

Republic of Somaliland

Chemistry Text Book



Grade: Form three

Introduction

This book is designed for **form 3 Somaliland secondary schools** and contains all of the topics they require. Each topic is supported theories, facts, explanations, analysis and exercises and it's based on newly revised curriculum.

The objective of this book is to satisfy educational needs of the learners and to help them to develop self-reliance and complete confidence in their abilities to understand and solve chemistry problems.

In the context of this objective, the book contains:

- Detailed notes, many examples, clear mathematical formulas, calculations, short clarifications,
- Summaries and exercises which address all topics covered in the text
- Chapter review questions at the end of each chapter allow students to practice

This book is an excellent teaching learning tool for both teachers and students.

The language is kept simple, to improve accessibility for all students, Care is taken to introduce and use all the special terms that students need to gain a complete understanding of the chemical concepts introduced.

In the text, key terms are highlighted in bold. The depth and breadth of each topic is pitched at the appropriate O level students.

The key objective of this book is to improve the quality of secondary chemistry education; all of it has been reviewed and revised, ensuring that the new specification is fully covered.

In addition to the main content in each chapter describing issues, applications or events, which put the chemical content introduced into a social context.

Chapter Exercise questions (EQs) in each chapter provide opportunities to check understanding. They often address misunderstandings that commonly appear in examination answers, and will help students to avoid such errors.

The key objective of Modern Education is to give learners the skills, knowledge and attitudes they will need to succeed in a rapidly evolving world. In most developing countries learning resources are scarce. It is therefore necessary that the teacher uses alternative methods such as: Collection from the environment and Improvisation

The teacher should have the capacity and attitude to improvise resources from locally available materials which are often considered waste or valueless. Improvisation helps reduce the cost of teaching and learning since improvised resources cost very little or have no cost at all. In addition improvisation helps demystify science and bring it home to the learner as part and parcel of everyday life.

to the learner as part and parcel of everyday life.

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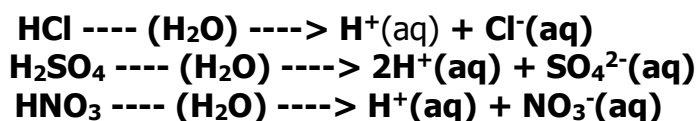
Chapter one: ACIDS BASES AND SALTS

Acids

Acids are compounds that contain Hydrogen (Hydrochloric, HCl; Sulphuric, H₂SO₄; Nitric, HNO₃). Acids are usually compounds of non metals with Hydrogen. .

When dissolved in water, these compounds break up to form hydrogen ions, which is actually what, determines the acidity of acids.

For example, the three acids above react with water in the following ways:



Definition of acid

An acid may be defined as follows:

- ✓ **An acid is any substance which when added water produces hydrogen ions [H⁺].(Arrhenius)**
- ✓ **An acid is substance that can donate a proton(Bronsted and Lowry concept)**

Most acids show acidic properties only when water is present.

Properties of acids

1. Acids have a sour taste.
Think the taste vinegar and lemon juice. But never try to taste laboratory acids.
They may burn you.
2. Acids turn the color of litmus into red.
3. They have PH numbers less than 7.
4. They react with **metals** forming **hydrogen gas** and a **salt**.
For example, when magnesium is dropped into dilute sulphuric acid, hydrogen quickly bubbles off. The bubbles stop when the reaction is over.
The equation for the reaction is:
Magnesium +sulphuric acid → magnesium sulphate + Hydrogen
Mg(s) + H₂SO₄ (aq) → MgSO₄(aq) + H₂(aq)

The magnesium has driven hydrogen out of the acid, and taken its place. Magnesium sulphate is called **a salt**.

When a metal takes the place of hydrogen in an acid, the compound that forms is called a salt.

Salts of sulphuric acid are called sulphates.

Salts of hydrochloric acid are called chlorides.

Salts of nitric acid are called nitrates.

5. They react with **metal carbonates** forming **salt, water** and **carbon dioxide**.

Hydrochloric acid reacts with calcium carbonate like this:

Calcium carbonate + hydrochloric acid → calcium chloride + water carbondioxide



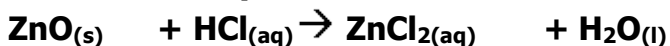
6. They react with **Alkalis** forming **salt** and **water**. For example:

Sodium hydroxide + hydrochloric acid → sodium chloride + water



7. They react with **metal oxides** forming **salt** and **water**.

Zinc oxide + hydrochloric acid → zinc chloride + water



What causes acidity?

Acids have a lot of properties in common. There must be something in them all that makes them to act alike. That "something" is **hydrogen ions**.

Acids contain hydrogen ions

Acids are solutions of pure compounds in water. The pure compounds are molecular. But in water, the molecules break up to form ions. They always give hydrogen ion. For example, in hydrochloric acid:



The more hydrogen ions there are in a solution, the more acidic it is. In other words, the more H^+ ions there are, the lower the PH number.

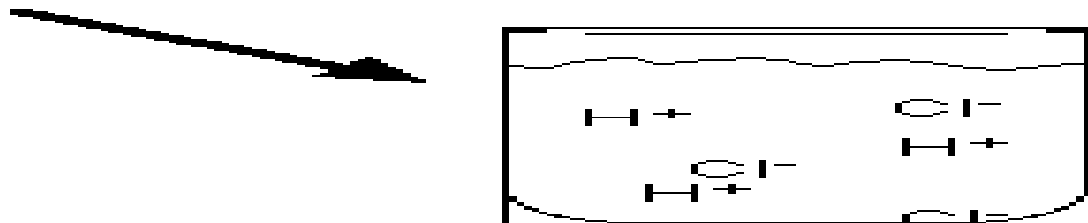
Strong acids and weak acids

As you know, an acid is a compound that produces H^+ in solution. How easily this hydrogen ion is released from the compound is the indication of the acid's strength.

Strong acids

Strong acids nearly all the acid molecules form ions.

For example, when hydrochloric acid is dissolved in water it completely breaks up to form hydrogen ions (H^+) and chloride ions (Cl^-).



So hydrochloric acid is typical strong acid.

Sulphuric acid is also a strong acid. When it dissolves in water, most of its molecules break up and form ions like this:



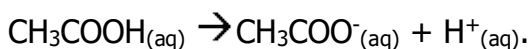
Other strong acids include:

- Nitric acid (HNO_3).
- Hydrobromic acid (HBr).
- Hydroiodic acid (HI).

Weak acid

In a weak acid, only some of the acid molecules form ions, H^+ . Ethanoic acid which is found in vinegar is a typical weak acid. When ethanoic acid is dissolved, only few of its molecules form ions.

Ethanoic acid $\rightarrow$ ethanoate ions + hydrogen ions



Other weak acids are:

- Carbonic acid (H_2CO_3).
- Citric acid

(The more hydrogen ions there are in a solution the more acidic it is and the lower the PH number)

Bases

A base is any substance that can neutralize an acid.

Metal oxides, metal hydroxides, metal carbonates, metal hydrogen carbonate and ammonia are all bases

Some bases dissolve in water. These bases are given special name which is **alkalis**.

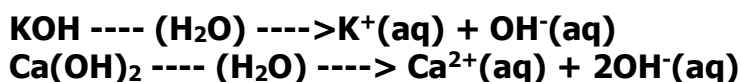
Alkalies

An alkali is a substance that gives OH^- **ions** when dissolved in water.

An alkali can neutralize an acid and destroy its acidity. It does this by removing the H^+ ions and converting them to water.

Alkalies are usually metal hydroxides (MOH). Examples include Sodium Hydroxide, NaOH, Calcium Hydroxide, $\text{Ca}(\text{OH})_2$.

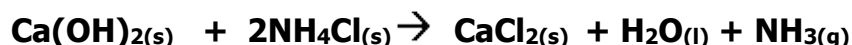
The two bases above react with water in the following ways:



Properties of alkalies

1. Alkalies have a bitter taste.
2. Their solutions turn litmus blue.
3. Their solutions have pH numbers greater than 7.
4. They react with acids to form a salt and water. The reaction is called neutralization.
5. All the alkalies except ammonia will react with ammonium compounds, driving ammonia out. For example:

Calcium hydroxide + ammonium chloride $\rightarrow$ calcium chloride + water + ammonia



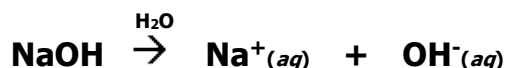
The reaction is used for making ammonia in the laboratory.

What causes alkalinity?

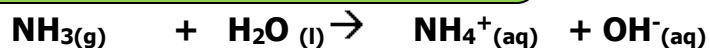
As we have seen, all acidic solutions contain hydrogen ions, H^+ . Alkalies also have something in common.

All alkaline solutions contain hydroxide ions, OH^- .

When alkalies dissolve in water, they release hydroxide ions (OH^-) into solution. For example, a typical alkali is sodium hydroxide (NaOH) produces ions like this:



In ammonia solution, the ammonia molecules react with water molecules to form ions like this:



(The more OH^- ions there are in a solution, the more alkaline it will be and the higher the PH number.)

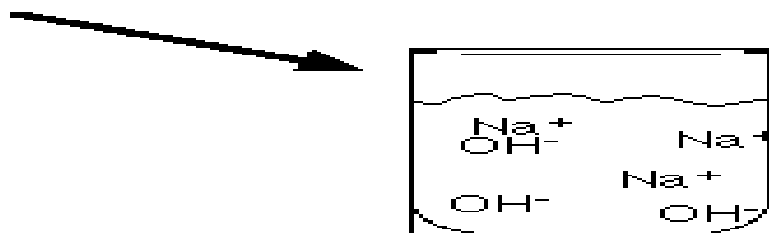
Strong alkalis and weak alkalis

Like acids, alkalis can also be strong or weak.

Strong alkalis:

Like acids alkali can also be strong or weak.

Sodium hydroxide is strong alkali because it exists almost completely as ion in solution. Ammonia solution is a weak alkali because only some ammonia molecules form ions.



Other strong alkalis include potassium hydroxide (KOH) which produces ions like this:



Other strong alkalis include the following:

- Lithium hydroxide, LiOH .
- Calcium hydroxide, $\text{Ca}(\text{OH})_2$.
- Barium hydroxide, $\text{Ba}(\text{OH})_2$

Weak alkali: the example of weak alkali is ammonia

Neutralization reaction

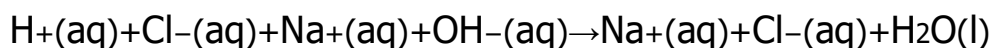
When acids and bases are added to each other they react to neutralize each other if an equal number of hydrogen and hydroxide ions are present.

When this reaction occurs -salt and water are formed.

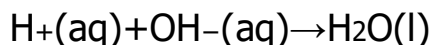
For example, when sodium hydroxide (an alkali) is added to hydrochloric acid (an acid), a neutralization reaction occurs. The end product of this neutralization is sodium chloride and water which are neutral substances.



Chemical reactions occurring in aqueous solution are more accurately represented with a net ionic equation. The full ionic equation for the neutralization of hydrochloric acid by sodium hydroxide is written as follows:



Since the acid and base are both strong, they are fully ionized and so are written as ions, as is the NaCl formed as a product. The sodium and chloride ions are spectator ions in the reaction, leaving the following as the net ionic reaction.



Tracking neutralization

When you add an alkali to an acid you can track the neutralization in several ways

- **PH.** As hydrogen ions are removed the pH of the solution rises
- **Conductivity:** as hydrogen ions get removed the solution is less able to conduct electricity. Conductivity reaches its lowest when neutralization is complete.
- **Temperature:** neutralization is exothermic so the temperature of the solution rises until the reaction is complete.

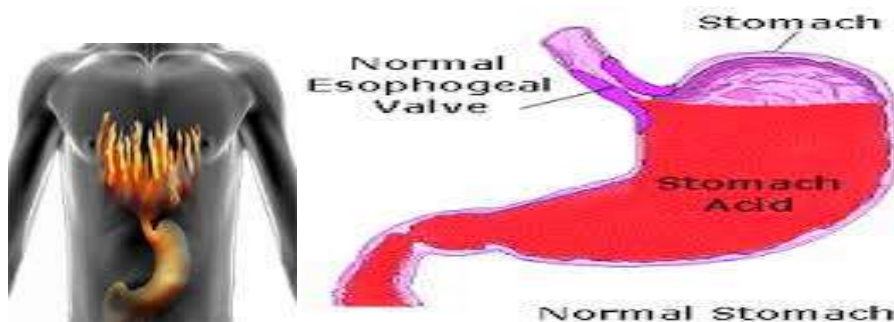
Daily-life uses of neutralization

1. Insect stings:



When a bee stings it injects an acidic liquid into the skin. The sting can be neutralized by rubbing on calamine lotion, which contains zinc carbonate, or baking soda, which is sodium hydrogen carbonate

2. Indigestion



To digest food properly, the liquid in your stomach must be acidic. But too much acidity leads to indigestion, which can be very painful.

To cure indigestion, you must take an ant acid tablet to neutralize the excess acid. For example, a drink of sodium hydrogen carbonate (baking soda), or indigestion tablet.

3. Soil treatment



Most plants grow best when the PH of the soil is close to 7. If the soil is too acidic, or too alkaline, the plants grow badly or not at all.

Chemicals can be added to soil to adjust its PH. Most often, soil is too acidic, so it is treated with quicklime (calcium oxide), slaked lime (calcium hydroxide), or chalk (calcium carbonate). These are all bases, and are quite cheap.

4. Factory waste



Liquid waste from factories often contains acid. If it reaches a river, the acid will kill fish and other river life. This can be prevented by adding slaked lime to the waste, to neutralize it.

Indicators

It will be too dangerous to taste a liquid to find out it is acidic. Chemists use substances called indicators which change color when they are added to acids or alkalis. For example, litmus turns red when added to an acid and turns blue when it is added to an alkali.

The following table shows some color indicators and their colors in acidic and basic medium.

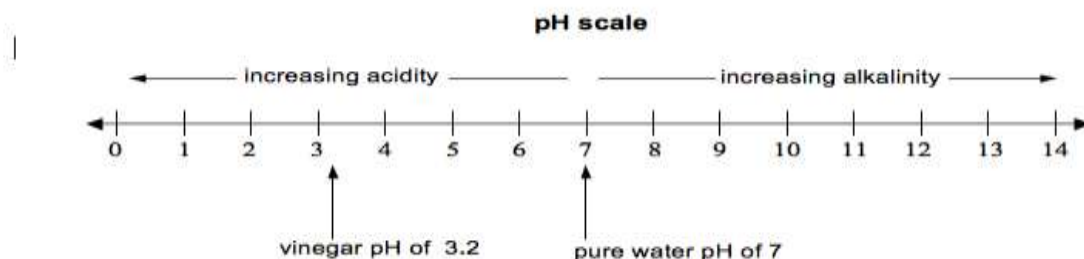
Indicators	Colour in acid solution	Colour alkali solution
Blue litmus	Red	Blue
Methyl orange	Pink	Yellow
Phenolphthalein	Colorless	pink
Red litmus	Red	Blue

These indicators tell chemists whether a substance is acid or alkali or neutral.

