.

#### Note:

- Some nitrates are **hydrated** and when heated first give out their water of crystallization; and then proceed to decompose as usual on further heating.

#### Examples:



On further heating;





- ✓ The mixture is heated
- ✓ Ammonia gas is given off.
- ✓ It is then passed through lumps of calcium oxide which dries it, calcium oxide being hygroscopic and does not react with the gas
 

**Note;** ammonia gas being an alkaline gas cannot be dried with concentrated sulphuric acid because the two would react

The gas cannot also be dried with anhydrous calcium chloride because the two react forming a complex compound
- ✓ The gas is then collected by upward delivery because it is less dense than air

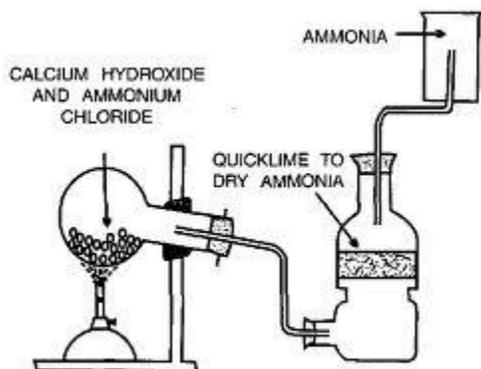
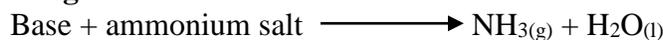


Fig. 9.1 Lab. preparation of dry ammonia

Ammonium chloride + calcium hydroxide



### Reagents.



### Note;

- the solid mixture of reagents is thoroughly ground in a mortar to *increase surface area for the reactions*

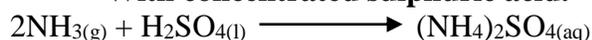
- A round-bottomed flask ensures **uniform distribution** of heat while heating the reagents.

- The flask should not be **thin-walled** because the pressure of ammonia gas liberated during heating may easily **crack or break** it.

- The flask is positioned **slanting downwards** so that as water condenses from the reaction, it does not run back to the hot parts of the flask and crack it.

Ammonia cannot be dried with the usual drying agents; concentrated sulphuric acid and calcium chloride as it reacts with them.

- **With concentrated sulphuric acid.**



- **With fused calcium chloride:**



- i.e. ammonia reacts forming **complex ammonium salt**.

### Other methods of preparing ammonia.

(b). Ammonia from caustic soda (sodium hydroxide) or caustic potash (potassium hydroxide)

Note:

- The slaked lime is replaced by either of the above solutions.
- Thus the solid reactant is ammonium chloride and the liquid reactant is potassium hydroxide.

#### Note;

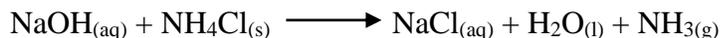
- The flask is not slanted. It is vertical and heated on a tripod stand and wire gauze.

Reason:

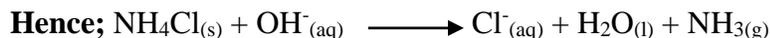
- No need of slanting since water produced is in liquid form and not gaseous. Thus there is no possibility of condensation of water on hotter parts.

### Equations

#### With caustic soda:



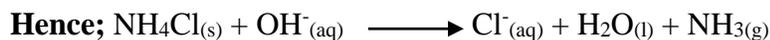
#### Ionic equation



#### With caustic potash:



#### Ionic equation



#### Note:

Ammonium sulphate could be used in place of ammonium chloride in either case.

### Equations:

#### With caustic soda:



#### Ionic equation;

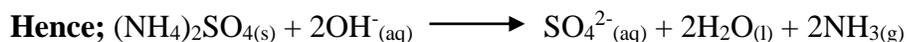




**With calcium hydroxide:**



**Ionicly;**



**Note:**

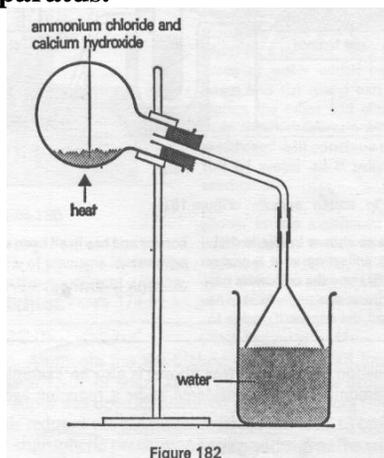
Reaction with calcium hydroxide however stops prematurely, almost as soon as the reaction starts.

**Reason;**

- Formation of insoluble calcium sulphate which coats the ammonium sulphate preventing further reaction.

### Preparation of ammonium solution.

(i). Apparatus.



(ii). Procedure:

- The apparatus is altered as above.
- The drying tower is removed and the gas produced is directly passed into water by an inverted funnel.

**Reasons for the inverted broad funnel.**

- It increases the **surface area** for the **dissolution** of thereby preventing water from “**sucking back**” into the hot flask and hence prevents chances of an **explosion**.

(iii). Equation.



**Note:**

- The solution cannot be prepared by leading the gas directly to water by the delivery tube.

**Reason:**

- Ammonia gas is **very soluble** in water and so water would rush up the delivery tube and into the hot flask causing it to crack.
- The rim of the inverted funnel is **just below** the water surface.

