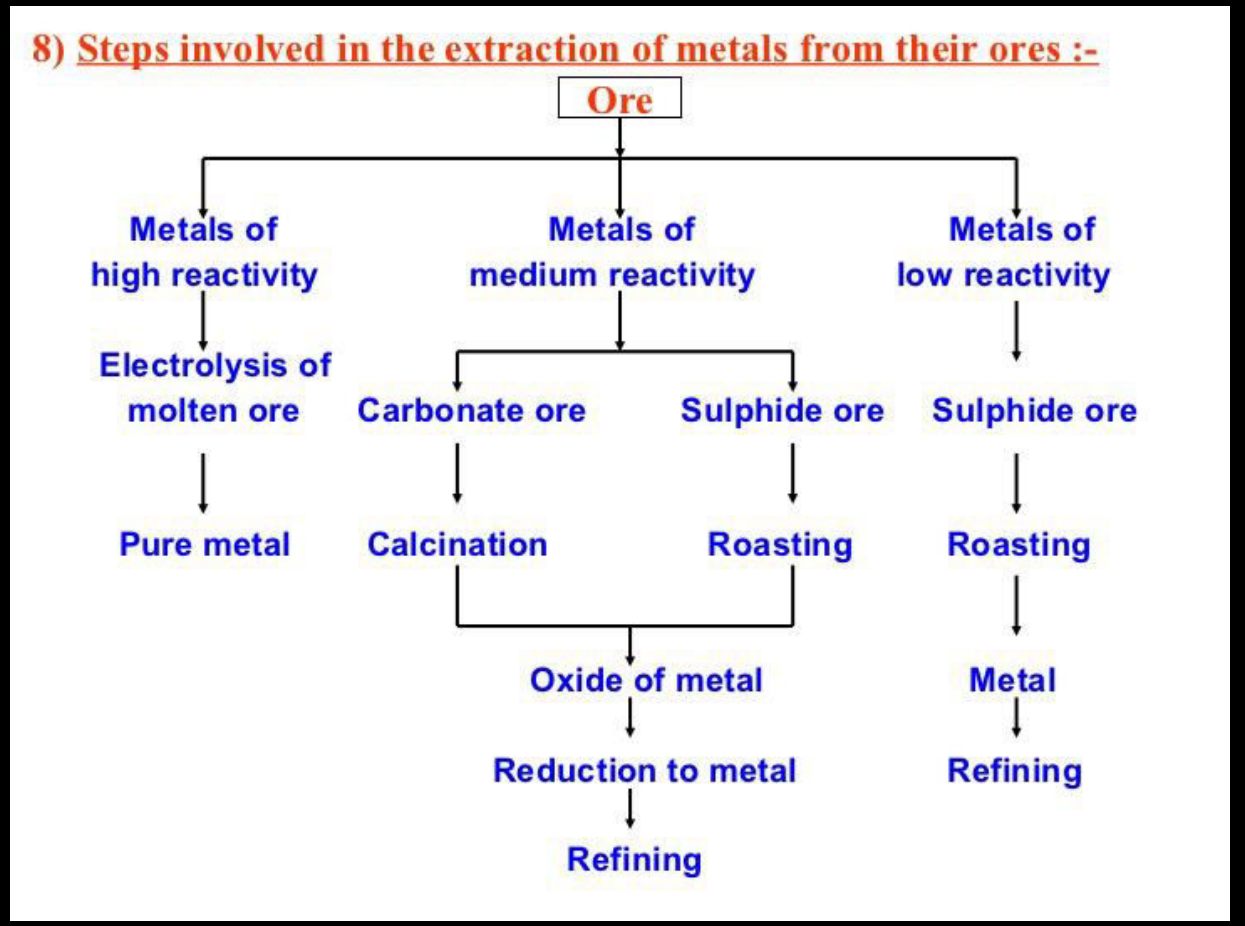
INDUSTRIAL INORGANIC CHEMISTRY

- Extraction of metals (copper and aluminium)
- The production of iron in blast furnace
- Rusting of iron and its prevention
- Ammonia and fertilisers
- Sufhur and sulfuric acid
- Chlor alkali industry
- Recycling metals

Extraction of metals/ Metallurgy



Extracting metals

Metal oxide + carbon ----- > metal + carbon dioxide

Some are so un-reactive that they exist in their NATIVE state e.g. gold, silver

If a rock is worth extracting we call the rock a METAL ORE

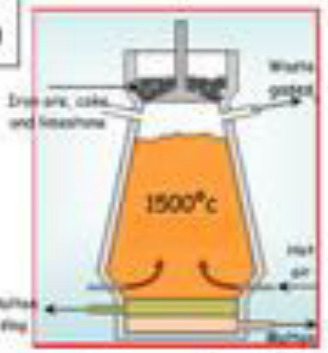
Metals that are less reactive than Carbon can be extracted from their oxide by reduction.