The PH scale

PH is a measurement of the acidity or alkalinity of a solution. It provides a value on a scale from 0 to 14 where 7 is neutral, less than 7 is acidic, and greater than 7 is alkaline (or basic). The closer you move towards 0, the more a solution is acidic, and the closer you move to 14, the more a solution is alkaline.

PH is often depicted on a graphical scale as shown below:



- When water has an equal concentration of H^+ ions and OH^- ions, it is said to be neutral ($pH=7$)
- When water has a greater concentration of H^+ ions, it is said to be acidic ($pH<7$)
- When a solution has a greater concentration of OH^- , it is said to be alkaline ($pH>7$)

PH Calculations

We are getting to the part with the calculations here. Now that we know what the pH scale is, let us learn how to compute the pH of a solution.

The equation you need for these calculations is simple:

$$PH = -\log [H^+]$$

In this equation, $[H^+]$ refers to the concentration of the hydrogen ions of the acid in the solution you are looking at. As a result, you will need to be given (or calculate) the concentration of acid present before you can do this problem. Once you know this, it's just a matter of plugging this equation into your calculator. $[H^+] = 10^{-PH}$

Example 1

What's the pH of a 0.01 M solution of HCl?

Solution

To solve this, you only need to realize that in this case $[H^+] = 0.01$. After that, it's just a matter

$$PH = -\log [H^+]$$

$$PH = -\log [1 \times 10^{-2}]$$

$$PH = 2$$

Example 2

Calculate the PH of a water solution containing 0.001M hydrogen ions.

Solution

Given

$$[H^+] = 0.001M = 1 \times 10^{-3}$$

$$PH = ?$$

$$PH = -\log [1 \times 10^{-3}] = -(-3) PH = 3$$

Example 3

What is the PH of an acid solution that contains a specific hydrogen ion concentration $[H^+]$ of $1.4 \times 10^{-5} M$?

Solution

$$\text{Given: } [H^+] = 1.4 \times 10^{-5} M$$

$$PH = ?$$

$$PH = -\log [1.4 \times 10^{-5} M]$$

$$PH = 4.85$$

Example 4

Calculate the hydrogen ion concentration of a hydrochloric acid solution whose PH number is 2.

Solution

Given

$$[H^+] = ?$$

$$PH = 2$$

$$[H^+] = 10^{-PH}$$

$$[H^+] = 10^{-2}$$

$$= 0.01$$

POH (power of hydroxide ion)

POH is the power of hydroxide ion concentration and it is given as follows:

$$POH = -\log [OH^-]$$

$$[OH^-] = 10^{-POH}$$

$$[H^+][OH^-] = 1 \times 10^{-14} \text{ at } 25^\circ C$$

$$-\text{Log} [\text{H}^+] - \text{log} [\text{OH}^-] = -(-14)$$

$$\text{PH} + \text{POH} = 14$$

Example 1

What is the POH of a sodium hydroxide solution that contains 1×10^{-13} M of hydroxide ions?

Solution

Given

$$[\text{OH}^-] = 1 \times 10^{-13} \text{ M}$$

$$\text{POH} = ?$$

$$\text{POH} = -\text{log} [\text{OH}^-]$$

$$\text{POH} = -\text{log} [1 \times 10^{-13}]$$

$$\text{POH} = 13$$

Example 2

What is the hydroxide ion concentration of a sodium hydroxide solution that has a POH of 12?

Solution

Given

$$[\text{OH}^-] = ?$$

$$\text{POH} = 12$$

$$[\text{OH}^-] = 10^{-\text{POH}}$$

$$[\text{OH}^-] = 10^{-12}$$

$$[\text{OH}^-] = 1 \times 10^{-12} \text{ M}$$

Neutral substances

Many substances do not affect the color of indicators, so they are not acids or bases. They are neutral.

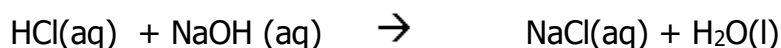
Neutral substances are substances that are neither acidic nor basic.

Examples are pure water, aqueous solution of sodium chloride and sugar.

Preparation of salts

A salt is a compound that is formed when the hydrogen ions of an acid is replaced by a metal ions.

Sodium chloride is produced as part of the neutralization reaction between hydrochloric acid and sodium hydroxide.



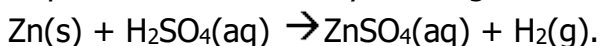
Salts can be classified as those which are soluble in water (**soluble salts**) or those which are not soluble in water (**insoluble salts**).

Soluble salts	Insoluble salts
All sodium, potassium and ammonium salts	
All nitrates	
All chlorides	<i>except</i> silver and lead chlorides
All sulphates	<i>except</i> calcium, barium and lead sulphates
Sodium, potassium and ammonium carbonates but all other carbonates are insoluble	

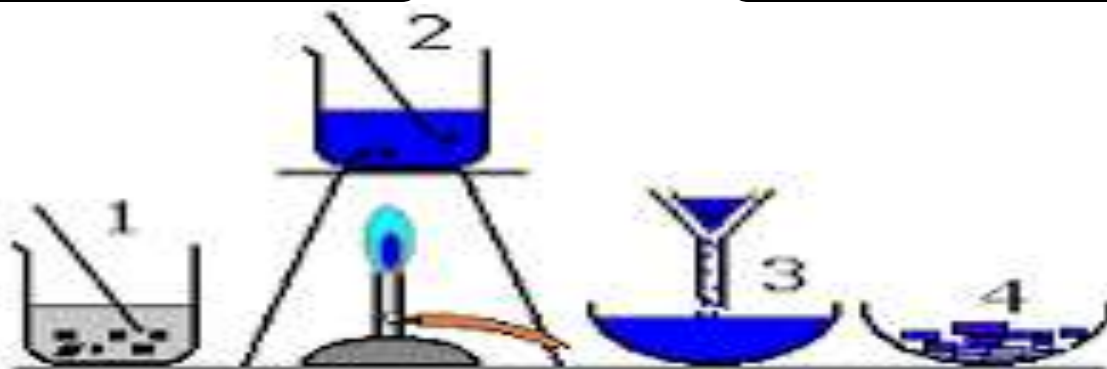
Methods of preparing soluble salts

1. Acid + metal

Some salts can be made by reacting dilute acids with metals. For example, zinc sulphate can be made by reacting dilute sulphuric acid with zinc:



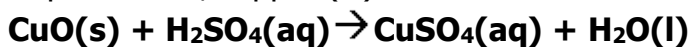
This method above is not suitable for all metals, or all acids. It is fine for magnesium, aluminum, zinc and iron. But the reaction for sodium, potassium and calcium with acid are dangerously violent. The reaction of lead is too slow, copper, silver and gold do not react acids at all. The diagrams below show the steps for preparing zinc sulphate salt.



2. Acid + insoluble

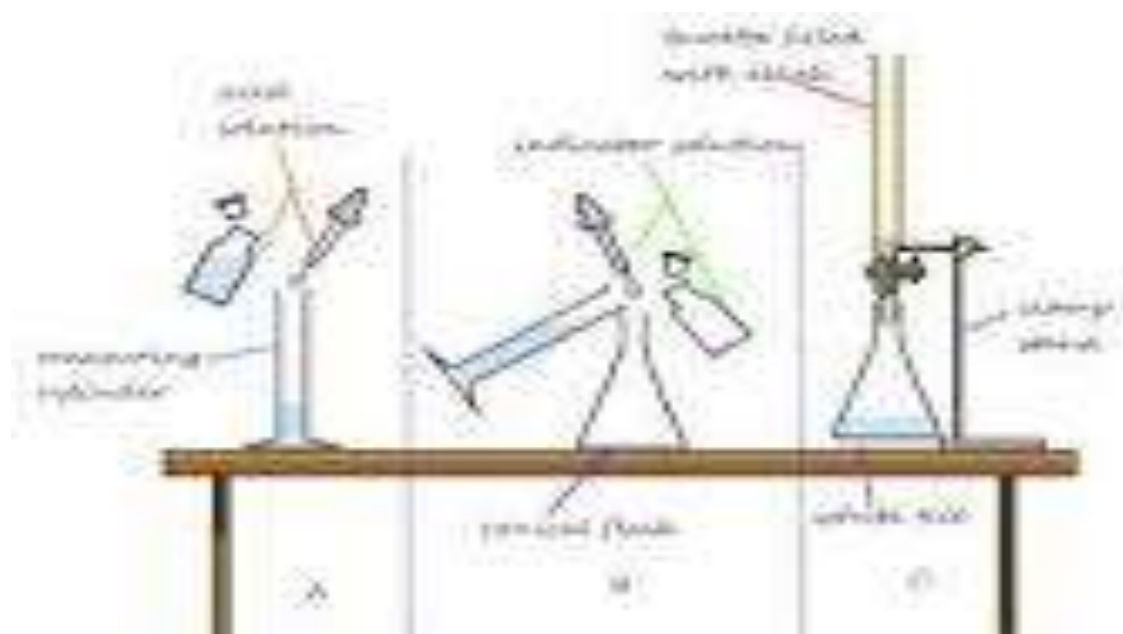
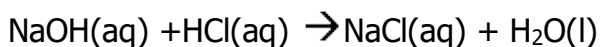
Some salts can also be made by reacting acids with insoluble bases.

Copper (II) oxide is an insoluble base. Although copper will not react with dilute sulphuric acid, copper (II) oxide will. The salt that forms is copper (II) sulphate:



3. Acid + Alkali

The reaction of sodium with acids is very dangerous. So sodium salts are usually made by starting with sodium hydroxide. This reaction can be used to make sodium chloride:



4. Acid + metal carbonate

This method can be used to prepare the salts of many metals. For example, sodium chloride can be made by reacting hydrochloric acid with sodium carbonate.

Hydrochloric acid + sodium carbonate $\rightarrow$ sodium chloride + water + carbon dioxide



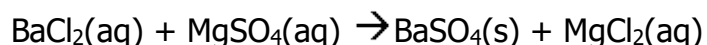
Making insoluble salts

The salts made so far have all been soluble. They were obtained as crystals by evaporating solutions. But not all salts are soluble.

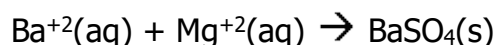
Insoluble salts can be made by process called **precipitation**.

For example, insoluble barium sulphate is precipitated when solutions of barium chloride and magnesium sulphate are mixed.

The equation for the reaction is:



You could also write in a shorter way as:



These are the steps for obtaining barium sulphate:

1. Solutions of barium chloride and magnesium sulphate are mixed.
A white precipitate of barium sulphate forms at once.
2. The mixture is filtered. The barium sulphate gets trapped in the filter paper.
3. Then it is rinsed with distilled water.
4. Then it is put in a warm oven to dry.

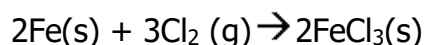
Barium sulphate could also be made from barium nitrate and sodium sulphate, for example, since these are both insoluble. As long as barium ions and sulphate ions are present, barium sulphate will be precipitated.

To precipitate an insoluble salt, you must mix a solution that contains positive ions with one that contains its negative ions.

Making salts by combining elements

Some salts contain just two elements. They can be made by direct combination of the elements. For example, you can make Iron (III) chloride, FeCl_3 , by heating iron in a stream of chlorine.

Iron + Chlorine $\rightarrow$ Iron (III) chloride



Types of salts

Salts are compounds formed when all or part of replaceable hydrogen of acid is replaced by metal. There are different types of salts

Normal salts: are salts formed when all replaceable hydrogen ions of an acid are replaced. Examples are sodium chloride potassium sulphate and zinc nitrate.

Acid salts: are salts formed when only some of replaceable hydrogen ions of acid are replaced examples are NaHCO_3 , $\text{Mg}(\text{HCO}_3)_2$ and KHSO_4

Double salts: are salts which dissociate in water to produce three different types of ions.

Chapter Review questions

Part one circle the correct answer.

- Which of the following statements is not true dilute hydrochloric acid and dilute sulphuric acid
 - Both react with magnesium
 - Both have PH less than 7
 - Both conduct electricity
 - Both have concentration H^+ equal to the concentration of OH^-
- WHICH OF THE FOLLOWING IS PROPERTY OF ALL ACIDS
 - Turn litmus paper into blue
 - React with salts to form a base
 - React with carbonates to produce carbon dioxide
 - React ammoniumsalts to produce ammonia gas
- $AgNO_3(aq) + KI(aq) \longrightarrow AgI(s) + KNO_3(aq)$**
The above equation is represents a reaction which could be classified as
 - Acid base reaction
 - Decomposition reaction
 - Precipitation reaction
 - Neutralization REACTION
- WHICH OF THE FOLLOWING IS NOT PROPERTY OF ABASE
 - Soapy to touch
 - Turn litmus paper into red
 - Have PH greater than 7
 - React acid to form salt and water
- The PH of a weak alkali is between
 - 0 and 14
 - 7 and 10
 - 4 and 7
 - 11and 14
- An acid react with alkali to give
 - Hydrogen and water
 - Hydrogen and salt
 - A salt only
 - Salt and water
- Hydrogen gas is evolved when an acid react with
 - Another acid
 - An alkali
 - A non-metal
 - A metal
- What are the products of this reaction $HNO_3 + NaOH$
 - Salt and water
 - Hydrogen and salt

- C. Water only
D. Hydrogen water and salt
9. Neutral solutions have PH of
A. 2
B. 7
C. 0
D. 14
10. A solution of PH of 8 is
A. Slightly basic
B. Slightly acidic
C. Strongly acidic
D. Strongly alkali
11. A WORD EQUATION IS SHOWN BELOW
Hydrochloric acid + magnesium $\longrightarrow$ magnesium chloride + hydrogen
What is the symbol equation for the reaction?
- A. $\text{HCl}_{(\text{aq})} + \text{Mg}_{(\text{s})} \longrightarrow \text{MgCl}_{2(\text{aq})} + \text{H}_{(\text{g})}$
B. $\text{HCl}_{(\text{aq})} + \text{Mg}_{(\text{s})} \longrightarrow \text{MgCl}_{2(\text{s})} + \text{H}_{2(\text{g})}$
C. $\text{HCl}_{(\text{aq})} + \text{Mg}_{(\text{s})} \longrightarrow \text{MgCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$
D. $\text{HCl}_{(\text{aq})} + \text{Mg}_{(\text{s})} \longrightarrow \text{MgCl}_{2(\text{s})} + \text{H}_{2(\text{g})}$
12. Which of the following is not characteristics of acid
A. Changes the color of an indicator
B. Has a bitter taste
C. Ionized in water
D. Produces hydrogen ion
13. Which gas is produced when acid react with metal carbonates
A. Hydrogen
B. Carbon dioxide
C. Nitrogen
D. Ammonia
E. None
14. The following table shows the preferred as soil PH conditions for different vegetables which vegetable grows over the largest PH range