**Tests for ammonia.**

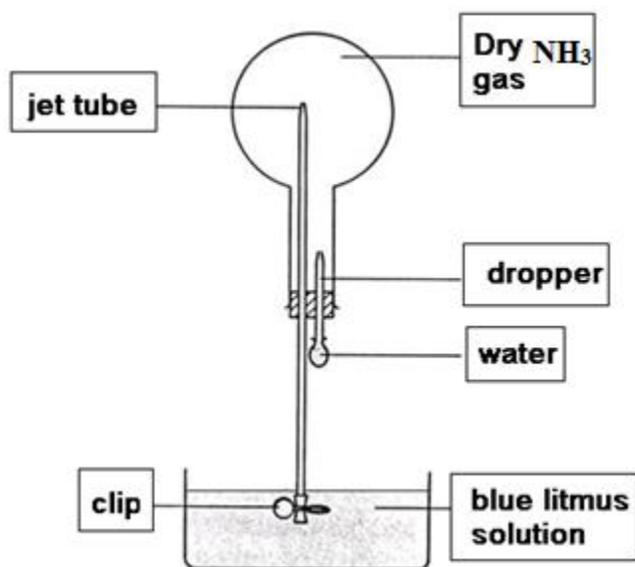
1. It is a colourless gas with a **pungent smell**.
2. It is the only common gas that is **alkaline** as it turns moist red litmus paper blue.
3. When ammonia is brought into contact with hydrogen chloride gas, **dense white fumes** of ammonium chloride are formed.

**Equation:**



**Fountain experiment.**

The apparatus is assembled as below (To show that ammonia gas is very soluble in H<sub>2</sub>O).

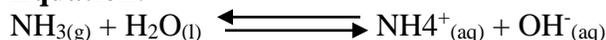


- A thick walled round bottomed flask is filled with ammonia gas and fitted with a jet tube with a clip and a dropper
- The flask is then inverted over water as shown in the diagram above.
- A few drops of water are allowed into the flask using a dropper and the few drops of water absorb the ammonia gas.
- Clip C is then opened and water rises up the tube by capillarity and comes out in form of a jet called a fountain.

### Properties and reactions of ammonia.

1. **Smell:** has a characteristic pungent smell.
2. **Solubility:** it is highly soluble in water. The dissolved ammonia molecule reacts partially with water to form ammonium ions ( $\text{NH}_4^+$ ) and hydroxyl ions ( $\text{OH}^-$ )

#### Equation:



- Formation of hydroxyl ions means that the aqueous solution of ammonia is (weakly) alkaline and turns universal indicator purple.

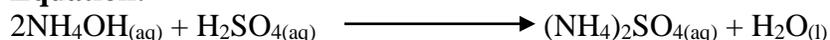
### 3. Reaction with acids.

- Sulphuric acid and concentrated ammonia solution are put in a dish and heated slowly.
- The mixture is evaporated to dryness.

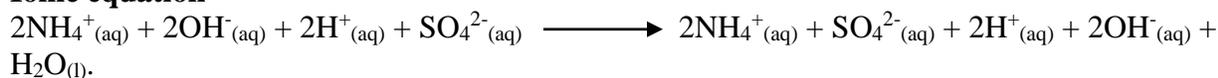
#### Observations:

- A white solid is formed.

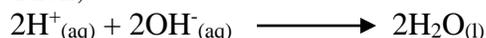
#### Equation:



#### Ionic equation



Then;



-

### 4. Reaction of ammonia with oxygen.

- Ammonia **extinguishes** a lighted taper because it does not support burning.
- It is **non-combustible**.
- However it burns in air enriched with oxygen with a **green-yellow flame**.

#### Experiment: Burning ammonia in oxygen.

#### Procedure:

- Dry oxygen is passed in the U-tube for a while to drive out air.
- Dry ammonia gas is then passed into the tube.
- A lighted splint is then passed into the tube.

#### Observations:

- A colourless gas is liberated.
- Droplets of a colourless liquid collect on cooler parts of the tube.

#### (iv). Explanations:

- The conditions for the reactions are:

- Dry ammonia and oxygen gas i.e. the gases must be **dry**.
  - All air must be driven out of the tube.
- Ammonia burns continuously in oxygen (air enriched with oxygen) forming nitrogen and water vapour i.e. ammonia is oxidized as hydrogen is removed from it leaving nitrogen.

**Equation:**



In the presence of a catalyst platinum, ammonia reacts with oxygen to form nitrogen monoxide and water

**Sample question:**

*Suggest the role of glass wool in the tube.*

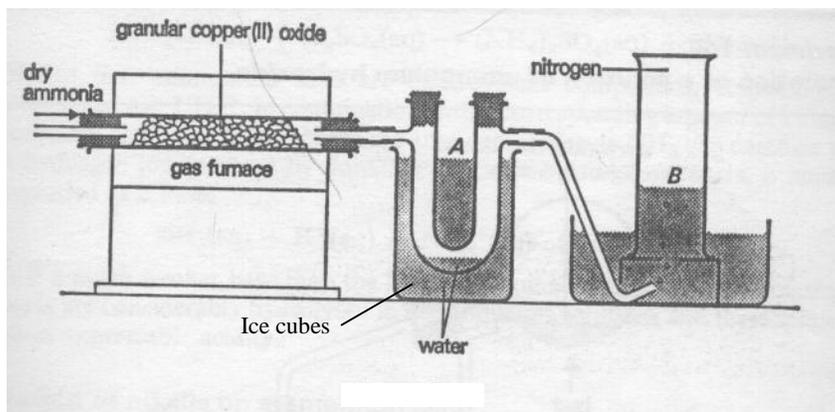
**Solution:**

- To slow down the escape of oxygen in the combustion tube, thus providing more time for combustion of ammonia.