Metals more reactive than Carbon can be extracted by electrolysis

Haematite

Coke (carbon)

You need:



Blast Furnace

Limestone (removes impurities)

Reactivity Series

"Please stop calling my aunts zebra into the lesson cos she goes Potty "

Potassium
Sodium
Calcium
Magnesium
Aluminium
(Carbon)
Zinc
Iron
Tin
Lead
Copper
Silver
Gold
Platinum

High-alloy steels (12-15% metals) e.g. stainless steel which contains chromium-nickel

Iron + other elements = STEEL

A metal that contains other elements is an ALLOY

Alloy

Pure Iron

1st reaction:
$C + O_2 \dashrightarrow CO_2$

2nd reaction:
$CO_2 + C \dashrightarrow 2CO$

3rd reaction:
$Fe_2O_3 + 3CO \dashrightarrow 2Fe + 3CO_2$

Low-alloy steels (1%-5% metals)

Carbon steels (0.03% - 1.5% carbon)

Iron produced in the blast furnace
= PIG iron
It is made up of 96% iron. This means it is very brittle and not very useful.

Blast furnace

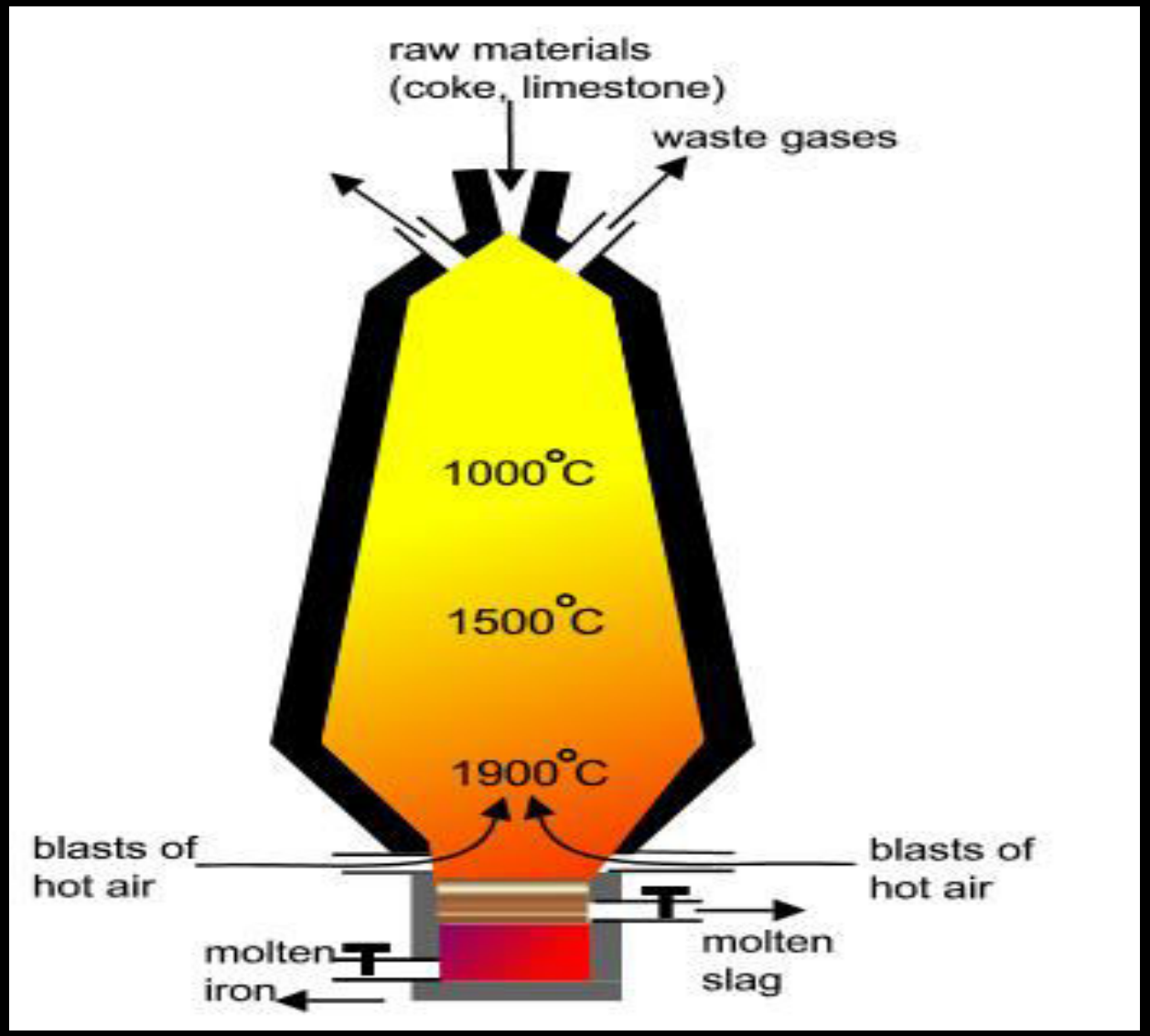
Extraction of iron

Step 1:
The Iron ore is crushed and broken into smaller pieces, concentrated with gravity separation process (washed with water).

Step 2:
The raw materials are added into the top of the furnace : Iron ore (Haematite, Fe_2O_3), Coke (Carbon, C), Limestone ($CaCO_3$)

Step 3:
It is heated by blowing hot air at the base of the blast furnace.

Step 4:
This causes the coke to burn to form Carbon Dioxide (CO_2). The process produces a very exothermic reaction, with the temperature rising as high as 1900 Degree Celscius.



Reactions in the Furnace:

- $C(s) + O2(g) \rightarrow CO2(g)$
- $CO2(g) + C(s) \rightarrow 2CO(g)$
- $Fe2O3 + 3CO \rightarrow 2Fe + 3CO2$
- $CaCO3(s) \rightarrow CaO(s) + CO2(g)$
- $CaO(s) + SiO2(s) \rightarrow CaSiO3(l)$

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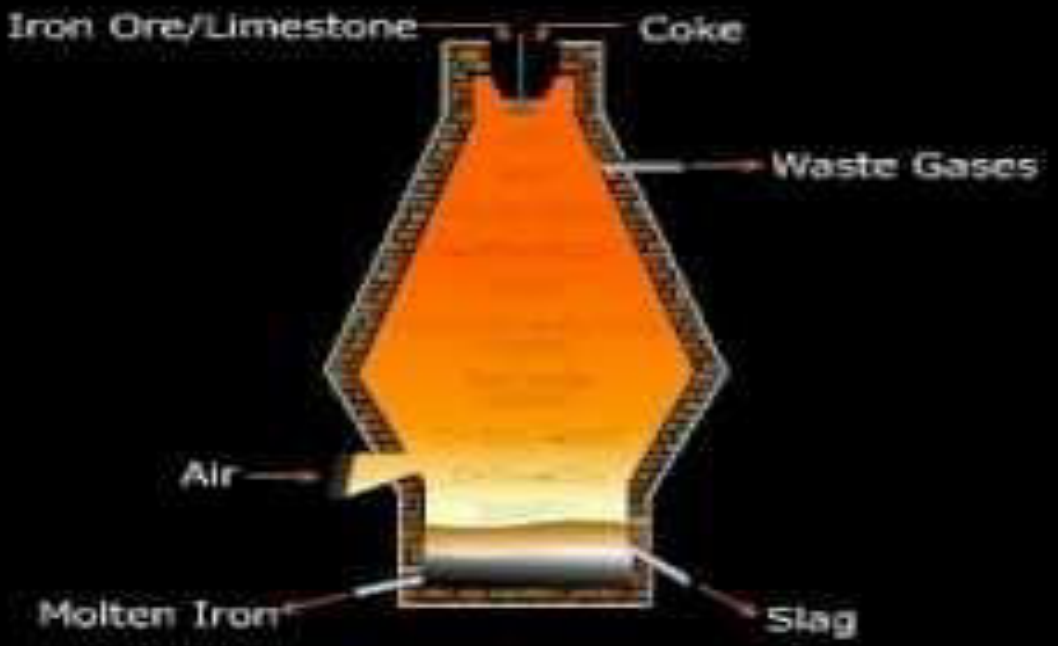
Blast Furnace

Close

Processes in the Blast Furnace

The charge is introduced into the furnace through the cup and cone arrangement



PAUSE REPLAY BACK NEXT

edurite

Rusting of iron

- The exposure of iron (or an alloy of iron) to oxygen in the presence of moisture leads to the formation of rust.
- The reaction of the rusting of iron involves an increase in the oxidation state of iron, accompanied by a loss of electrons. Rust is mostly made up of two different oxides of iron that vary in the oxidation state of the iron atom. These oxides are:

1.Iron(II) oxide or ferrous oxide. The oxidation state of iron in this compound is +2 and its chemical formula is FeO.

2.Iron(III) oxide or ferric oxide, where the iron atom exhibits an oxidation state of +3. The chemical formula of this compound is Fe_2O_3.