Vegetable	SOIL PH
A	4.5---6
B	5.5---6.5
C	5.5---7.5
D	6.0....7.5

PART TWO STRUCTURED QUESTIONS

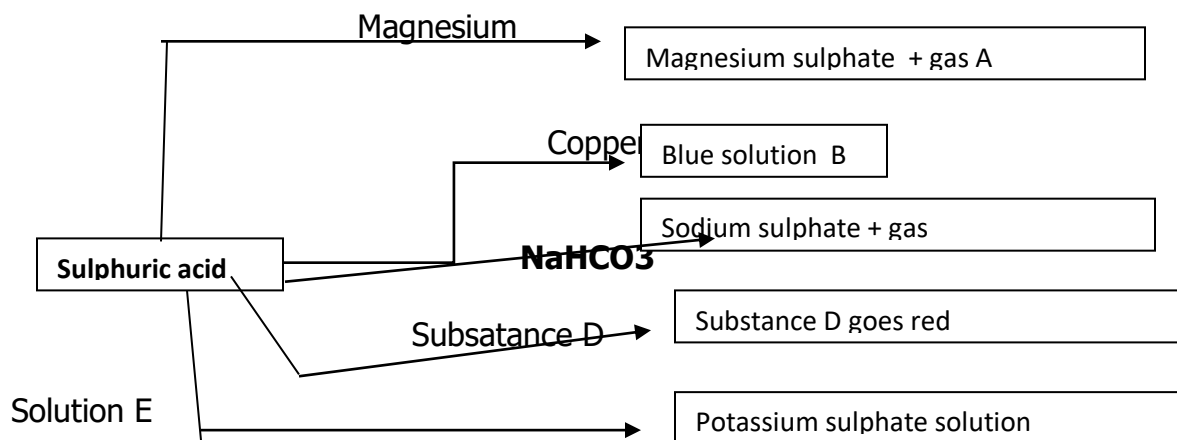
1. GIVE DEFINATION FOR AN ACID?

.....

GIVE DIFFINATION FOR A BASE

.....

2. The diagram below shows some dilute sulphuric acid use it to answer the that flow.



Name and give the formula of each of following

- GAS A.....
- Solution B.....
- Gas C.....
- Substance D.....
- Solution E.....
- How would you test gas A?
.....

3. School students tested some solutions with universal indicator they wrote down their Ph
But they forget to write their names of the solution can you help them by matching the PH to the correct solution **1 5 7 14**

Solution tested	PH
DISTILLED WATER	
SULPHURIC ACID	
SODIUM HYDROXIDE	
VINEGAR	

4. What are salts? Give three examples we can prepare soluble salts and one we can prepare insoluble salts?

.....

 5, what is precipitation Reaction? Givean example?

-
 5. Name the acid and metal would you use for making
 A. Zinc chloride
 B. Magnesium sulphate
 C. How would you obtain lead nitrate starting with insoluble lead (ii) carbonate ?

6. Show witch property described acid and which describe base

Property	Base or acid
Ph less than 7	
Bitter taste	
Ph greater than 7	
Dissolve in water to produce OH ⁻ ions	
Dissolve in water to produce H ⁺ ions	
Turn litmus paper into red	
React metals to produce hydrogen gas	

7. Name three weak acid and three strong acids?

8. Write the balanced chemical equation for the reaction between sodium hydroxide and
 A. Sulphuric acid
 B. Nitric acid

9. Complete the following table which covers the different methods for preparing salts.

Method of preparation	Name of the salt formed	Two substance used in preparation
Aid +alkali	Sodium sulphate	and
Acid +metal		and hydrochloric acid
Acid +insoluble base	Copper ii sulphate	and,.....
Acid +carbonate	Magnesium.....+.....	and
Precipitation	Bariumsulphate	and.....

Chapter two: METALS

There are over 105 different elements, only 84 are metals and 21 are non-metals. This means that over three-quarters of the elements are metals.

Properties of metals

Physical properties of metals

1. They are **strong** under tension and compression. That means they can withstand stretching and crushing without breaking.
2. They are **malleable**. That means they can be hammered and bent into shape without breaking.
3. They are **ductile**. They can be drawn out to make wires.
4. They are **sonorous**. They can make a ringing noise when you strike them.
5. They are **shiny** when polished.
6. They are good conductors of electricity and heat.
7. They have high melting and boiling points (they are all solid at room temperature, except mercury).
8. They have high densities. They feel "heavy".

Chemical properties

1. They react with oxygen to form oxides. For example, magnesium burns in air to form magnesium oxide. Metal oxides are bases, which mean they react with acids to form salts.

Magnesium + oxygen → magnesium oxide

2. When metals form ions, the ions are positive. for example, in the reaction between magnesium and oxygen, magnesium ions (Mg^{+2}) and oxide ions (O^{-2}) are formed.

Some of the metals.

Metal	Symbol	Metal	Symbol
Aluminum	Al	Cobalt	Co
Calcium	Ca	Nickel	Ni
Copper	Mg	Mercury	Hg
Gold	Au	Platinum	Pt
Iron	Fe	Chromium	Cr
Lead	Pb	Scandium	Sc
Magnesium	Mg	Beryllium	Be
Potassium	K	Strontium	Sr

Silver	Ag	Radium	Ra
Sodium	Na	Cadmium	Cd
Tin	Sn	Rubidium	Rb
Zinc	Zn	Cesium	Cs

All metals are different

The properties we see in the last page are typical of metals. But not all metals have all these properties. For example



Iron is a typical metal. It is used for gate like these because it is both malleable and strong. It is used for anchors because of its high density. It melts at 1530°C . But unlike most other metals it is magnetic. Sodium is quite different. It is so soft that it



Can be cut with a knife. It melts at only 98°C .

It is so light that it floats on water, but it reacts

Immediately with the water, forming a solution.

No good for gates.

Gold melts at 1064°C . Unlike most other

metals it doesn't form an oxide—it is very

un-reactive. But it is malleable and ductile

And looks attractive. So it is used for making

Jewelers.



(No two metals have exactly the same properties.)

Comparing metals with non—metals

Only 21 of the elements are non-metals. Non-metals are quite different from metals. They usually have these properties:

- They are not strong, or malleable, or ductile, or sonorous. In fact, when solid non-metals are hammered, they break up—they are brittle.
- They have lower melting and boiling points than metals. (one of them is a liquid, and eleven are gases at room temperature).
- They are poor conductors of electricity. Graphite (carbon) is the only exception. They are also poor conductors of heat.
- They have low densities.
- Like metals, most of them react with oxygen to form oxides. For example sulphur reacts with oxygen to form sulphur dioxide as shown by this equation:

$$\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)}$$
 Sulphur + Oxygen → Sulphur dioxide.
 But unlike metal oxides, non—metal oxides are not bases. Many of them dissolve in water to give **acidic solutions**.
- When they form ions, the ions are negative. Hydrogen is an exception—it forms the ion H^+ .

Some of the non –metals

Bromine	Br
Carbon	C
Chlorine	Cl
Helium	He
Iodine	I
Oxygen	O
Sulphur	S
Nitrogen	N

Metals and reactivity

On the last pages you saw that all metals are different. The next few pages compare the way some metals react, to see how different they are.

The reaction between metals and Oxygen

This table shows the results of the reactions of some metals with oxygen

Metal	Behavior	Order of reactivity	Product
Sodium	Catches fire with only a little heating. Burns fiercely with a bright yellow flame.		Sodium peroxide, Na_2O_2 , a pale yellow powder.
Magnesium	Catches fire easily. Burns with a blinding white flame.		Magnesium oxide, MgO , a white powder.
Iron	Does not burn, but the hot metal glows brightly in oxygen, and gives off yellow sparks.		Iron oxide, Fe_3O_4 , a black powder.
Copper	Does not burn, but the hot metal becomes coated with a black substance.		Copper oxide, CuO , a black powder
Gold	No reaction, no matter how much the metal is heated.	Least reactive	

If a reaction takes place, the product is an oxide. Sodium reacts the most vigorously with oxygen. It is the **most reactive** of the five metals. Gold does not react at all—it is the **least reactive** of them. The arrow on the table shows the **order of reactivity**.

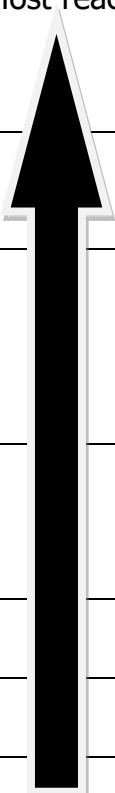
The reaction between metals and water

Metals also show differences in the way they react with water. For example:

Sodium reacts violently with cold water, floating over the surface. Hydrogen gas and a clear solution of sodium hydroxide are formed. But the reaction between calcium and cold water is slower. Hydrogen bubbles off and a cloudy solution of calcium hydroxide forms. Magnesium reacts very slowly with cold water, but vigorously when heated in steam: it glows brightly. Hydrogen and solid magnesium oxide are formed.

This table shows the results for other metals too:

Metal	Reaction	Order of reactivity	Products
Potassium	Very violent with cold water. Catches fire.	Most reactive	Hydrogen and a solution of potassium hydroxide, KOH.
Sodium	Violent with cold water		Hydrogen and a solution of sodium hydroxide, NaOH.
Calcium	Less violent with cold water		Hydrogen and calcium hydroxide Ca(OH)_2 , which is only slightly soluble in water.
Magnesium	Very slowly with cold water, but vigorous with steam.		Hydrogen and solid magnesium oxide, MgO.
Zinc	Quite slow with steam		Hydrogen and solid zinc oxide, ZnO.
Iron	Slow with steam		Hydrogen and solid iron oxide, Fe_3O_4
Copper Gold	No reaction	Least reactive	



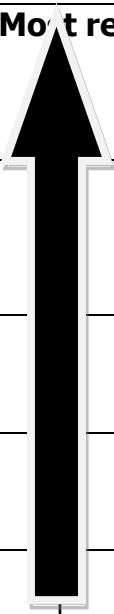
Notice that the first three metals in the list produce hydroxides. The others produce oxides, if they react at all.

Now compare this table with one on the opposite page. Is sodium more reactive than iron each time? Is iron more reactive than copper each time?

The reaction of metals with dilute hydrochloric acid

Some metals react with dilute acid, some don't. When they do, they drive out or displace hydrogen from the acid. Hydrochloric acid has the formula HCl(aq) .

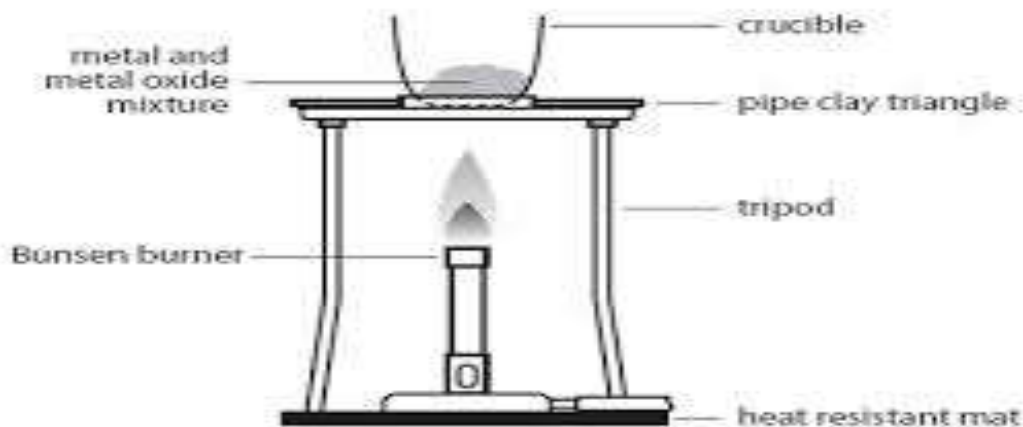
Compare these results:

Metal	Reaction with hydrochloric acid	Order of reactivity	Products
Magnesium	Vigorous		Hydrogen and a solution of magnesium chloride, MgCl_2 .
Zinc	Quite slow		Hydrogen and a solution of zinc chloride, ZnCl_2 .
Iron	Slow.		Hydrogen and a solution of iron(II) chloride, FeCl_2 .
Lead	Slow, and only if the acid is concentrated.		Hydrogen and a solution of lead(II) chloride, PbCl_2 .
Copper Gold	No reaction, even with concentrated acid.		
		Least reactive	

Now compare this table with the last two tables. Is iron always more reactive than copper? Is magnesium always more reactive than iron.

Competition between metals for oxygen

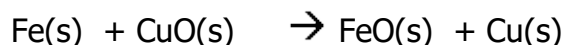
The reaction with oxygen, water, and hydrochloric acid show that iron is more reactive than copper. Now look at this experiment.



In this experiment powdered iron and copper (II) oxide are mixed and then heated together. The reaction starts as soon as the mixture is heated. The mixture glows, even after the Bunsen is removed. Iron (II) oxide and copper are formed.

Here iron and copper are competing for oxygen. Iron wins:

Iron + copper oxide $\rightarrow$ Iron oxide + copper



By taking away the oxygen from copper, iron is acting as a reducing agent. Other metals behave in the same way when heated with oxides of less reactive metals.

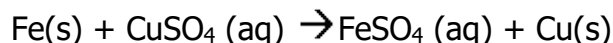
(When a metal is heated with oxide of a less reactive metal, it will remove the oxygen from it. The reaction is exothermic.)

Displacement of one metal by another

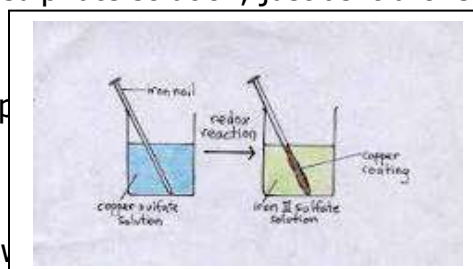
An iron nail is placed in copper (II) sulphate solution. Soon copper appears on the nail. The solution turns green.