**5. Ammonia as a reducing agent.**

- It reduces oxides of metals below iron in the reactivity series.

Experiment: reaction between ammonia and copper(II) oxide.



**Procedure:**

- Copper(II) oxide is heated strongly and dry ammonia is passed over it.
- The products are then passed through a U-tube immersed in cold water (ice cubes).

**Observations.**

- The copper (II) oxide **glows** as the reaction is **exothermic**.
- A **colourless liquid** collects in the U-tube.
- A **colourless gas** is collected over water.
- The **black** copper(II) oxide changes to **brown** copper metal.

**Explanations.**

- Ammonia gas **reduces** copper(II) oxide to copper and is itself **oxidized** to nitrogen and water.

**Equation:**

- The water produced condenses in the U-tube immersed in cold (ice) water.
- The resultant nitrogen is collected by downward displacement of water.
- The nitrogen gas collected is ascertained indirectly as follows:
  - A lighted splint is extinguished and the gas does not burn; thus it is not oxygen, hydrogen, or carbon (II) oxide.
  - It has neither smell nor colour; it is not ammonia, chlorine, sulphur (IV) oxide or nitrogen (IV) oxide.
  - It is not carbon dioxide because it does not turn lime water into a white precipitate.

**Note:**

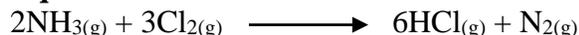
- This experiment proves that ammonia contains nitrogen.

**6. Reaction with chlorine.****Procedure:**

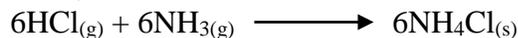
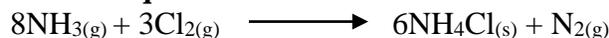
- Ammonia gas is passed into a bell jar containing chlorine.

**Observations:**

- The ammonia catches fire and burns for a while at the end of the tube.
- The flame then goes out and the jar then gets filled with dense white fumes of ammonium chloride.

**Equations:**

Then;

**Overall equation:****7. Ammonia solution as an alkali.**

- Solution of ammonia in water contains hydroxyl ions.

**Equation:**

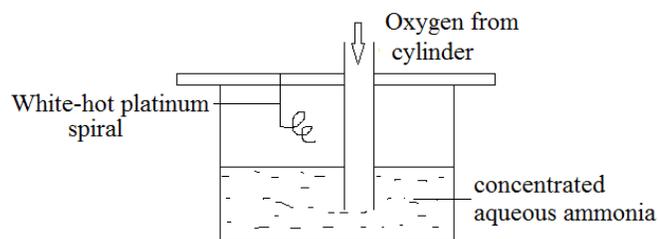
- Thus it has many properties of a **typical alkali**.
- Ammonia salts are similar to metallic salts.
- The group ( $\text{NH}_4^+$ ) precipitates in the reaction as a whole without splitting in any way.
- It exhibits unit valency in its compounds and therefore called a **basic radical**.

**Note:**

- It cannot exist freely as ammonia gas (NH<sub>3</sub>) which is a compound.
- Like other alkalis, ammonia solution precipitates insoluble metallic hydroxides by double decomposition when mixed with solution of salts of the metals.

## 8. Reaction with air in the presence of platinum wire.

### Apparatus:



### Procedure:

- Concentrated ammonia solution is put in a conical flask.
- The platinum (or even copper) wire is heated until white-hot.
- Oxygen gas or air is then passed through the ammonia solution.
- The red-hot platinum (copper) wire is then put into the flask containing the concentrated ammonia.

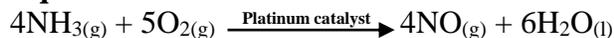
### Observations:

- The hot platinum wire **glows**.
- Red-brown fumes are evolved.

### Explanations:

- The hot platinum coil glows when it comes into contact with the ammonia fumes, which come from the concentrated ammonia solution.
- Reaction between ammonia and oxygen takes place on the **surface of the platinum** wire that acts as a **catalyst**.
- A lot of **heat** is produced in the reaction that enables the platinum coil to continue glowing.
- Ammonia is oxidized to **nitrogen(II) oxide**.

### Equation:



- *Red-brown fumes* of nitrogen(IV) oxide are produced due to further **oxidation** of the nitrogen(II) oxide to form nitrogen(IV) oxide.