Factors that Affect the Rusting of Iron

- Many factors speed up the rusting of iron, such as the moisture content in the environment and the pH of the surrounding area. Some of these factors are listed below.
- Moisture: The corrosion of iron is limited to the availability of water in the environment. Exposure to rains is the most common reason for rusting.
- Acid: if the pH of the environment surrounding the metal is low, the rusting process is quickened. The rusting of iron speeds up when it is exposed to acid rains. Higher pH inhibits the corrosion of iron.
- Salt: Iron tends to rust faster in the sea, due to the presence of various salts. Saltwater contains many ions that speed up the rusting process via electrochemical reactions.
- Impurity: Pure iron tends to rust more slowly when compared to iron containing a mixture of metals.

Rust prevention

- Alloys that are Resistant to Rusting
- Galvanization
- Cathodic Protection
- Coatings

Ammonia and fertilisers

- Ammonia is an alkaline gas, which turns damp **red litmus paper blue**.
- physical properties - colorless gas
  - has a sharp or pungent smell
  - very soluble in water (because it reacts with water)
  - less dense than air

chemical properties - when ammonia reacts with water, ammonium ions and hydroxide ions are produced. the hydroxide ions produced make the solution of ammonia alkaline. the solution is only weakly alkaline because it is a reversible reaction - resulting in low concentration of hydroxide ions.

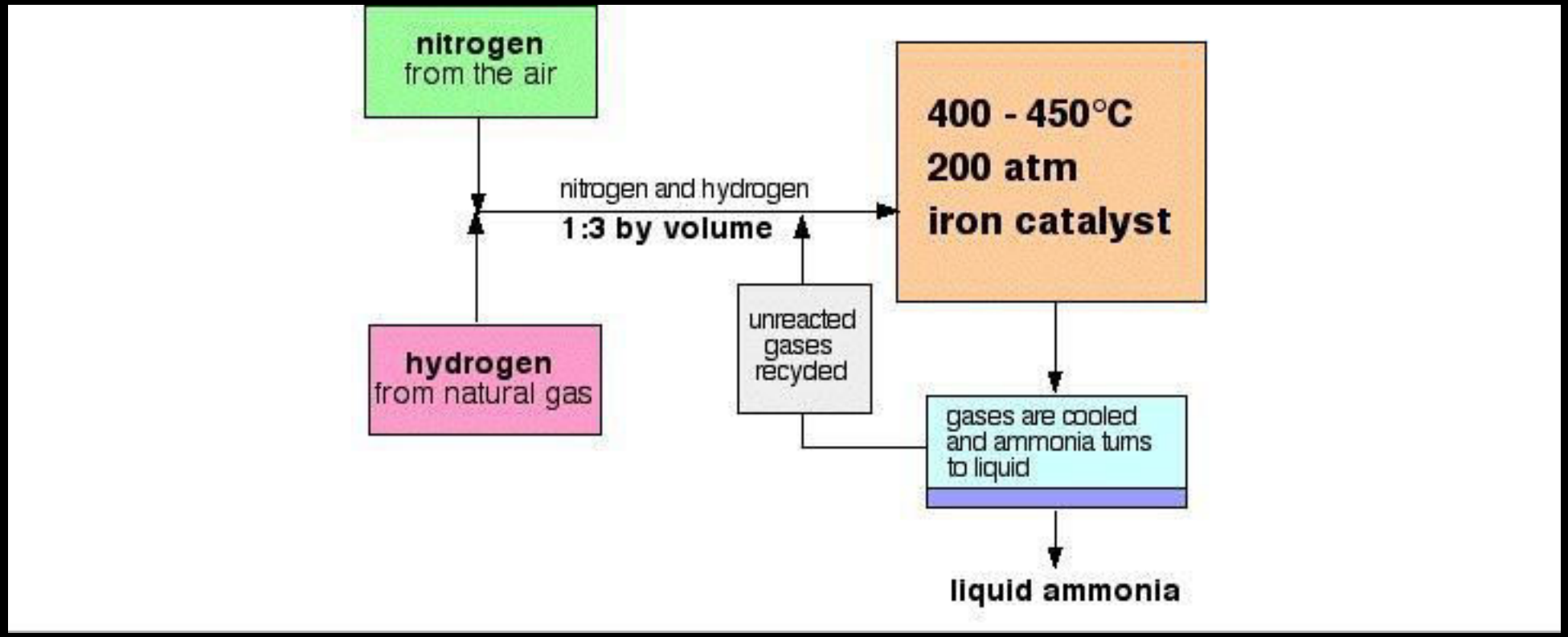
Ammonia and fertilisers

Manufacture of ammonia by the Haber Process

- The essential conditions:
  A temperature of about 450°C
  A pressure of about 200 atmospheres
  An iron catalyst
- **N_2 (g) + 3H_2 (g) ⇌ 2NH_3 (g)**
- This reaction is a reversible reaction.