Here iron and copper are competing to be the compound in solution. once again iron wins. It drives out or displaces copper from copper(II) sulphate solution, just as it drove it from its oxide. Green iron(II) sulphate is formed:

Iron + copper (II) sulphate $\rightarrow$ Iron (II)sulphate + Copper



Other metals displace less reactive metals in the same way



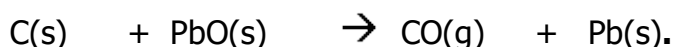
(A metal will always displace a less reactive metal from solutions of its compounds.)

Competition between metals and carbon for oxygen

If magnesium oxide is mixed with powdered carbon and heated in a crucible, no reaction takes place. But, when lead (II) oxide is used instead, silvery beads of molten lead appears.

In competition against magnesium for oxygen, carbon loses. But against lead, a less reactive metal, it wins. It reduces the lead (II) oxide to lead. In this way it proves itself more reactive than lead:

Carbon + lead (II) oxide $\rightarrow$ carbon monoxide + lead



(Carbon is more reactive than some metals. It will reduce their oxides to metals)

The reactivity series

What is a reactivity series?

In the last pages we compared the reactions of different metals. You saw how some were always more reactive than others. In fact, we can list them in order of their reactivity. This list is called reactivity series.

Reactivity series is a list that puts metals in order of their reactivity.

Here it is.

Potassium, K Sodium, Na Calcium, Ca Magnesium, Mg Aluminum, Al	Most active	Above this line, metal oxides can't be reduced by carbon.
Zinc, Zn Iron, Fe Lead, Pb	Increasing reactivity	Metals above this line react with acids, displacing hydrogen.

Copper, Cu
Silver, Ag
Gold, Au

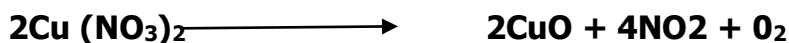
Least reactive

Useful things to remember about the reactivity series

1. The more reactive a metal, the more it "likes" to form compounds. So only copper, silver and gold are ever found as elements in the earth's crust. The other metals are always found as compounds.
2. When a metal reacts, it gives up electrons to form ions. The more reactive the metal, the more easily it gives up electrons.
3. The more reactive the metal, the more stable its compounds. Stable means difficult to break down. For example, when you heat sodium nitrate you get nitrite:

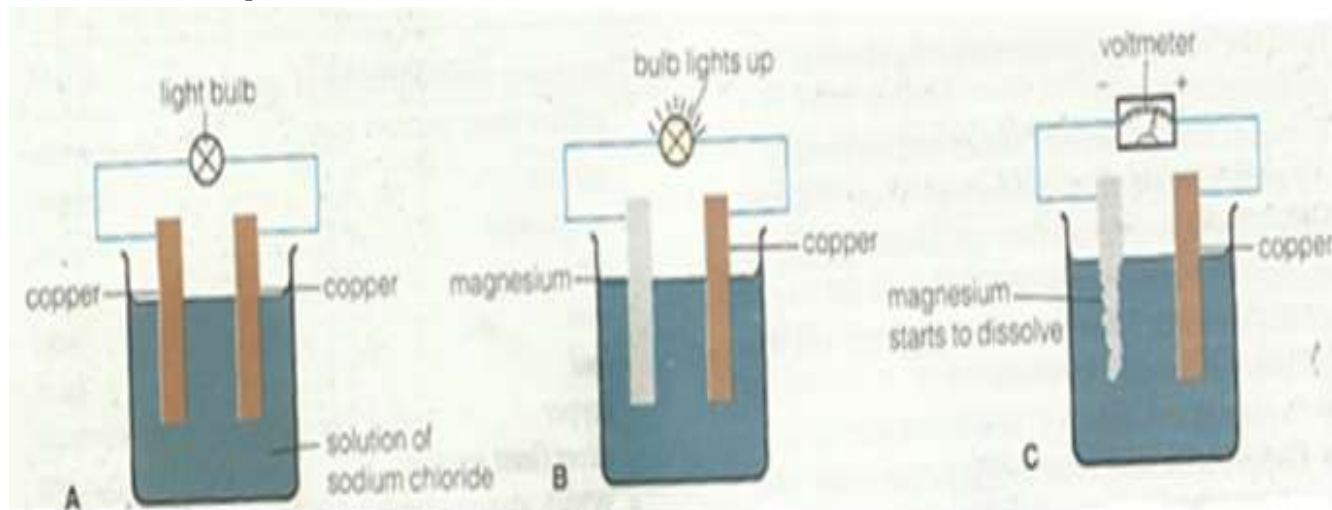


But copper (ii) nitrate breaks down more further to the oxide giving off nitrogen dioxide.



4. The more reactive the metal, the more difficult it is to extract from its compounds (since the compounds are stable). For the most reactive metals you will need the toughest method of extraction: electrolysis.
5. The less reactive metals have been known and used since ancient times, because they are easiest to extract. But aluminum was scarcely used before 1886, when it was first extracted by electrolysis.
6. If you stand two metals in an electrolyte and join them up with a copper wire, you will get a current. Electrons flow from the more reactive metal to the less reactive one.

The reactivity series and cells



The above pictures show cells of different metals.

In cell A, two copper strips are wired up to light bulb and placed in a solution of sodium chloride. The solution is an electrolyte—it can conduct electricity. But nothing happens.

In B, one copper strip is replaced by a magnesium strip, the bulb lights up. Electricity is being produced. Electrons are following through the wires even though there is no battery.

In C, the bulb is replaced by a voltmeter. This measures the push or voltage that makes the electrons flow. It is 2.7 volts. The needle shows the direction of the electron flow.

In A, the bulb does not light because there is no electron flow in the circuit but in B, the bulb lights because electrons do. But where do the electrons come from? The answer is this:

Magnesium can give up electrons more readily than copper. So magnesium atoms give up electrons and go into solution as ions. The electrons flow along the wire to the copper strip. This arrangement is called a **cell**. The magnesium strip is the negative pole of the cell. The copper strip is the positive pole.

A cell consists of two different metals and an electrolyte. In the cell, chemical energy produces electricity. The more reactive metal becomes the negative pole from which electrons flow.

The experiment can be repeated with other metals.

As long as the strips are made of different metals, electrons will flow. But the voltage changes with the metals, as this table shows.

Metal strips	Volts
Copper and magnesium	2.70
Copper and iron	0.78
Lead and zinc	0.64
Lead and iron	0.32

Of these metals, copper and magnesium are furthest apart in the reactivity series. They give the highest voltage. Lead and iron are closest, so they give the lowest voltage.

(The further apart the metals are in the reactivity series, the higher the voltage of the cell)

Extraction of metals from their ores

Metals are found in rocks in the earth's crust. A few metals, such as gold and copper occur as the free metal in nature, uncombined. They are said to occur **(native)**.

Only the metals which are very un-reactive in nature can withstand the action of air and water for thousands of years without being converted into compounds.

Most metals occur as compounds. Rock containing the metal compound is mined. Then machines are used to crush and grind the rock. Next a chemical method must be found for extracting the metal. All these stages cost money. If the rock contains enough of the metal compound to make it profitable to extract the metal, the rock is called an ore.

The method used to extract a metal from its ore depends on the position of the metal in the reactivity series.

The following table shows the methods of extraction of some metals from their ores.

Metal	Method of extraction
Potassium Sodium Calcium Aluminum	Electrolysis
Zinc iron lead	Heating carbon or carbon monoxide
Copper	Roasting in air
Silver Gold	Occur naturally as elements

Examples of different methods of extraction

1. Electrolysis

This is used for extracting aluminum from aluminum oxide or alumina obtained from bauxite. The alumina must be melted first.



2. Heating with carbon monoxide

This is used for extracting iron from haematite in the blast furnace.

Iron (II) oxide + carbon monoxide $\rightarrow$ iron + carbon dioxide



3. Heating with carbon

The zinc ore or zinc blende is first converted into zinc oxide. Then this is heated in a furnace with coke.

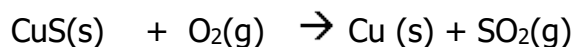
Zinc oxide + carbon $\rightarrow$ Zinc + carbon monoxide



4. Roasting in air

This is used for extracting copper from chalcopyrite.

Copper (II) sulphide + Oxygen $\rightarrow$ Copper + Sulphur dioxide



Making Use of metals

Pure metals and alloys

The way a metal is used depends on its properties:

- ✓ Pure aluminum can be rolled into very thin sheets, which are quite strong but easily cut. So it is used for milk bottle tops and cooking foil.
- ✓ Pure lead is soft, and bends easily without being heated. It also resists corrosion. So it is used to seal off brick work around chimneys.
- ✓ Pure copper is easily drawn into wires, and is an excellent conductor of electricity. So it is used for electrical wiring around the home.

Sometimes a metal is most useful when it is pure. For example, copper is not nearly such a good conductor when it contains impurities.

But many metals are more useful when they are not pure. Iron is the most—widely metal of all and it is almost never used pure. This is because pure iron is no good for building things. It is too soft and stretches easily.

But when a little carbon (0.5%) is mixed with it, the result is mild steel. This is hard and strong. It is used for buildings, bridges, ships and car bodies. Likewise, when nickel and chromium are mixed with iron, the result is stainless steel. This is also hard and rust-proof. It is used for car parts, kitchen sinks, and cutlery.

You can see that the properties of the iron have been changed by mixing other substances with it.

The properties of any metal can be changed by mixing other substances with it. These mixtures are called **alloys**.

The added substances are usually metals, but sometimes non-metals like carbon or silicon. An alloy is usually made by melting the main metal and then dissolving the other substances in it.

Turning a metal into an alloy increases its range of uses.

Uses of pure metals

Pure metal	Uses	Properties that make it suitable
Sodium	A coolant in nuclear reactors. Extraction of titanium.	Conducts heat well. Melts at only 98°C , so the hot metal will flow along pipes. It is more reactive than titanium and melts easily.
Aluminum	Overhead electricity cables(with a steel core for strength) CDs and CD-ROMs.	A good conductor of electricity(not as good as copper, but cheaper and much lighter); resists corrosion. Provides a cheap reflective coating.
Zinc	Coating iron, to give galvanized iron.	Protects the iron from rusting.
Tin	Coating steel cans or 'tins'	Unreactive and non-toxic. Protects the iron from rusting.
Nickel	Electroplating steel	Resists corrosion, sticks well to steel, shiny and attractive to look at.
Titanium	Teeth implants and replacement hip joints.	Light, strong, resists corrosion, non-toxic and ductile so can be easily shaped

Uses of alloys

There are thousands of different alloys. The following table summarizes some alloys, their components, special properties and their uses.

Alloy	Made from	Special properties	Uses
Cupronickel	75% copper 25% nickel	Hard-wearing, attractive silver color	'silver' coins.
Stainless steel	70% iron 20% chromium 10% nickel	Doesnot rust	Car parts, kitchen sinks, cutlery, tanks and pipes in chemical factories.
Manganese steel(Hadfield steel)	85% iron 13.8% manganese 1.2% CARBON	Very hard	Springs
A titanium alloy	92.5% titanium 5% aluminum 2.5% tin	High strength at high temperatures	Jet engine components.
Brass	70% copper 30% zinc	Harder than copper, does not corrode	Musical instruments
Bronze	95% copper 5% tin	Harder than brass, doesn't corrode, sonorous.	Statues, ornaments, church bells.
Solder	70% tin 30%lead	Low melting point	Joining wires and pipes.

More about Aluminum

Aluminium is the most **abundant** metal on Earth. Despite this, it is **expensive**, largely because of the amount of **electricity** used up in the extraction process.

Aluminium ore is called **bauxite** which is Aluminium oxide mixed with impurities like sand and iron oxide. The impurities make it reddish brown. The bauxite is purified to yield a white powder, aluminium oxide, from which aluminium can be extracted.

Extraction of Aluminium

The following are the steps for obtaining Aluminium:

1. When a company discovers Bauxite, it does feasibility study to see if mining is profitable. If yes, the mine is constructed and mining begins.
2. Bauxite usually lies near the surface, so it is easy to dig up. Everything gets coated with red-brown bauxite dust.
3. From the mine, the ore is taken to a bauxite plant, where it is treated to remove the impurities. The result is white **Aluminiumoxide**, or **alumina**.

The extraction process:

The extraction of Aluminium is done by **electrolysis**. But first the aluminium oxide must be made **molten** so that electricity can pass through it. Aluminium oxide has a very high melting point (over $2,000^{\circ}\text{C}$), so it would be expensive to melt it. Instead, it is **dissolved** in **molten cryolite(sodium hexafluoro aluminate, Na_3AlF_6)**, an aluminium compound with a lower melting point than aluminium oxide. The use of cryolite reduces some of the energy costs involved in extracting aluminium.

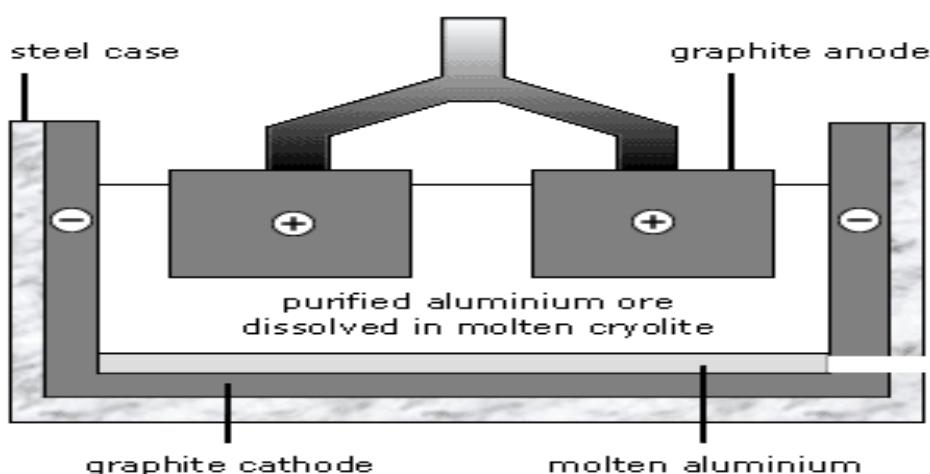
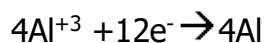


Diagram showing a cell for Aluminium extraction.

The diagram shows an aluminium oxide electrolysis tank. Both the negative electrode (cathode) and positive electrode (anode) are made of **graphite**, a form of carbon.

When the alumina(aluminium oxide) dissolves, it decomposes into aluminium ions (Al^{+3}) and Oxide ions (O^{-2}) which are free to move.