### Equation:



### Uses of ammonia gas and its solution:

1. Ammonia gas is used in the manufacture of **nitric acid** and nylon.
2. Ammonia gas is important in the preparation of **ammonium salts** used as **fertilizers**.
3. It liquefies fairly easily (B.P is  $-33^{\circ}\text{C}$ ) and the liquid is used as a **refrigerant** in large cold storages and ice cream factories.
4. Liquid ammonia is injected directly into the soil as a high nitrogen content **fertilizer**.
5. Ammonia solution is used in **laundry work** as a **water softener** and a **cleansing agent** (stain remover)
6. Ammonia is used in the manufacture of sodium carbonate in the **Solvay process**.
7. Ammonia is used in "**smelling salts**". It has a slightly stimulating effect on the action of the heart and so may prevent fainting
8. Ammonia from ammonium carbonate acts on the heart and prevents fainting and dizziness.
9. Liquid ammonia is a source of hydrogen, it is easy to transport in cylinders and is readily decomposed to nitrogen and hydrogen by heated catalysts.
10. Ammonia is used in the manufacture of plastics like nylon

### INDUSTRIAL MANUFACTURE OF AMMONIA-The Haber process.

- Most of the world's supply of ammonia is from the synthesis of Nitrogen and hydrogen in the Haber process.

#### Raw materials

##### ✓ Nitrogen

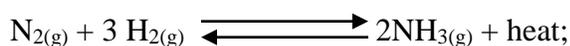
- Usually obtained from liquid air by fractional distillation

##### ✓ Hydrogen

- Obtained from water gas in the **Bosch process**.

- Also from crude oil (cracking)

#### General equation



#### Note:

- Nitrogen and hydrogen combine in the ratio 1:3 respectively to form two volumes of ammonia gas plus heat.

-The reaction is **exothermic** releasing heat to the surrounding.

#### Conditions

##### ✓ High pressures

- The process is favoured by high pressures and thus a pressure of approximately 200 to 300 atmospheres is used.

**Reason:**

- The volume of gaseous reactants from equation is higher than volume of gaseous products. Thus increased pressure leads to a decrease in volume, favoring the production of more ammonia.

**Note:**

Such high pressures are however uneconomical.

✓ **Low temperatures**

- Low temperatures favour production of ammonia;

**Reason:**

- The reaction is exothermic (releases heat to the surrounding) hence lower temperature will favour the forward reaction (shift the equilibrium to the right), producing more ammonia.

✓ **Catalyst**

- The low temperatures make the reaction slow and therefore a catalyst is used to increase the rate of reaction

- The catalyst used is *finely divided iron*; impregnated with Aluminium oxide ( $\text{Al}_2\text{O}_3$ )

**The chemical processes****Step 1: Purification**

-The raw materials, nitrogen and hydrogen are passed through a purification chamber in which impurities are removed.

-The main impurities are  $\text{CO}_2$ , water vapour, dust particles,  $\text{SO}_2$ ,  $\text{CO}_2$  and  $\text{O}_2$ ;

**Reason:**

The impurities would poison the catalyst

**Step 2: Compression**

- The purified Nitrogen and Hydrogen gases are compressed in a compressor at 500 atmospheres.

**Reasons:**

✓ To increase chances of molecules reacting;

✓ To increase rate of **collision** of the reacting particles.

✓ To increase pressure (attain desired pressures); and hence increase concentration of reacting particles.

**Step 3: Heat exchanger reactions**

- Upon compression the gaseous mixture, nitrogen and hydrogen are channeled into a heat exchanger; which **heats them** increasing their temperature.

- This enables the reactants (hydrogen and nitrogen) to attain the **optimum temperatures** for the succeeding reactions (in the catalytic chamber)

- From the heat exchanger the gases go to the catalyst chamber.

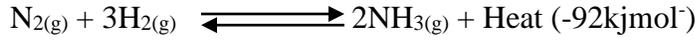
**Step 4: Catalytic chamber.**

- The gases then combine in the ratio of 1:3 ( $\text{N}_2:\text{H}_2$  respectively), to form ammonia.

- This reaction occurs in presence of a **catalyst**; which speeds up the rate of ammonia formation;

- The catalyst is **finely divided iron** impregnated with **aluminium oxide** ( $\text{Al}_2\text{O}_3$  increases the catalytic activity of iron).

**Equation in catalytic chamber**



- Only about 6-10% of the gases combine.
- Due to the high heat evolution involved, the products are again taken back to the heat exchanger; to cool the gases coming from the catalytic chamber.

**Step 5: Heat exchanger**

- The gases from the catalytic chamber enter the heat exchanger a second time.