Manufacture of ammonia by the Haber Process

The essential conditions:



The Haber process

nitrogen + hydrogen $\rightleftharpoons$ ammonia (iron catalyst)

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

A relatively cool temperature will shift the equilibrium to the right hand side.

Fertilisers

- Fertilisers contain inorganic ions which supply plants with elements they need, including **nitrogen, phosphorus and potassium**. Plants obtain these inorganic ions from compounds in the soil through their roots. These three ions are also called Primary Macronutrients.
- Nitrogen: Plants need nitrogen to make amino acids and protein for strong stems and healthy leaves (Most important element, plants take in nitrogen in the form of nitrate and ammonium ions)
- Phosphorus: Helps to develop the roots and ripen crops.
- Potassium: Helps plants to survive frost and resist diseases.

Fertilisers

- **Ammonium Salts**: Ammonium (NH_4^+) is an ion. Ammonium salts are formed by the neutralisation of acids by ammonia.

ex. Ammonia + Sulphuric Acid= Ammonium Sulphate

- Ammonium nitrate is one of the most common fertilizers. It is made by neutralizing nitric acid with ammonia solution

ex. $HNO_3 (aq) + NH_3 (aq) \rightarrow NH_4NO_3 (aq)$

nitric acid + ammonia --> ammonium nitrate

- Ammonium sulphate is made by similar neutralization:

ex. $H_2SO_4 (aq) + 2NH_3 (aq) \rightarrow$ ammonium sulphate

Fertilisers

- leaching of nitrates from the soil is a problem, while leaching of phosphates and potassium salts is not:
- Leaching of nitrates is when the soil gets excessively wet and can no longer hold any more water, so as water moves down the soil, it carries nitrogen with it. The nitrogen then gets into our water system.
- Effects of leaching of nitrates:
- blue baby syndrome - causes anoxia or internal suffocation; nitrates are convertedto nitrites and combine with the haemoglobin so your blood cannot carry oxygen around the body .
- eutrophication - fertilizers help algae to grow over the river like a carpet blocking out light and preventing photosynthesis from occuring. Fish die due to oxygen starvation.
- high concentration of nitrates in drinking water has been linked to cancers
- The reason why the leaching of phosphates and potassium salts are not a problem is because they are not as soluble in water as nitrates. Therefore, these salts do not get absorbed in the water in large quantities.

Fertilisers

- **NPK Fertilisers**
- Fertilisers contain **nitrogen, potassium** and **phosphorus**.
- Nitrogen promotes healthy **leaves**, potassium promotes **growth** and healthy **fruit** and **flowers** and phosphorus promotes healthy **roots**.
- Fertiliser compounds contain the following water-soluble ions:
  - Ammonium ions, NH_4^+ and nitrate ions, NO_3^-, are sources of soluble nitrogen.
  - Phosphate ions, PO_4^{3-} are a source of soluble phosphorus.
  - Most common potassium compounds dissolve in water to produce potassium ions, K^+.

Fertilisers

- **Displacement of ammonia**
- Ammonia can be **displaced** from its salts by the addition of an alkali substance.
- Farmers regularly add basic substances such as **calcium** hydroxide to their soil to neutralise any excess soil acidity.
- If **too** much of the basic substance is added or if it has been added **too** soon after fertiliser has been added, then an ammonia displacement reaction may occur.
- This involves the loss of nitrogen from the fertiliser, nullifying its effectiveness as a fertiliser.
- For example, the salt ammonium chloride is used extensively in fertilisers and reacts with calcium hydroxide:
- $2NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2NH_3 + 2H_2O$

Sulphur and sulphuric acid

- **Sources of sulfur**
- Sulfur is found in its elemental state underground in the USA, Mexico and Poland.
- It is also a by-product from the removal of sulfur from **petroleum** and **natural** gas.
- Sulfur can also be obtained from **sulfide** ores.
- **Uses of sulfur**
- The main use of sulfur is in making **sulphuric** acid which is a very important chemical used in many industries.
- It is also used extensively in making **rubber** tyres more flexible (vulcanising), where the rubber is heated with sulfur.
- **Sulfur dioxide**
- Sulfur dioxide can be made by the direct combination of sulphur with oxygen.
- This is the method used in the first stage of the manufacture of sulfuric acid:
- $S + O_2 \rightarrow SO_2$
- **Uses of sulfur dioxide**
- As a bleach in the manufacture of wood pulp for paper.
- As a preservative for foods and drinks by killing bacteria.
- Sulfites are often added to foods and these release sulfur dioxide in acidic conditions.