At the cathode: the aluminium ions receive electrons to form neutral aluminium atoms:



The aluminium atoms collect together, and drop to the bottom of the cell as molten metal. This run off at intervals.

At the anodes: the oxide ions give up electrons:



Oxygen gas bubbles off. But unfortunately it attacks the graphite anodes and eats them away, so they must be replaced from time to time.

Some properties of Aluminium

1. Aluminium is bluish-silver, shiny metal.
2. Unlike most metals, it has a low density—it is 'light'.
3. It is a good conductor of heat and electricity.
4. It is malleable and ductile.
5. It is non-toxic.
6. It is not very strong when pure, but it can be made stronger by mixing it with other metals to form alloys.

Some uses of Aluminium

The properties we mentioned above, lead to a wide range of uses for aluminium. These include making the materials listed below:

1. cans.
2. Cooking foil.
3. Saucepans.
4. Racing bikes.
5. TV aerials.
6. Aeroplanes and ships.

More about Iron

Iron is the second most abundant metal in the earth's crust. The chief ore of Iron is **Haematite**. It is mainly iron (III) oxide, **Fe₂O₃**.

Extraction of iron

This section looks at the use of the Blast Furnace in the extraction of iron from iron ore, and the conversion of the raw iron from the furnace into various kinds of steel.

Iron is extracted from its ore by **Blast furnace**, a tall oven.

The raw materials for the extraction of iron are:

1. **Iron ore:** as mentioned above, this is mainly Haematite, Fe₂O₃, mixed with sand.
2. **Limestone:** this is mainly calcium carbonate, CaCO₃.
3. **Coke:** this made from coal, and is almost pure carbon.

These three substances are mixed together to give a mixture called **charge**.

What is happening in the blast furnace?

The blast furnace is like a giant chimney, at least 30M tall. It is made of steel, and lined with fireproof bricks. This is how the process proceeds:

The limestone is used to convert silica (this is the main impurity in iron ore) to calcium silicate, which is molten at the temperature of the blast furnace. This collects as molten slag at the base of the furnace. If the silica is not removed in this way, it clogs up the furnace.

These three substances are mixed to give a mixture called charge the charge is heated in blast furnace. Several reactions take place in it.

The coke reacts with oxygen in the air.



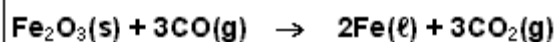
The limestone decomposes to form calcium oxide and carbon dioxide



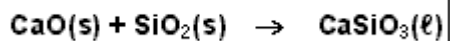
The carbon dioxide reacts with more coke giving carbon monoxide



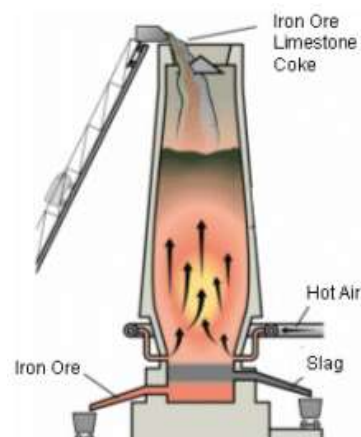
In the hottest part of the furnace (at the bottom) where the temperature reaches 1 800 C⁰, carbon monoxide (CO) reduces the iron oxide:



The calcium oxide combines with the silica, which is acidic, and forms the slag which collects at the bottom of the furnace and floats on the molten iron:



The molten iron is then tapped from the bottom of the furnace and the slag is tapped off from another opening.



Chapter Review questions

Circle the correct answer

1. which of the following metals would displace zinc from solution of zinc sulphate
A. Iron C. magnesium
B. Lead D. copper
2. When copper is extracted from its ore Cu_2S the method used for is
A. electrolysis C. reduction
B. roasting D. None
3. Which of the following elements burns in air to produce an acidic oxide
A. Aluminum C. magnesium
B. Sodium D. sulphur
4. $\text{Mg} + \text{SO}_2 \longrightarrow \text{MgO} + \text{S}$
In this reaction magnesium is acting as
A. Catalyst C. oxidizing agent
B. Reduced agent D. acid
5. Which of the following symbols is the correct formula for potassium
A. P C. Po
B. K D. Ca
6. Malleable means
A. Can be hammered and bent into any shape
B. Can be made wires
C. Make a ringing noise
D. None
7. The only metal that is liquid at room temperature is
A. Sodium C. silver
B. Mercury D. None
8. Rust is formed iron and which other element
A. Hydrogen C. nitrogen
B. Sulphur D. oxygen
9. Which of the following metals can be found in pure state in nature

- A. Lithium B. iron C. gold D. Aluminum
10. In its simplest form bronze is an alloy which is made of
- A. Copper and tin C. copper and Zinc
B. Copper, zinc and nickel D. copper, tin and lead
11. The ability of metal to be drawn into wires is a measure of its
- A. Ductility C. hardness
B. Malleability D. strength
12. $\text{Ca} + \text{FeSO}_4 \longrightarrow \text{CaSO}_4 + \text{Fe}$ in this reaction;
- a. Ca is more reactive than Fe
b. Fe is displaced from the solution
c. The reaction is displacement reaction
d. All of the above
13. Which of the following is **NOT** properties of metals
- a. Malleable
b. Sonorous
c. Weak under tension and compression
d. Shiny when polished
14. When One metal is mixed to another it is called:
- a. Isotopes
b. Alloy
c. Allotrope
d. None
15. The chemical formula of Aluminium Carbonate is:
- a. $\text{Al}_2(\text{CO}_3)_3$
b. $\text{Al}(\text{CO}_3)_3$
c. $\text{Al}_4(\text{CO}_3)_3$
d. None
16. Iron and copper are competing to be the compound in solution so;
- a. Iron win
b. Copper win

- c. No one win
17. Reactivity series of metals are :
- K, Na, Ca, Mg & Al. respectively
 - K, Na, Mg, Ca, respectively
 - K, Mg, Na, Ca, Mg & Al
18. Which substance does not form copper (II) sulphate with warm, dilute sulphuric acid?
- copper
 - copper (II) carbonate
 - Copper (II) oxide
 - copper (II) hydroxide
19. Which one of the following metals reacts most vigorously with cold water
- magnesium
 - copper
 - zinc
 - sodium
20. Which one of the following metals could produce zinc metal, if heated with zinc oxide?
- Aluminium
 - Copper
 - Iron
 - Lead
- 21.. Which one of the following metal mixtures could be separated from one another by adding dilute hydrochloric acid? Assume that the metals simply mixed, and have not formed an alloy.
- Sodium and potassium
 - sodium and calcium
 - magnesium and zinc
 - magnesium and copper

Structured questions

1. **Iron is below magnesium in the** reactivity series. Write an ionic equation for the reaction that occurs when magnesium metal is placed in solution of iron (ii) sulphate?
-
2. A strip of zinc is placed in a beaker of copper (ii) sulphate solution then it is left for a day
- A.** What is the color of solution before the reaction takes place
-

- B. What is the color of solution after the reaction takes place
.....
- C. What is the color of the metal before the reaction takes place
.....
- D. Write the half equation to show what happened to copper (ii) sulphate
.....
- E. Write the half equation to show what happened to zinc atoms?
.....
- F. Which half equation shows reduction
.....
- G. Which half equation shows oxidation
.....
- H. How was the oxidation number of copper changed
.....
- I. How was the oxidation number of zinc changed
.....
- J. Which ion was spectator ion**
.....

3. The table below shows some properties of elements

substance	Reaction with water
A	Reacts with steam but no reaction with cold water
B	Reacts cold water producing a flame
C	Reacts cold water producing bubbles and fizzing sound

- A. Arrange the elements in increasing order of reactivity?**
.....
- B. Write the equation for the reaction between B and cold water?**
.....

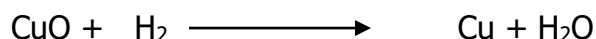
Chapter three: Electrolysis

Oxidation and reduction reactions (Redox reactions)

What is oxidation and what is reduction?

If a substance gains oxygen during reaction it is oxidized

If a substance loss oxygen during reaction it is reduced. For example

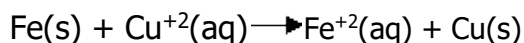


The copper oxide is losing oxygen it is being reduced

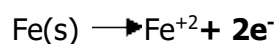
The hydrogen is gaining oxygen it is being oxidized

Reduction and oxidation always takes place together

Oxidation means losing electrons. Reduction means gaining electrons. For example, the reaction between iron and copper (II) sulphate is a typical redox reaction.

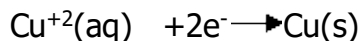


Iron metal(Fe) loses two electrons to form iron ions (Fe^{+2}).



This is called oxidation.iron is said to be oxidized.

Copper ions (Cu^{+2}) gain 2 electrons to form copper metal (Cu).



This is called reduction. The copper is said to be reduced.

To help you remember that oxidation means losing electrons and reduction means gaining electrons, remember OIL and RIG.

(Oxidation Is Loss, Reduction Is Gaining—of electrons).

Oxidation and reduction always occur together. They are called redox reactions.

Oxidation numbers

Oxidation number is the apparent charge that an element has in a compound or ion.

Oxidation numbers are sometimes called **oxidation states**. Oxidation numbers are important in tracking electron movement in redox reactions.

Rules for calculating oxidation numbers

1. The oxidation number of uncombined elements is zero.
2. Elements just bonded to identical atoms, like **O₂** and **H₂**; also have an oxidation number of zero.
3. The oxidation number of a monatomic ion is equal to the charge on that ion.
4. The sum of the oxidation numbers of a compound ion is equal to the charge on that ion.
5. The sum of the oxidation numbers of neutral compound is equal to zero.
6. The oxidation number of oxygen is -2 (except in peroxides where it is -1, and in O₂ where it is 0).
7. The oxidation number of hydrogen is +1 (except in metal hydrides where it is -1).

Example 1

Determine the oxidation number of carbon in the carbon dioxide, CO₂.

Solution

The sum of the oxidation numbers in CO₂ = 0

$$x + 2(-2) = 0$$

$$x - 4 = 0$$

$$x = +4$$

Therefore, the oxidation number of carbon in CO₂ is +4

Example 2

What is the oxidation number of sulphur in potassium sulphate, K₂SO₄?

Solution

The sum of the oxidation numbers in K₂SO₄ = 0

$$2(+1) + X + 4(-2) = 0$$

$$2 + x - 8 = 0$$

$$x - 6 = 0$$

$$x-6=0$$

$$X=+6$$

Example 3

Determine the oxidation number of carbon CO_3^{-2}

Solution

The sum of the oxidation numbers in $\text{CO}_3 = -2$

$$X + 3(-2) = -2$$

$$X - 6 = -2$$

$$X = -2 + 6$$

$$X = +4$$

Example 4

What is the oxidation number of hydrogen in H_2O ?

The sum of the oxidation number of $\text{H}_2\text{O} = 0$

$$2(x) + 1(-2) = 0$$

$$2x - 2 = 0$$

$$2X = 2$$

$$X = 2/2$$

$$X = +1$$

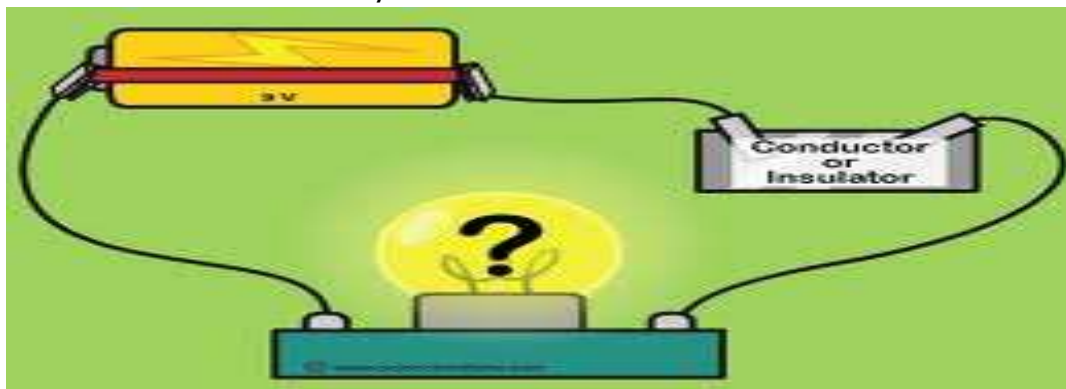
Conductors and insulators

Substances that allow electricity to pass through them are called **conductors**.

Substances that do not allow electricity to pass through them are called **non-conductors or insulators**.

Testing substances to see if they conduct

The circuit below can be used to test if it conducts electricity. The substance we want to test is simply connected into the circuit. If the bulb lights, it means that the connected substance allows electricity to pass and hence it is a conductor. In contrast, if the bulb does not light, that means the substance is an insulator and does not conduct electricity.



The following points show the results from a range of tests:

1. **The only solids that conduct are the metals and graphite:** this conduct because of their free electrons. The free electrons are getting pumped out of one end of the solid by the battery. Electrons then flow in the other end and through the spaces left behind.
2. **Molecular or covalent substances are non-conductors:** this is because they contain no free electrons, or other charged particles, that can flow through them.

Examples of molecular substances are: ethanol, paraffin, petrol, sugar, plastics and sulphur. These substances never conduct whether solid or liquid.

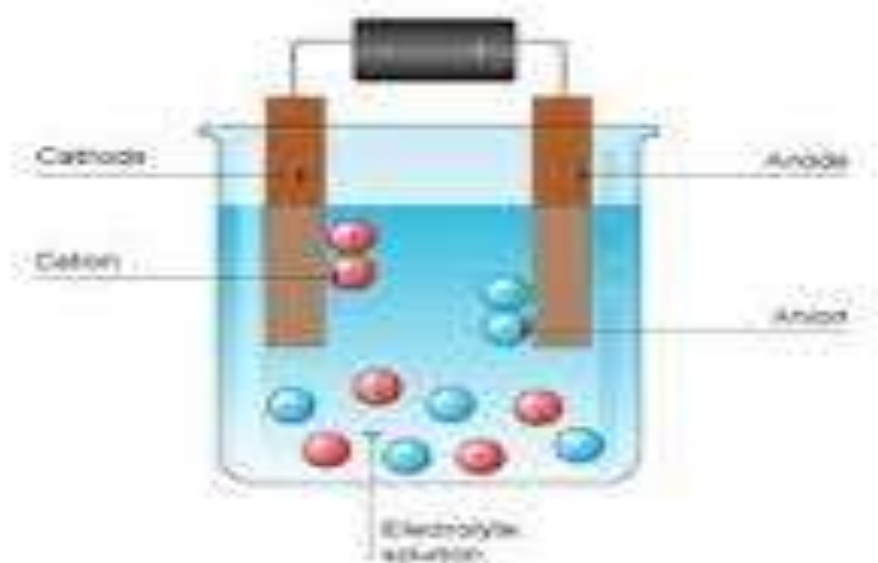
3. **Ionic substances do not conduct when solid. However, they conduct when melted or dissolved in water, and they decompose at the same time:**

An ionic substance contains no free electrons. However, it contains ions which are also charged particles. The ions become free to move when the substance is melted or dissolved in water, and this reason that it conducts electricity.