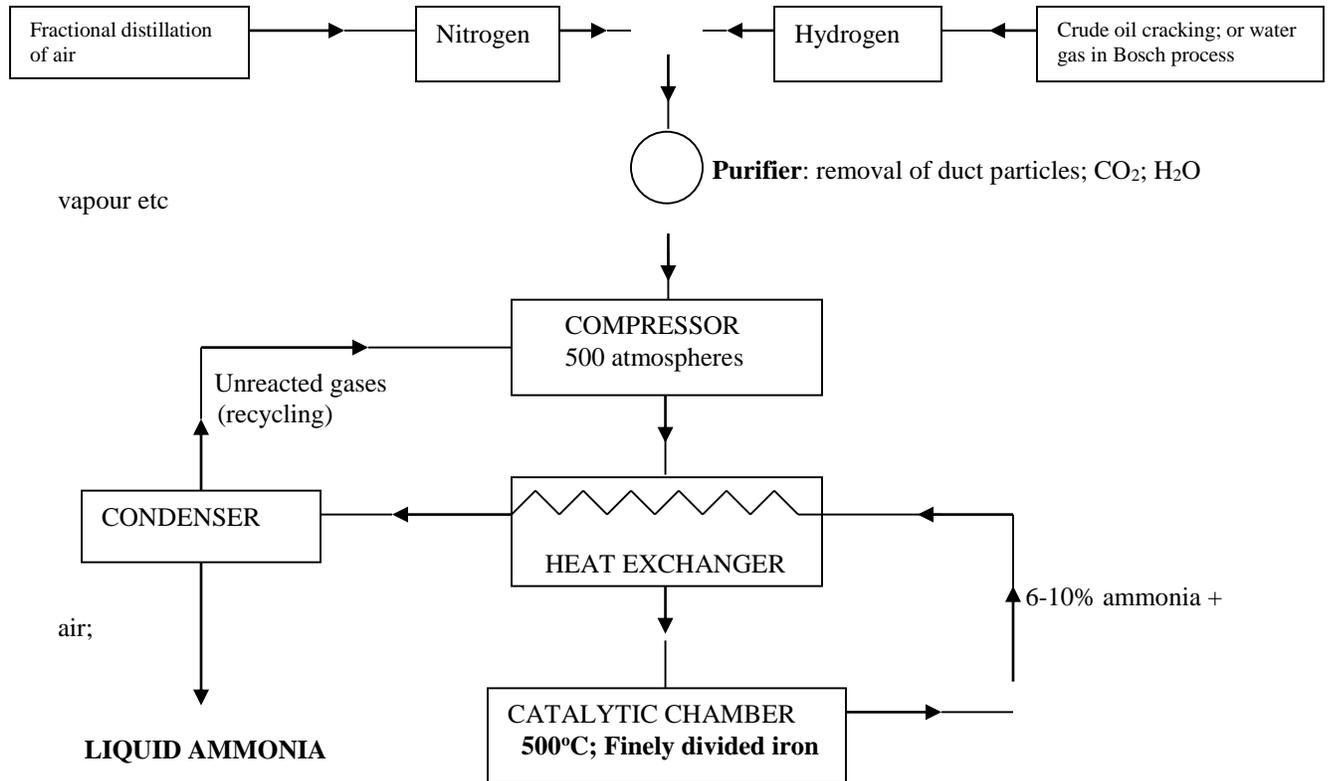
**Reason:**

- To **cool** the gases coming from the catalytic chamber, thus reduce cost of condensation.
- The gaseous mixture; ammonia and uncombined nitrogen and hydrogen are the passed through a condenser.

**Step 6: The condenser reactions (cooling chamber)**

- The pressure and the low temperatures in this chamber liquefy ammonia, which is then drawn off.
- The uncombined (unreacted) gases are recirculated back to the compressor, from where they repeat the entire process.

**Summary: flow chart of Haber process.**



### **Citing a Haber process plant**

- When choosing a site for this industrial plant, the following factors are considered:

1. Availability of raw materials (natural gas and crude oil)
2. Presence of cheap sources of energy.
3. Availability of transport and marketing.
4. Availability of appropriate technology and labour force.

### **Ammonium salts as fertilizers**

- Ammonium salts are prepared by the reaction between ammonia and the appropriate acid in dilute solution followed by evaporation and crystallization

#### **(a). Ammonium sulphate**

- Is prepared by absorbing ammonia in sulphuric acid.

#### **Equation:**

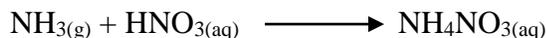


**Note:** It is a cheap fertilizer.

#### **(b). Ammonium nitrate**

- Is prepared by neutralization nitric acid by ammonia.

#### **Equation:**



- As there is some danger of exploding during storage, ammonium nitrate is mixed with finely powdered limestone ( $\text{CaCO}_3$ ).

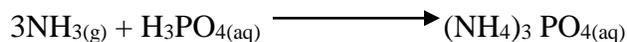
-The mixture, sold as nitro-chalk is much safer.

#### **(c). Ammonium phosphate**

- It is particularly useful as it supplies both nitrogen and phosphorus to the soil.

- It is prepared by neutralizing **orthophosphoric** acid by ammonia

#### **Equation:**

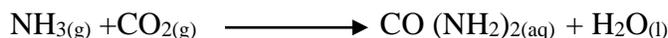


#### **(d) Urea $\text{CO}(\text{NH}_2)_2$**

- Is made from ammonia and carbon (IV) oxide

- Its nitrogen content by mass is very high; nearly 47%

#### **Equation:**

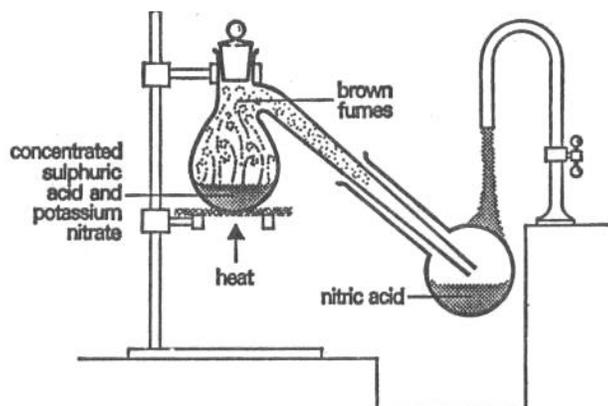


### **Nitric (V) acid**

- Is a monobasic acid (has only one replaceable Hydrogen atom); and has been known as strong water (aqua forty).