CONTACT PROCESS

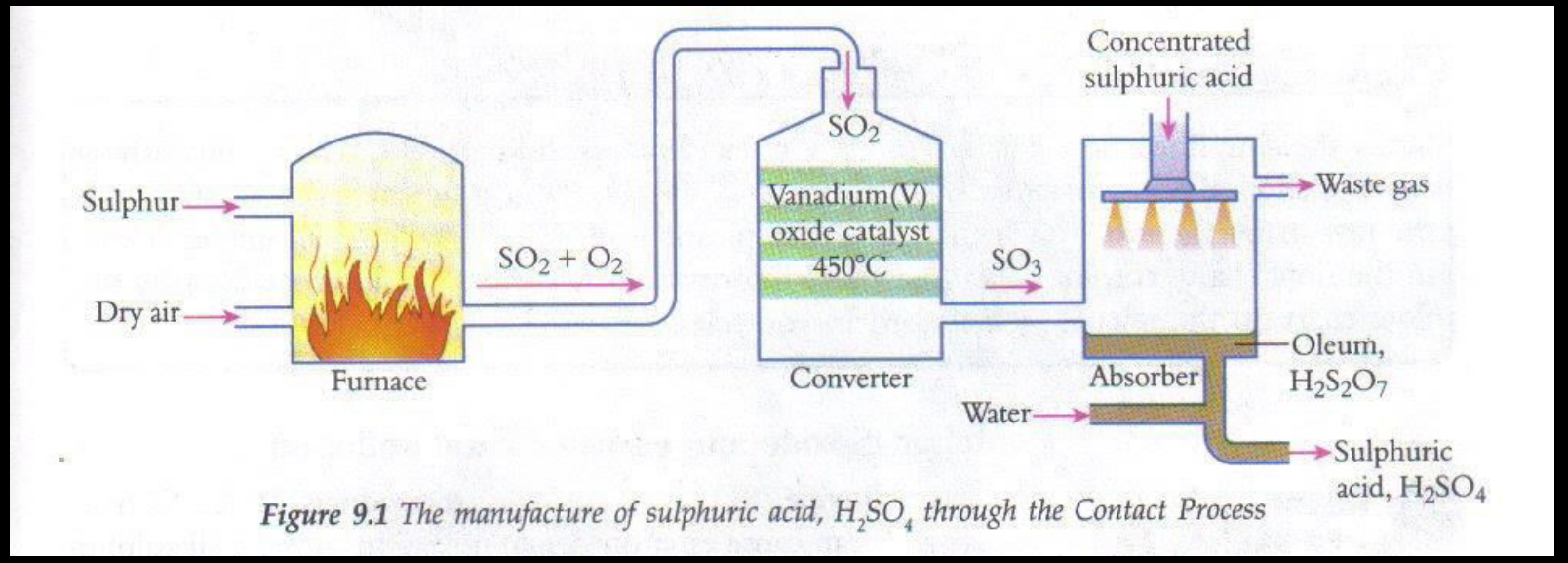


Figure 9.1 The manufacture of sulphuric acid, H_2SO_4 through the Contact Process

- **STEPS:**

1. Sulfur is burned in air or oxygen to produce sulphur dioxide $S(s) + O_2(g) \rightarrow SO_2(g)$
2. Sulfur dioxide reacts with oxygen to form sulfur trioxide

   The reaction is catalyzed by vanadium pentoxide at 450°C and 2 atmospheres.