Electrolysis of molten lead bromide

Electrolysis is the most powerful way to decompose ionic compounds.

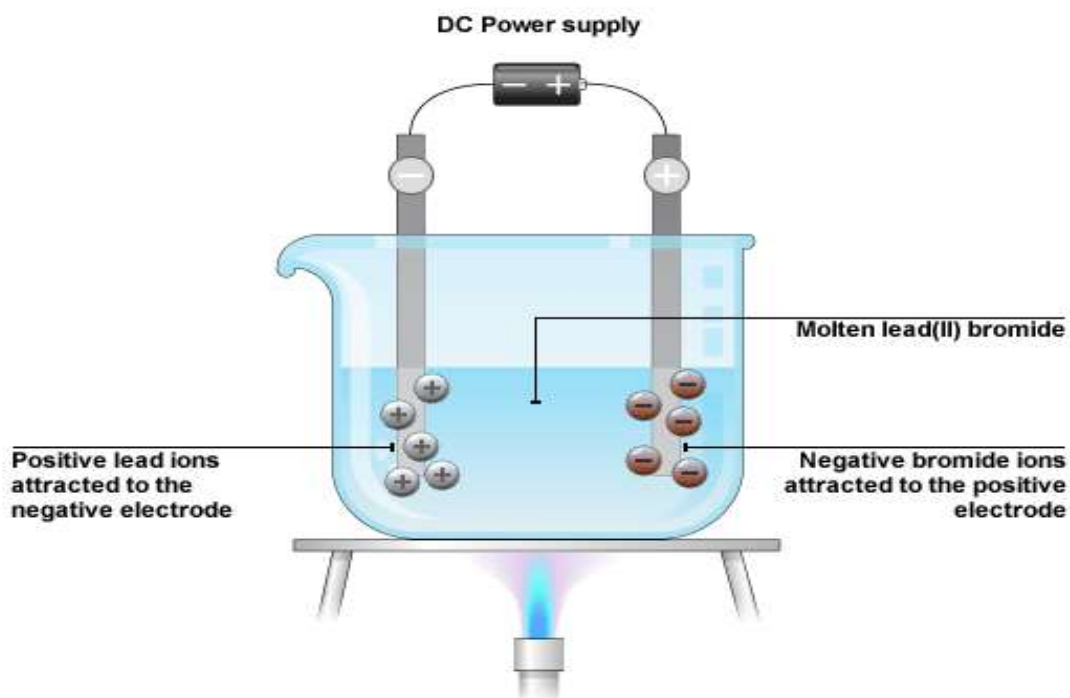
The decomposition caused by electricity is called electrolysis and the liquid that decompose is called electrolyte the apparatus for electrolysis is shown below



The graphite rods carry the current into and out of the molten lead bromide. Conducting rods like these are called electrodes.

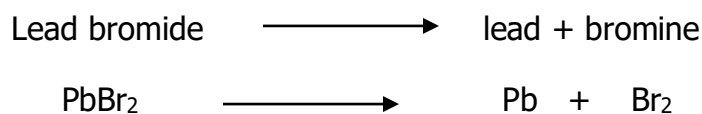
The electrode joined to the negative terminal of battery is called CATHODE and it is negative charged.

The electrode joined to the positive terminal of battery is called ANODE and it is positively charged particles. When the switch is open no electricity can flow.



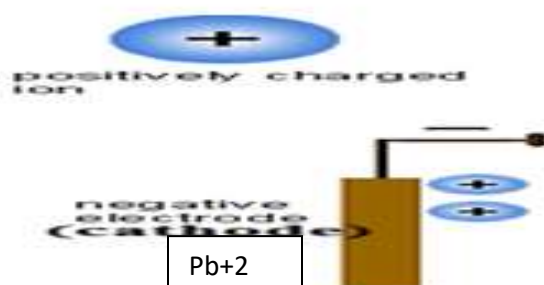
The apparatus for the electrolysis of molten lead bromide is shown above.

Electrolysis: once the switch is closed, electrolysis starts and bromine vapor bubbles out of the molten lead bromide, around the anode. After some, time a bead of molten lead forms at the cathode.



When lead bromide melts its lead ions and bromide ions become free to move. When the switch is closed the electrodes become charged and the ions are immediately attracted to them.

Opposite charges attract so the lead ions attracted to the cathode and bromide ions to the anode.

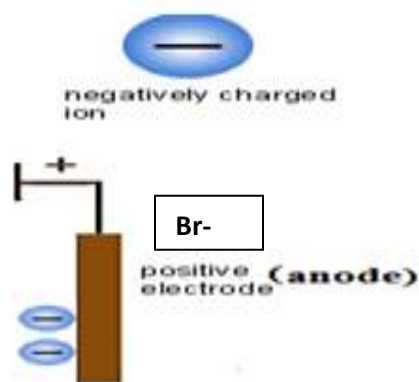


At the cathode: lead ions receive electrons and become lead atoms

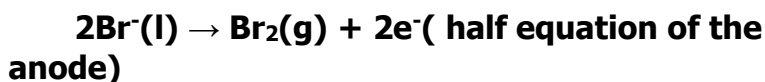


(Reduction occur at the cathode because lead atoms gain electrons)

At the positive electrode (Anode)



At the anode bromide ions give up electrons to become bromine atoms. These pair together to form molecule

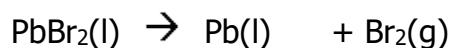


(Oxidation occur at the anode because bromide ions loss electrons to become atoms)

The electrical energy from the battery has caused a

chemical change.

Lead bromide $\rightarrow$ lead + bromine



Why the molten lead bromide conducts: during the electrolysis, each lead ion takes two electrons from the cathode. At the same time, two bromide ions each give an electron to the anode. The effect is the same as if two electrons flowed through the liquid from the cathode to the anode. In other word, the lead bromide is acting as a conductor of electricity.

The electrolysis of other ionic compounds

All ionic compounds can be electrolysed, when they are molten. These are some points to remember about the process:

1. The electrolyte always decomposes. So electrical energy is causing a chemical change.
2. The electrodes are usually made of graphite or platinum. These substances are unreactive or inert. That means they will not react with the electrolyte or the product of electrolysis.
3. Metals always form positive ions. These ions always go to the cathode. So they are called **cations**.
4. Non-metals except hydrogen always form negative ions. These ions always go to the anode. So they are called **anions**.

Remember

- Cations(+) go to the cathode (-)
- Anions (-) go to the anode (+)

The following table shows more about this:

Electrolyte	The decomposition	At the cathode	At the anode
Sodium chloride NaCl	$2\text{NaCl(l)} \rightarrow 2\text{Na(l)} + \text{Cl}_2\text{(l)}$	$2\text{Na}^+ + 2\text{e}^- \rightarrow 2\text{Na}$	$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$
Potassium iodide KI	$2\text{KI(l)} \rightarrow 2\text{K(l)} + \text{I}_2\text{(l)}$	$2\text{K}^+ + 2\text{e}^- \rightarrow 2\text{K}$	$2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$
Copper(II) bromide CuBr₂	$\text{CuBr}_2 \rightarrow \text{Cu(l)} + \text{Br}_2$	$\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu}$	$2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$

5. Electrolysis is the most powerful way to decompose an ionic compound. So it is used in industry to extract reactive metals such as sodium and aluminium from their ores.

Electrolysis of solutions

When a salt such as sodium chloride is dissolved in water, its ions become free to move. So the solution can be electrolysed. But the products may be different from when you electrolyse the molten salt, because water itself produces ions. Although water is a molecular, a tiny fraction of its molecules form ions.



During electrolysis, these H^+ and OH^- ions compete with the metal and non-metal ions from the dissolved salt, to receive or give up electrons. So which ions win?

These are the rules.

At the cathode:

1. The more reactive a metal, the more it 'likes' to exist as ions. So if a metal is very reactive, its ions remain in solution. The H^+ ions accept electrons, and hydrogen molecules are formed.
2. The ions of less reactive metals will accept electrons and form metal atoms, leaving the hydrogen ions in solution.

At the anode:

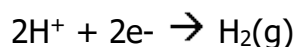
3. If ions of a halogen are present (Cl^- , Br^- , I^-), they will give up electrons more readily than the OH^- ions do. Molecules of chlorine, bromine are formed
4. If no halogen ions are present, OH^- ions will give up electrons more readily than other non-metal ions do, and oxygen is formed.

Electrolysis of sodium chloride solution:

A concentrated solution of sodium chloride contains Na^+ ions and Cl^- ions from the salt, and H^+ and OH^- ions from the water. When the solution is electrolysed using appropriate electrolysis apparatus, positive ions go to the cathode and negative ions go to the anode.

The solution contains Na^+ ions and Cl^- ions from the dissolved salt and H^+ ions and OH^- from the water. The positive ions go to the cathode and the negative ions go to the anode

At the cathode: it is the H^+ ions which accept electrons since sodium is more reactive than hydrogen:

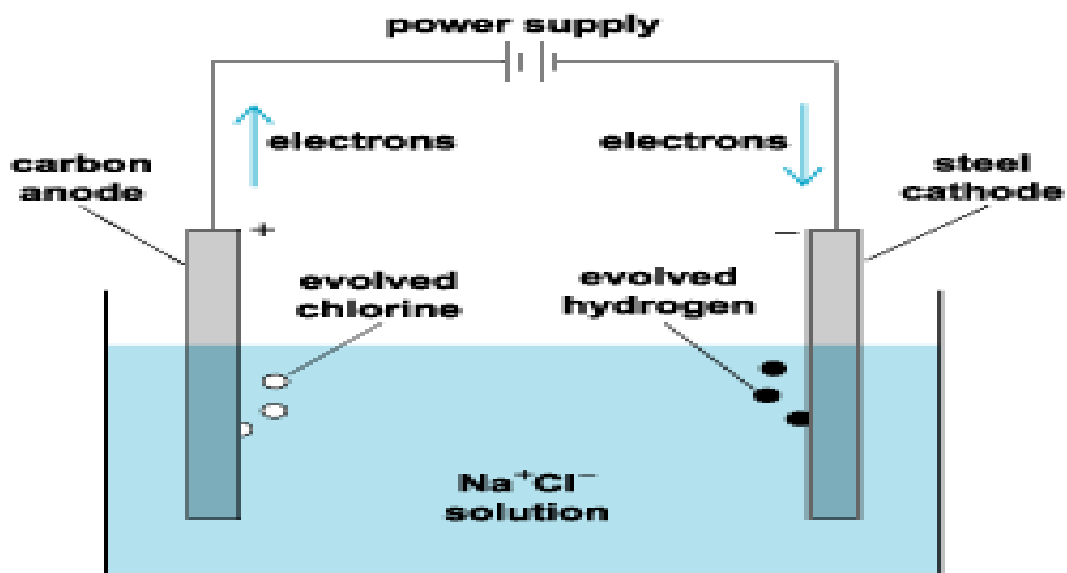


Hydrogen gas bubbles of while Na^+ remains in solution.

At the anode: The Cl^- ions give up electrons more readily than the OH^- ions do. Chlorine gas bubble off:



The OH^- ions remain in solution.



When the hydrogen and chlorine bubble off, Na^+ and OH^- ions are left behind: a solution of sodium hydroxide is formed.

Other salt solutions

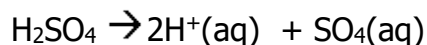
Here are the results of electrolysing some other salt solutions. Check them out. Do they obey the rules given above?

Electrolyte	Cathode	Anode
Potassium bromide, KBr(aq)	Hydrogen $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	Bromine $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$
Sodium iodide NaI	Hydrogen $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	Iodine $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$
Magnesium sulphate MgSO_4	Hydrogen $4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2(\text{g})$	Oxygen $4\text{OH}^- \rightarrow 2\text{H}_2\text{O(l)} + \text{O}_2 + 4\text{e}^-$

Lead(II) nitrate, $\text{Pb}(\text{NO}_3)_2$	Lead $2\text{Pb}^{+2} + 4\text{e}^- \rightarrow 2\text{Pb} + 2\text{e}^-$	Oxygen $4\text{OH}^- \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2 + 4\text{e}^-$
Copper (II) chloride, CuCl_2	Copper $\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu}$	Chlorine $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$
Silver nitrate $\text{AgNO}_3(\text{aq})$	Silver $4\text{Ag}^+ + 4\text{e}^- \rightarrow 4\text{Ag}$	Oxygen $4\text{OH}^- \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2 + 4\text{e}^-$

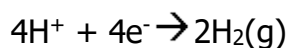
Electrolysis Of Dilute sulphuric acid

Sulphuric acid has the formula H_2SO_4 . In water it forms ions:

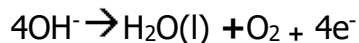


As you have seen, water also produces ions. So a dilute solution of the acid contains H^+ ions from both water and acid, OH^- ions from the water and SO_4 ions from the acid. The solution can be electrolysed using appropriate apparatus.

At the cathode: this time there are no metal ions to compete with. So hydrogen gas is formed and bubbles off:



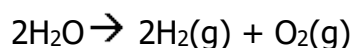
At the anode: the OH^- and SO_4 ions compete to give up electrons. As you would expect from rule 4, the OH^- ions win:



The oxygen bubbles off. The SO_4^{-2} ions are left behind in solution.