- It is a compound of hydrogen, oxygen and nitrogen.

## Laboratory preparation of nitric (V) acid



### Reagents

- Nitric acid is prepared in the laboratory by action of concentrated sulphuric acid on solid nitrates e.g. potassium nitrate ( $\text{KNO}_3$ ) and sodium nitrate ( $\text{NaNO}_3$ )

### Procedure

- 30-40 grams of small crystal of  $\text{KNO}_3$  are put in a retort flask.
- Concentrated sulphuric acid is added just enough to cover the nitrate; and then heated (warmed) gently.
- The apparatus is all glass.

### Reason:

- Nitric(V) acid would attack rubber connections.
- The neck of the retort flask is inserted into a flask that is kept cool continually under running water; this is where nitric acid is collected.

### Note:

The cold water running over the collection flask is meant to cool (condense) the hot fumes of nitric (V) acid.

### Observations and explanations

- Fumes of nitric are observed in the retort;

### Equation



-These fumes of nitric acid appear brown due to the *presence of nitrogen(iv) oxide gas formed by thermal decomposition of nitric acid.*

### Equation



- Pure nitric(V) acid is colourless but may appear yellow (brown) due to the presence of Nitrogen(IV) oxide.
- The brown colour can be removed by blowing air (oxygen) through the acid.
- Fuming nitric acid boils at 83°C and is 99% pure; while concentrated nitric acid is only 70% acid and 30% water.

**Note:** Nitric acid is usually stored in dark bottles.

**Reason:**

- To avoid its decomposition by light to nitrogen(IV) oxide, oxygen and water.
- The reaction in the retort flask is a typical displacement reaction; in which the more volatile nitric acid is displaced from nitrates by the less volatile sulphuric acid.
- The nitric acid distills over because it is more volatile than sulphuric acid.

**Properties of concentrated nitric acid**

- Nitric acid readily gives oxygen and therefore is called an oxidizer.
- The acid is usually reduced to nitrogen (IV) oxide and water.

**Effects of heat on concentrated nitric acid**

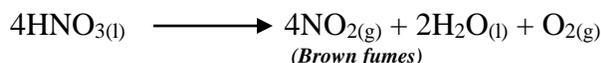
**Observations**

- Brown fumes are seen in the hard glass tube.
- Colourless gas is collected over water.

**Explanations**

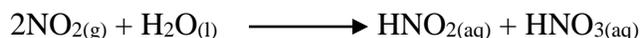
- Sand soaked in concentrated nitric acid produces nitric solid vapour on heating.
- The hot glass wool catalyzes the decomposition of nitric acid to nitrogen (IV) oxide (brown fumes), water vapour and oxygen.

**Equation**



- The so formed nitrogen(IV) oxide dissolves in water forming both nitric and nitrous acids.

**Equation:**



- The oxygen gas is collected over water; and with the solution becoming acidic.

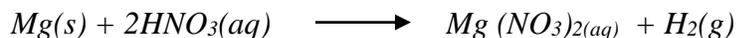
## CHEMICAL PROPERTIES OF NITRIC ACID

### a) As an acid:

Nitric acid behaves as an acid when dilute and a strong oxidizing agent when concentrated

#### As an acide;

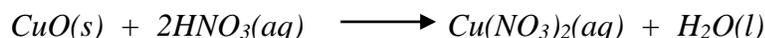
- i) Dilute Nitric acid does not release hydrogen when reacted with metals instead Nitrogen oxide or ammonia are formed. The exception is magnesium which reacts with cold dilute nitric acid to form magnesium nitrate and hydrogen.



#### Observation:

The solid dissolved with effervescence of a colourless gas that burns with a pop sound.

- ii) Nitric acid reacts with bases to form salts.



#### Observation:

Black solid dissolves to give a pale blue solution.

- iii) With carbonates:

Here the reaction produces carbon dioxide



#### Observation:

*The white salt dissolves with effervescence of a colourless gas that turns lime water milky to form a colourless solution.*



#### Observation:

*The green solid dissolves with effervescence to give a pale blue solution.*

### b) Properties of Nitric acid as an oxidizing agent:

- i) Reaction with saw dust

- Saw dust contains compounds of carbon Hydrogen and oxygen.

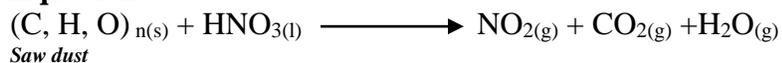
**Procedure:**

- Some saw dust is heated in an evaporating dish and some few drops of concentrated nitric (V) acid on it (this is done in a fume cupboard)

**Observation:**

- A violent reaction occurs, the saw dust catches fire easily and a lot of brown fumes of nitrogen (IV) oxide given off.  
- Nitric (V) acid oxidizes the compounds in saw dust to CO<sub>2</sub> and water; and itself it is reduced to nitrogen (IV) oxide and water.