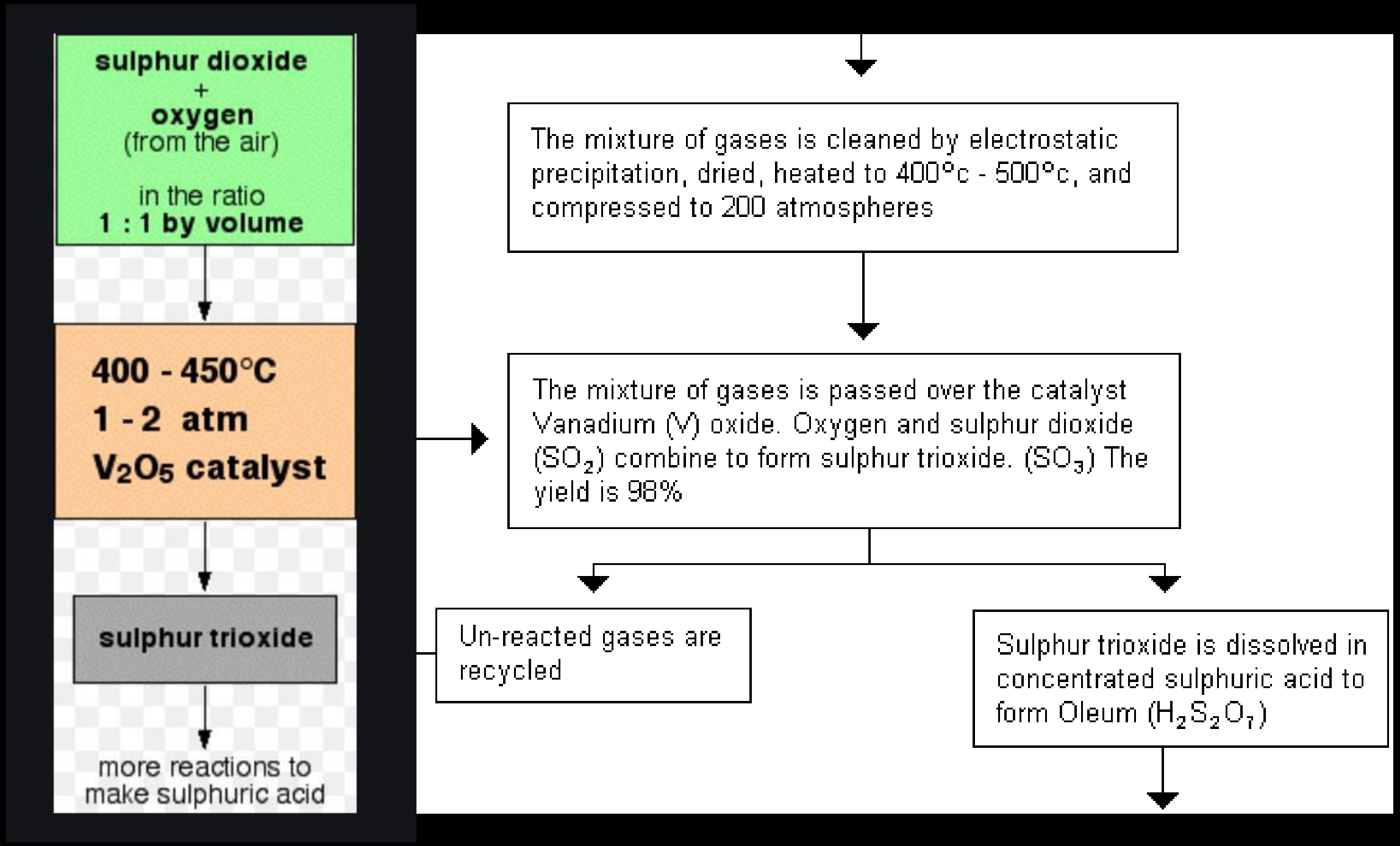
$$2\,SO_{2\,(g)} + O_{2\,(g)} \xrightleftharpoons{V_2O_5\ /\ 450^{\circ}C\ /\ 2\ atm} 2\,SO_{3\,(g)} \qquad \Delta H < 0$$