The overall result is that the water decomposes rather than the acid:

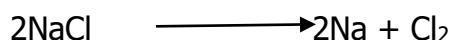
Water $\rightarrow$ hydrogen + oxygen



Uses of electrolysis

Extracting reactive metals

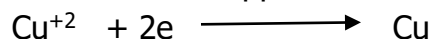
Reactive metals like to exist as compounds so it is difficult to extract them from their compounds. Electrolysis delivers the energy that is needed for example it is used to extract sodium from molten sodium chloride



Purifying metals

When a solution of copper sulphate solution is electrolysed using carbon or platinum electrodes copper is obtained at the cathode and oxygen at the anode. But when copper electrodes are used there is a different result

At the cathode: copper ions become atoms

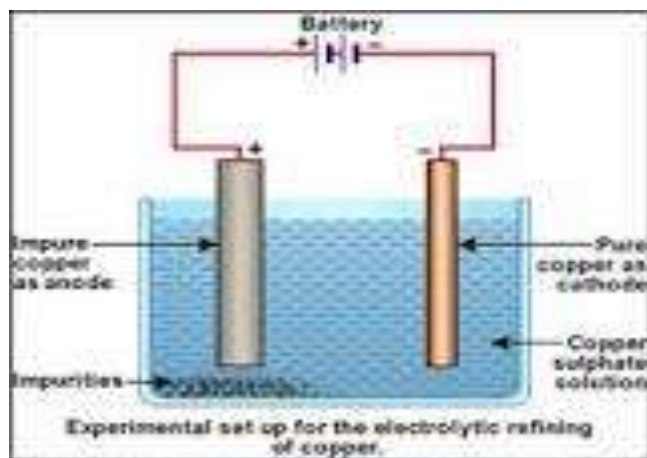


The copper atoms cling to the cathode

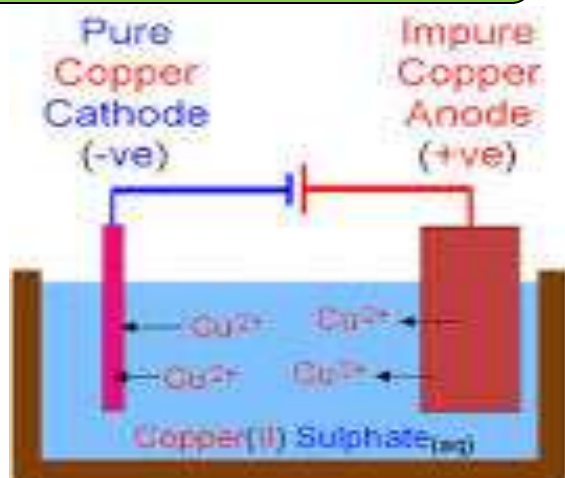
At the anode: the copper anode dissolves forming copper ions



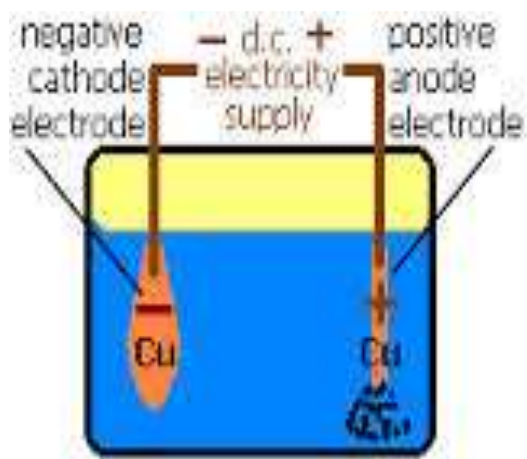
So the anode wears away while the cathode grows thicker. This process is used in industry to purify copper



The anode is made of impure copper and the cathode is pure copper the electrolyte is solution of copper sulphate solution



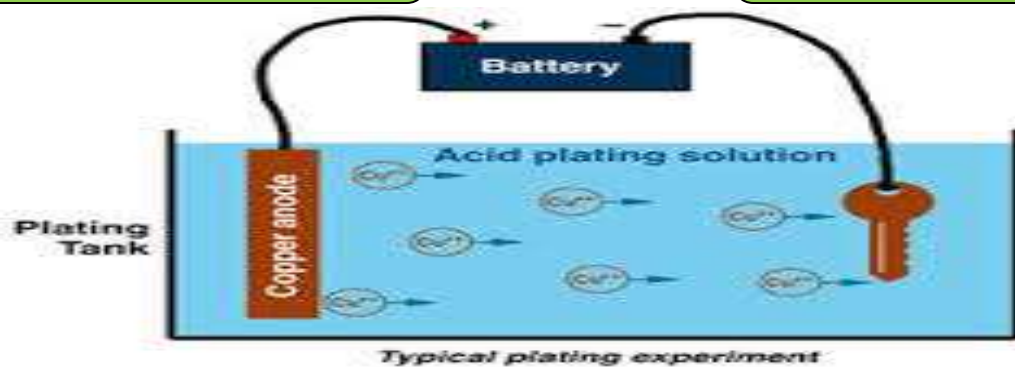
The anode dissolves and the impurities drop to the floor of the cell. A layer of pure copper builds up on the cathode.



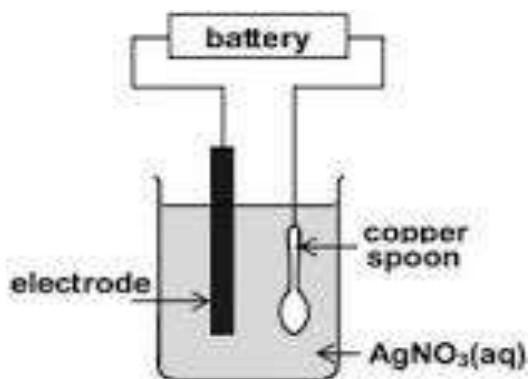
eventually the cathode is removed. The slime on the cell floor is checked for silver and other precious metals

Electroplating

Electrolysis is also used to coat one metal with another to make it look better or to prevent it corrosion. For example it is used to coat steel car bumpers with chromium or steel cans with tin or cheap metal jewellery with silver. This process is called electroplating.



Typical plating experiment

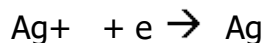


at the anode: the silver dissolves forming ions in solution.



At the cathode :

Silver ions receive electrons forming a coat of silver on the spoon.



When the layer of silver is thick enough the spoon is removed

In general to electroplate any object with metal X the set up is

Cathode: object to be electroplated

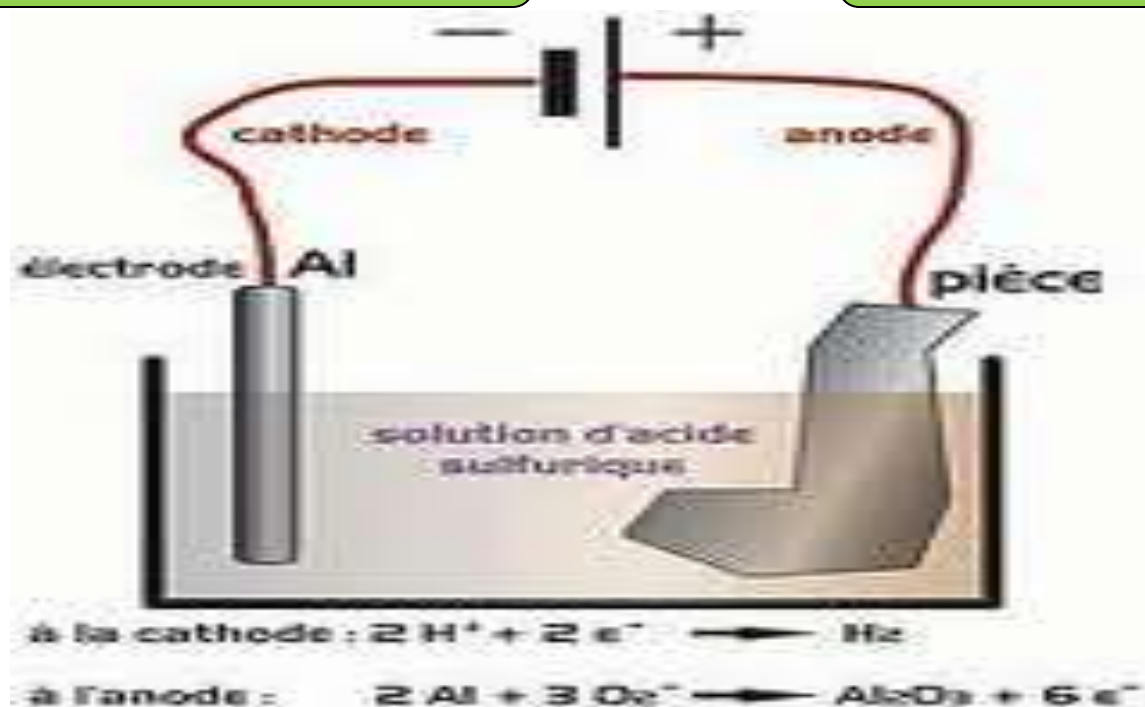
Anode: metal X

Electrolyte: solution of a compound of X

Anodizing aluminium

Although aluminium is quite reactive corrosion in air is not a problem this is because a thin layer of aluminium oxide quickly forms and acts as a barrier to oxygen the layer can be made thicker to give even more protection by a process called ANODIZING.

The aluminium is used as the anode of the cell and dilute sulphuric acid as the electrolyte.



Calculations on electrolysis

Faraday's law

The mass of a substance formed at each electrode during electrolysis is governed by two factors:

- The size of the current.
- How long the current flows.

This relationship was stated by faraday and is known as faraday's laws of electrolysis.

Faraday's law of electrolysis states:

The mass of a substance produced at an electrode during electrolysis is directly proportional to the size of the current and how long it flows.

Coulombs

The coulomb is a measure of quantity of electricity. 1 coulomb is the quantity of electricity which passes if 1 ampere(amp) flows for 1 second.

Number of coulombs= current (in amps) X time (in sec).

Faraday's constant

A flow of electricity is a flow of electrons. 1 faraday is the quantity of electricity which represents 1 mole of electricity passing a particular point in the circuit.

1 faraday= 96500 coulombs

Example 1

Find the quantity of electricity used up when 10 amperes of electricity passed through an electrolyte for 1 hour and 30 minutes.

Solution

Given: $I = 10 \text{ A}$

Required Q

$T = 1\text{h and } 30 \text{ min}$

$Q = It$

$= 60\text{min} \times 30\text{min}$

$Q = (10\text{A}) (5400\text{s})$

$= 90\text{min} \times 60 \text{ s/min}$

$Q = 5400\text{As or } 5400\text{C.}$

$= 5400\text{s}$

Example 2

What time is required to use 1930C of electricity when the current delivery is 20A?

Solution

Given: $Q = 1930\text{C}$

required t

$I = 20\text{A}$

$t = Q/A$

$t = 1930\text{C}/20\text{A}$

$t = 96.5\text{s}$

Example 3

A solution of copper (II) sulphate is electrolysed using copper electrodes. How much copper will be deposited by a current of 2 amps flowing for 20 minutes?

Solution**Step 1: calculate the charge of the current:**

Charge (in coulombs) = time (seconds) X current (amps)

$$Q = It = (2A) (20\text{min} \times 60\text{sec}/\text{min}) = 2400\text{AS} = \mathbf{2400C}$$

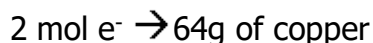
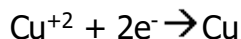
Step 2: convert the charge into moles of electrons:

$$96500C = 1 \text{ mole } e^-$$

$$2400C = X$$

$$X = 2400C \times 1 \text{ mole}^- / 96500C$$

$$X = 0.025 \text{ moles of electrons}$$

Step 3: Now use the equation for the reaction at the cathode to find how much copper is deposited:

$$X = (0.025 \text{ mole}^-)(64\text{g}) / 2\text{mole} \mathbf{X = 0.8 g}$$

Alternatively this problem can be solved by the equation:

$$\mathbf{M = \frac{I \times T \times M}{F \times z}}$$

F x z

I = CURRENT M = MASS

T = TIME F = FARADAY

Z = MUNBER OF MOLES OF ELECTRON

$$x = \frac{2 \times 20 \times 60}{96500 \times 2} = \mathbf{0.8 \text{ grams}}$$

Example 4

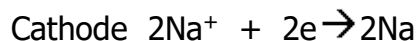
1 kg of sodium is produced by electrolysing sodium chloride. What volume of chlorine is produced at the same time?

Solution**Step 1. Calculate the number moles of sodium**

Mole=mass/molar mass

$$= 1000/23$$

$$\text{Moles} = 43.48$$

Step 2. Calculate the number of moles of chlorine by using half equations

1 mole of Cl_2 give up 2 moles of electrons which is enough to produce 2 moles of sodium atoms. But 43.48 moles of sodium is produced so $43.48/2$ moles or 21.74 moles of chlorine is produced

Step 3:Convert volume into mole

1 mole of any molecules has a volume of 24dm³ at rtp

So the volume of chlorine obtained is $24 \times 21.74 = 521.76 \text{ dm}^3$

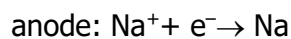
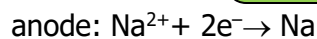
$$\text{Unknown volume} = \frac{\text{mass} \times 24\text{dm}^3 / \text{mol}}{\text{Molar mass} \times \text{number of electrons}}$$

Chapter Review Questions

Part one circle the correct answer

- Decomposition caused by electricity is called
 - Thermal decomposition
 - Electrolyte
 - electrolysis
 - none
- In an **electrolytic** cell the electrode at which the electrons enter the solution is called the _____ ; the chemical change that occurs at this electrode is called _____.
 - Anode, oxidation
 - Anode, reduction
 - Cathode, oxidation
 - Cathode, reduction
 - Cannot tell unless we know the species being oxidized and reduced
- The half-reaction that occurs at the anode during the electrolysis of molten sodium bromide is:
 - $2 \text{Br}^- \rightarrow \text{Br}_2 + 2 \text{e}^-$
 - $\text{Br}_2 + 2 \text{e}^- \rightarrow 2 \text{Br}^-$
 - $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$
 - $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$
 - $2 \text{H}_2\text{O} + 2 \text{e}^- \rightarrow 2 \text{OH}^- + \text{H}_2$
- Which **one** of the following lists shows the ions present in copper (II) sulfate solution?

A	Cu^{2+}	SO_4^{2-}		
B	Cu^+	SO_4^{2-}		
C	Cu^+	SO_4^{2-}	H^+	OH^-
D	Cu^{2+}	SO_4^{2-}	H^+	OH^-
- During the electrolysis of molten sodium chloride, the correct electrode equations are
 - cathode: $\text{Na}^{2+} + 2\text{e}^- \rightarrow \text{Na}$; anode: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$
 - cathode: $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$; anode: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$



6. During electrolysis:

A. Reduction occurs at the cathode; oxidation occurs at the anode.

B, Reduction occurs at the anode; oxidation occurs at the cathode.

C. Reduction occurs at both electrodes.

D. Oxidation occurs at both electrodes.

7. Which **one** of the following statements is **wrong**?

A. The negative electrode in an electrolysis circuit is called the cathode.

B. Electrolysis is a chemical change caused by passing an electric current through

a compound which is either molten or in solution.