**Equation:**



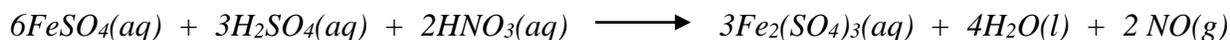
- Warm concentrated nitric (V) acid oxidizes pure carbon and many other compounds containing carbon.

**Equation:**

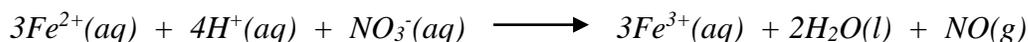


ii) With Iron(II) sulphate solution

Nitric acid oxidizes the green solution of iron (II) sulphate to a yellow solution of iron (III) sulphate



**Or**



**Observation:**

The pale green solution turns pale brown.

iii) With Carbon

Addition of concentrated nitric acid to carbon produces brown fumes of Nitrogen dioxide and carbon dioxide gas is also produced.



**Observation:**

*Brown fumes, and the black solid dissolves to give a colourless solution.*

iv) With sulphur

Concentrated Nitric acid oxidizes sulphur to sulphuric acid and it is itself reduced to Nitrogen dioxide gas.



**Observation:**

*The yellow solid dissolves with effervescence of reddish brown fumes to form a colourless solution.*

**v) With hydrogen sulphide**

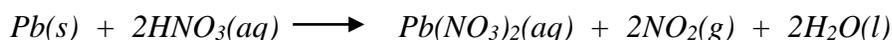
When hydrogen sulphide is bubbled through a fairly concentrated solution of Nitric acid, a pale yellow precipitate of sulphur is formed. Here the nitric acid oxidizes the hydrogen sulphide to sulphur and it is itself reduced to nitrogen dioxide.



**Observation:**

*A yellow precipitate*

**vi) Warm concentrated Nitric acid oxidises lead and zinc forming Nitrogen dioxide gas**

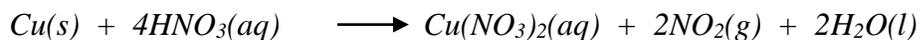


**Observation;**

The grey beads dissolved with effervescence of reddish brown fumes to form a colourless solution

**vii) With copper metal**

Concentrated Nitric acid oxidizes copper metal to copper (II) ions.



**Observation:**

*A reddish brown solid dissolved with effervescence of brown fumes and a pale blue solution is formed*

**Test for a nitrate ion**

**Brown ring test**

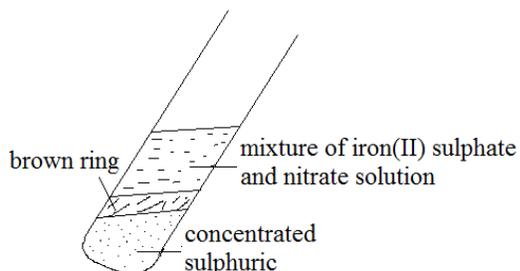
**Procedure.**

- An unknown solid is dissolved then acidified using dilute H<sub>2</sub>SO<sub>4</sub>.
- Some freshly prepared FeSO<sub>4</sub> solution is then added.

- The test tube is then held at an angle and concentrated sulphuric acid is added slowly (drop wise) to the mixture.

### Observations

- The oily liquid (conc.  $\text{H}_2\text{SO}_4$ ) is denser than water hence sinks to the bottom.
- A brown ring forms between the two liquid layers if the solid is a nitrate.



### Explanations:

- Suppose the solution tested is  $\text{KNO}_3$ , the conc.  $\text{H}_2\text{SO}_4$  and the  $\text{KNO}_3$  reacted to produce  $\text{HNO}_3$ .

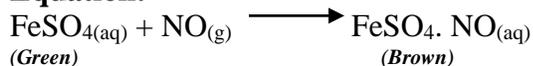
### Equation:



- The  $\text{NO}_3^-$  from nitric acid oxidizes some of the  $\text{FeSO}_4$  to  $\text{Fe}_2(\text{SO}_4)_3$  ( $\text{Fe}^{2+}$  to  $\text{Fe}^{3+}$ ) and itself reduced to  $\text{NO}$  by the  $\text{Fe}^{2+}$

-The  $\text{NO}$  so formed reacts with more  $\text{FeSO}_4$  to give a brown compound ( $\text{FeSO}_4 \cdot \text{NO}$ ) which appears as a brown ring.

### Equation:



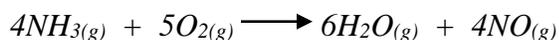
### Uses of nitric acid

- Large quantities are used in fertilizer manufacture.
- Manufacture of explosives (TNT)
- Manufacture of dyes
- Making nitrate salts
- Etching of metals.
- Manufacture of nylon and terylene
- Refining precious metals
- An oxidizing agent.

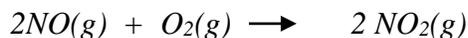
### INDUSTRIAL MANUFACTURE OF NITRIC ACID (The Ostwald's process)

*There are three stages in this process.*

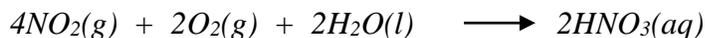
a) Ammonia is oxidized by oxygen in air over heated platinum catalyst at  $800^\circ\text{C} - 900^\circ\text{C}$ .



b) The gases are cooled and the colourless nitrogen monoxide combines with more oxygen to form brown fumes of Nitrogen dioxide.



c) Finally the nitrogen dioxide is dissolved in water with excess oxygen and the nitric acid produced is concentrated by evaporation.