3. Sulfur trioxide is bubbled through a previously prepared sulphuric acid to form oleum (fuming sulphuric acid)

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

4. Oleum is added to water to produce sulphuric acid of the required concentration.

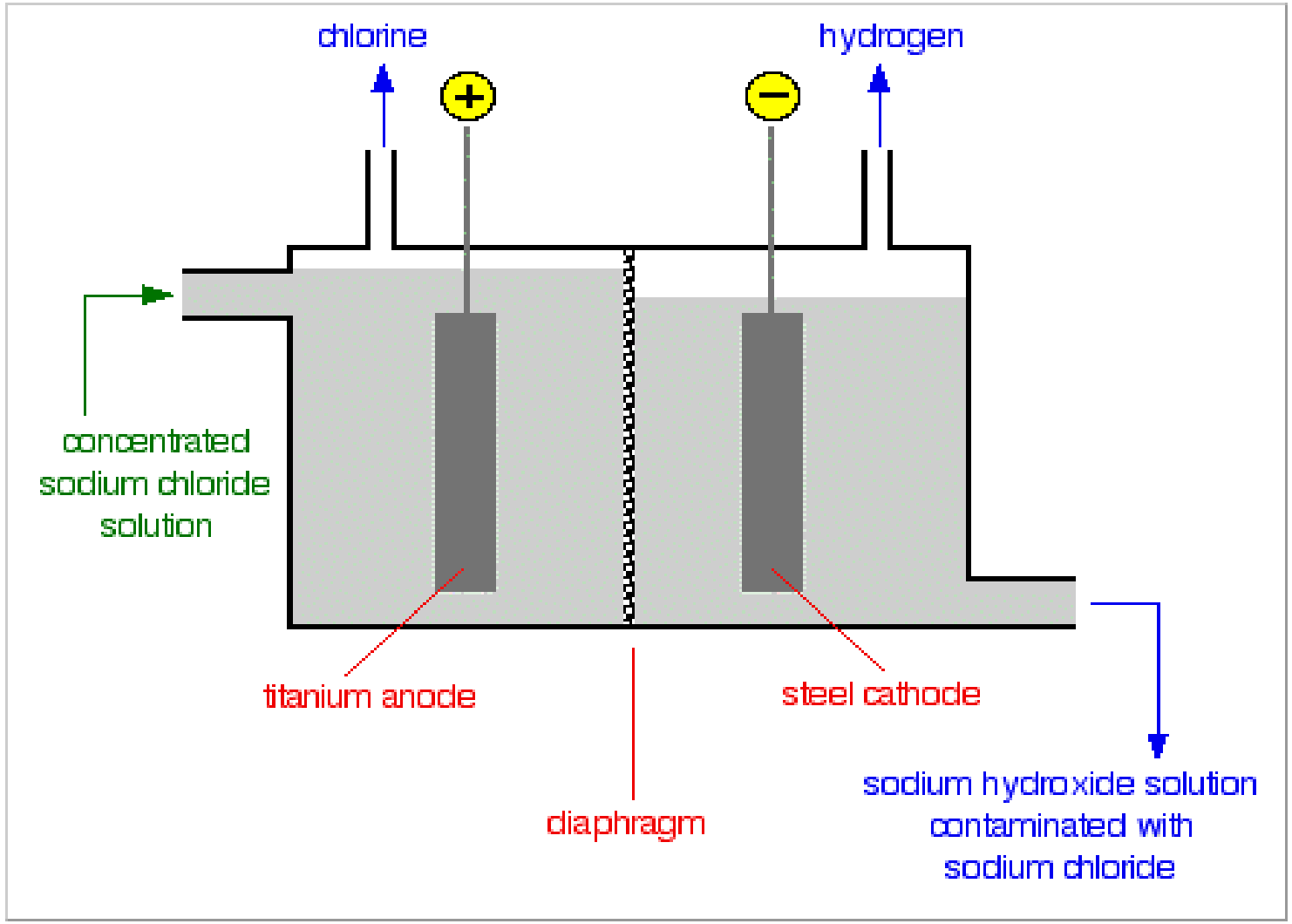
(Oleum is always added to water and not the other way round because the reaction is highly exothermic and the acid could spit out of the container causing severe burns). $H_2S_2O_7 + H_2O \rightarrow 2\,H_2SO_4$


sulphur dioxide
+
oxygen
(from the air)
in the ratio
1 : 1 by volume
400 - 450°C
1 - 2 atm
V_2O_5 catalyst
sulphur trioxide
more reactions to make sulphuric acid
The mixture of gases is cleaned by electrostatic precipitation, dried, heated to 400°c - 500°c, and compressed to 200 atmospheres
The mixture of gases is passed over the catalyst Vanadium (V) oxide. Oxygen and sulphur dioxide (SO_2) combine to form sulphur trioxide. (SO_3) The yield is 98%
Un-reacted gases are recycled
Sulphur trioxide is dissolved in concentrated sulphuric acid to form Oleum ($H_2S_2O_7$)

Chlor alkali industry

- **THE SOLVAY PROCESS**
-
- The Solvay process electrolyses Brine (salt water) to create Hydrogen, Chlorine and Sodium Hydroxide.
- Salt water is a cheap raw material.
- Hydrogen is used in making Margarine, as a fuel and in the Haber Process.
- Sodium Hydroxide is used to make soap, oven cleaner, bleach and products for unblocking drains.
- Chlorine is used to sterilise water and to make pesticides, Hydrochloric Acid and PVC.
-

Chlor alkali industry



Recycling metals

- Scrap **metal recycling** involves the recovery and processing of scrap **metal** from end-of-life products or structures, as well as from manufacturing scrap, so that it can be introduced as a raw material in the production of new goods. It can be recycled repeatedly with no degradation of its properties.
- Some of the most common manufactured items that use a high concentration of scrap metal include automobiles, aircraft, appliances, and industrial containers, ductwork, and plumbing. Recycled **aluminum** and steel are commonly reused as new food packaging.

Recycling metals

- **The advantages of recycling compared to producing metals from metal ore include:**
- more economic - less energy is needed to produce a metal.
- less damage to the environment - fewer quarries and mines, less noise and less heavy traffic.
- saves valuable raw materials - reserves of metal ores will last longer.