- The resultant  $\text{HNO}_3$  is only 55%-65% concentrated.
- It is made more concentrated by careful distillation of the solution.

### The process of distillation (increasing the concentration).

- Concentrated sulphuric acid is added to the dilute nitric acid.
- The heat produced (when dilute sulphuric acid reacts with water) vapourises the nitric acid.
- The resultant nitric acid vapour is condensed.

#### Note:

- Nitric acid is stored in dark bottles.

#### Reason:

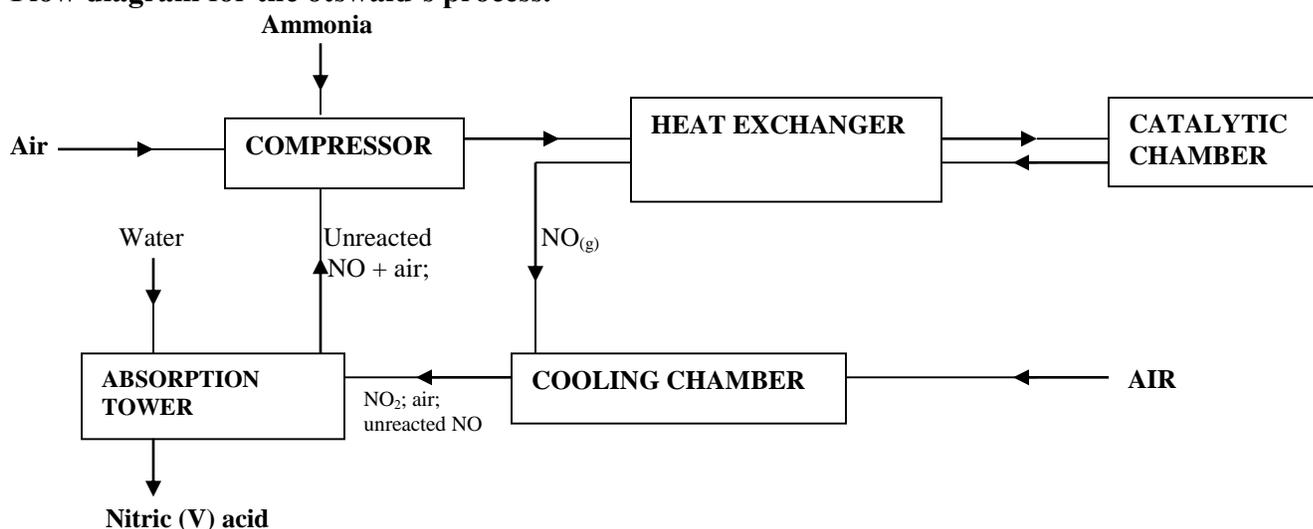
- To prevent its **decomposition** since it undergoes slow decomposition when exposed to light.

- Dilute nitric acid has higher ions concentration than concentrated nitric acid.

#### Reason.

- Dilute nitric acid is a stronger acid hence ionizes fully to yield more hydrogen ions than concentrated nitric acid.
- Dilute nitric acid is ionic whereas concentrated nitric acid is molecular;
- Dilute nitric acid is more (highly) ionized than concentrated nitric acid.

### Flow diagram for the Ostwald's process.



## **Pollution effects of nitrogen compounds.**

### **1. Acid rain**

- Nitrogen monoxide is produced in internal combustion engines on combination of nitrogen and oxygen.
- Nitrogen monoxide oxidized to nitrogen dioxide which dissolves in water to form nitrous and nitric acids.
- Nitric acid eventually reaches ground as acid rain and causes:
  - Loss of chlorophyll (chlorosis) from leaves
  - Corrosion of stone buildings and metallic structures, weakening them and destroying beauty.
  - Leaching of vital minerals from soils. These are converted into soluble nitrates and washed away from top soil. This leads to poor crop yields.

### **2. Smog formation.**

- Nitrogen dioxide also undergoes series of chemical reactions in air to produce one of the major components of **smog**.
- Smog reduces visibility for motorists, irritates eyes and causes breathing problems.

### **3. Eutrophication:**

- Refers to enrichment of water with excess nutrients for algal growth.
- Presence of nitrate ions from nitrogen fertilizers in a water mass encourages rapid growth of algae.
- This eventually leads to reduction of dissolved oxygen in water, killing aquatic animals like fish.
- Presence of **nitrate ions** in **drinking water** may also cause **ill health** to humans. This is because they are converted into carcinogenic compounds.

### **Prevention.**

1. Recycling unreacted gases in manufacture of nitric acid to prevent release into environment.
2. **Treating sewage** and **industrial effluents** to remove nitrogen compounds before releasing to rivers and lakes.
3. Fitting exhausts systems of vehicles with **catalytic converters** which convert nitrogen oxides into harmless nitrogen gas.
4. Adding lime to lakes and soils in surrounding regions to reduce acidity.
5. Applying fertilizers at right and in the correct proportion to prevent them from being washed into water masses.

***END***