C. During electrolysis, positive ions move towards the anode.

D. Covalent compounds don't undergo electrolysis.

8. The only solids that conducts electricity are

A. non-metals and graphite C. metals and diamond

B. Metals and graphite D. none of them

9. The oxidation number of carbon in **CaCO₃** is

A. + 4 C. 0

B. - 4 D. None

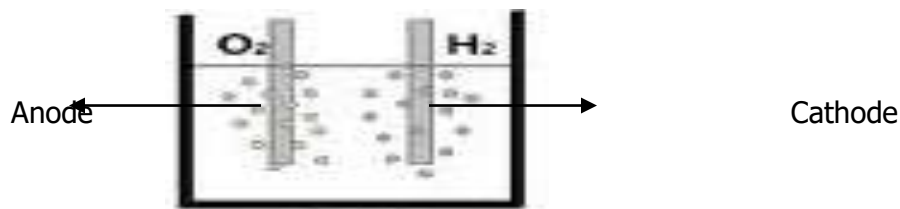
10. The oxidation number of **Mg⁺²** is

A. 0 C. +2

B. -2 D. None

Part two answer all the following questions

1. The Diagram Below shows the Electrolysis of Lithium Chloride Solution Using Graphite Electrodes.



A. List The Four Ions Present in the solution?

B. Name the product formed at each Electrode

i. Anode : _____

ii. Cathode : _____

C. Write down the half Equation for the reactions at each Electrode

Anode: _____

Cathode: _____

2. Define oxidation and reduction in both terms of electrons and oxygen

Oxidation I

ii.....

Reduction I

ii.....

3. What are the uses of electrolysis?

A.....B..... C.....

D.....

4. Define these words

• Electrolysis

• Anode.....

- Cathode.....
- Cation.....
- Anion.....
- Electrolyte.....
- Electroplating.....
- Anodizingaluminum
.....

5. This questions is about electrolysis of lead bromide

I. Why lead bromide is heated until it is molten?

.....
.....
.....

II. What are equations at the electrodes called?

.....

III. Do lead ions go to the cathode or anode? Explain the reason?

.....

IV. Do lead ions gain or lose electron?

V. Is the reaction at the cathode is called reduction or oxidation? Why?

.....

.....

VI. Write the balanced chemical equation for the overall reaction?

.....

6. 200 gram of potassium is produced by electrolyzing potassium fluoride. What volume of fluorine is produced?

.....
.....
.....

7. What two substances depends on the amount of substance formed during electrolysis?

A.....B.....

8. How much charge is needed to reduce 1 mole of Al^{+3} ions?

.....

.....

.....

9. An electric current of 100 Ampere is passed through a molten sodium chloride for 5 hours. Calculate the volume of chlorine produced at the same time?

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11. This question is about the electrolysis of silver nitrate solution

A. What ions are present in the solution?

B. Write the half equations for the each electrode?

At the cathode

.....

.....

At the anode

.....

.....

12. Write an equation for the reaction you expect at each electrode when you electrolyse a solution of A. sodium nitrate solution B. potassium iodide solution

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13. Aluminum is extracted by electrolysis of molten aluminum oxide

A. Write the equation for the reaction at the cathode

.....

B. How many moles of electrons are needed to obtain one mole of aluminum

.....

C. What mass of aluminum will be obtained if a current of 25000 A flows for 24 hours?

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14. This question is about the electrolysis of an aqueous solution of lithium chloride

A. Write down the names and symbols of all the ions present in the solution

.....
.....

B. What will be formed at the cathode

.....

C. What will be formed at the anode

.....
D. Write an equation for the reaction at each electrode

Cathode.....Anode.....

.....
E. Name two other electrolytes that will give up the same electrolysis product as this one?

.....
.....
.....
.....

CHAPTER 4: NON-METALS

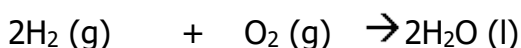
Hydrogen

Hydrogen is the lightest of all the elements. It is so light that there is none in the air. It has escaped into the outer atmosphere. But overall, it is the most common element in the universe.

The properties of hydrogen

1. It is the lightest of all gases. It is about 20 times lighter than air.
2. It is colorless and has no smell.
3. It is almost insoluble in water.
4. It combines with oxygen to form water. A mixture of the two gases will explode when lit. So take care. The reaction is:

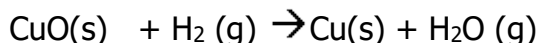
Hydrogen + Oxygen $\rightarrow$ water



This reaction gives out so much energy that it is used to fuel space rockets where the reaction takes place in **fuel cells**.

5. Hydrogen acts as a **reducing agent**, by removing oxygen. For example, copper (II) oxide is reduced to copper, by heating it in a stream of hydrogen. The hydrogen is oxidized to water:

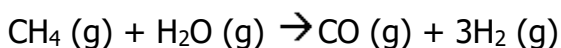
Copper(II)oxide + Hydrogen $\rightarrow$ copper + steam



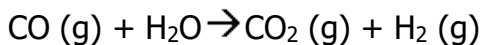
How hydrogen is made in industry

It is made from North Sea gas, which is mainly methane.

The North Sea gas is mixed with steam and passed over a catalyst:



This is a reversible reaction. A high temperature and pressure are needed to give a good yield. Then the carbon monoxide is converted to carbon dioxide in another reversible reaction.



The carbon dioxide is removed by scrubbing the gases with an alkali. This leaves the hydrogen.

Uses of Hydrogen

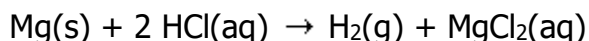
1. It is used to 'harden' vegetable oils to make margarine.
2. It is used as fuel for space rockets.

Hydrogen may become an important fuel for cars and homes in the future, as we run out of oil and gas. It has two big advantages:

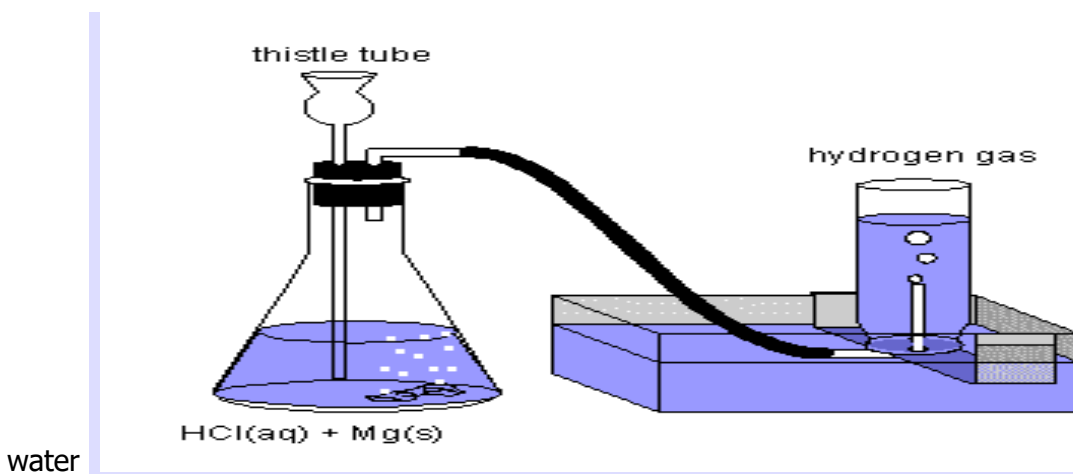
- Its reaction with oxygen produces just water. **No pollution!**
- It is a '**renewable**' **resource**. It can be made by electrolyzing acidified water. As cheaper sources of electricity are developed, this may become an attractive option.

How can pure hydrogen gas be prepared and detected?

Hydrogen can be made by reaction of any **active metal** like magnesium or zinc with a strong acid such as H_2SO_4 or HCl , for example,



Since hydrogen gas is almost insoluble in water, it can be collected by displacement of



Hydrogen gas is recognized by the 'pop' when it burns.

The 'pop' is the sound of a small explosion.

Hydrogen gas is highly flammable!

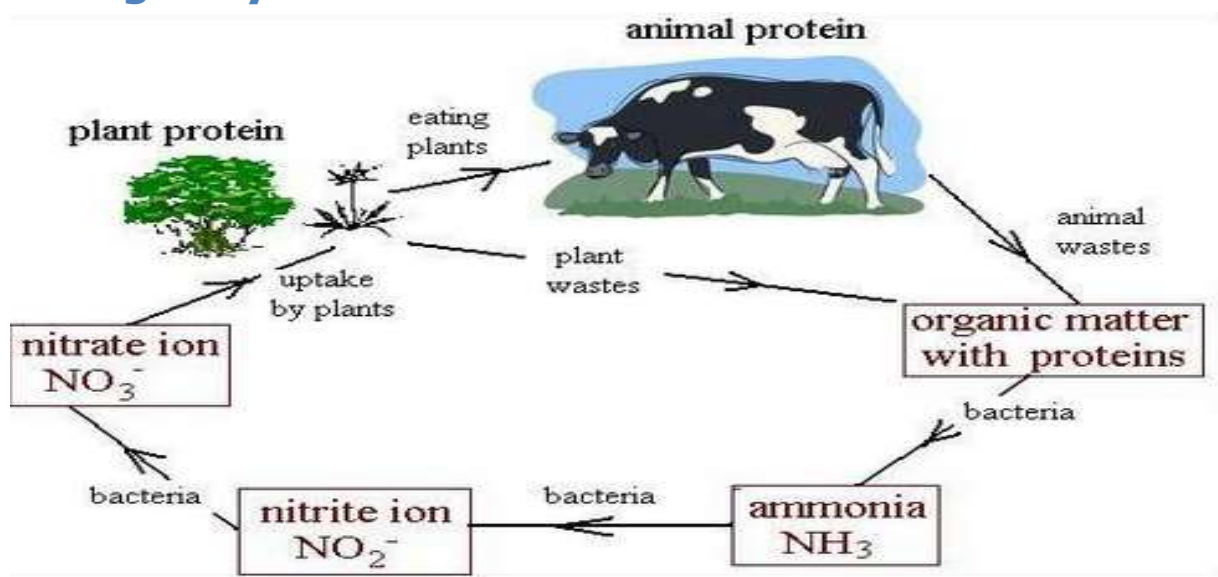
Hydrides

They are compounds formed by a metal and hydrogen, in which the hydrogen has an oxidation number -1 . Their name is formed by two words. The first word is the name of the metal the second word is the root of plus the suffix **-ide**. The most common ones are **CaH_2 , NaH , AlH_3 , and MgH_2**

Nitrogen

Nitrogen is a colorless, odorless, un-reactive gas—and we can't survive without it. It continually circulates between the air, the soil, and living things in a set of processes called the **nitrogen cycle**.

Nitrogen cycle



Simple Nitrogen Cycle Diagram

1. The conversion of nitrogen from air to nitrates in the soil:

- In the heat of a lightening flash, nitrogen and oxygen react to form nitrogen oxides. These dissolve to form acid rain, which reacts with compounds in soil to make nitrates.
- Some bacteria can also convert nitrogen to nitrates. They are **nitrogen-fixing**. They live in soil, or swellings called root nodules on the roots of pea, bean, and clover plants.
- Humans turn nitrogen into fertilizers in factories. These are compounds such as ammonium phosphate. Farmers spread them on soil.

2. **From nitrates to proteins:** plants take in nitrates through their roots and convert them to proteins (but pea, bean, and clover plants don't need to do this since they can fix nitrogen from the air directly.)
3. **From plants to animals:** animals eat, and in turn are eaten by other animals. The proteins are broken down into amino acids during digestion, built up again to form new proteins. In this way proteins are passed along the food chain. You need them to build up your body tissues.

4. **Excretion and decay:** animals excrete ammonium compounds. Bacteria also feed on the remains of dead plants and animals, producing ammonium compounds. These are converted into nitrates by **nitrifying bacteria**.
5. **From nitrates back to nitrogen:** bacteria called **denitrifying bacteria** complete the cycle. They live in heavy, wet soils. They break down nitrates, releasing back into the air.

The properties of nitrogen

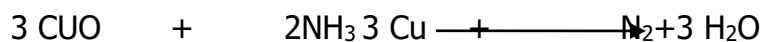
1. It is a colorless gas with no smell.
2. It is only slightly soluble in water.
3. It is very unreactive compared with oxygen. The strength of the nitrogen-nitrogen triple bond makes the N_2 molecule very un-reactive
4. However, it will react hydrogen to form ammonia, with the help of high pressure, moderate temperature, and a catalyst. The reaction is reversible:

$$N_2 (g) + 3H_2 (g) \rightarrow 2NH_3 (g)$$
 This reaction is the first step in making nitrogen fertilizers.
5. It will also combine with oxygen at high temperatures to form oxides of nitrogen. This happens naturally during thunderstorms. It also happens inside car engines and power stations furnaces. The results of these are air pollution and acid rain.

How nitrogen is made in the lab

Nitrogen gas can be made in the lab by the following reactions

Copper (ii) oxide + ammonia $\longrightarrow$ copper + nitrogen + water



Uses of nitrogen

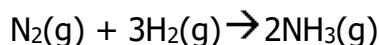
1. Manufacture ammonia and nitric acid.
2. Manufacture of fertilizers such as ammonium nitrate and ammonium sulphate.
3. To flush out food for packaging and keeping food fresh (since it is inert)

4. Quick—freezing food, freeze liquid in damaged pipes, and shrink—fit machine parts(since liquid nitrogen is very cold).

Ammonia

Ammonia is a gas with the formula **NH₃**.

In industry: it is made in industry by reacting nitrogen with hydrogen.



In the lab: it is made in the lab by heating any ammonium compound with a strong base. The base drives out or displaces ammonia from the compound.

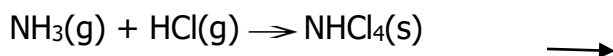
For example: Ammonium chloride + calcium hydroxide → calcium chloride + water + ammonia



This reaction can be used as a test to identify an ammonium compound. If a compound gives off ammonia when heated with a strong base, it must be an ammonium compound.

The properties of ammonia

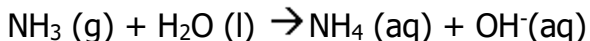
1. It is a colorless gas with a strong choking smell.
2. It is less dense than air.
3. It is easily liquified by cooling to -33°C or by compressing. So it is easy to transport in tanks.
4. It reacts with hydrogen chloride gas to form a white smoke. The smoke consists of tiny particles of solid ammonium chloride:



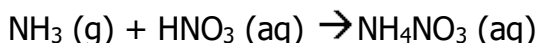
The reaction can be used to test whether a gas is ammonia.

5. It is very soluble in water.

6. The solution turns red litmus blue—it is alkaline. That means it contains hydroxide ions. Some of the ammonia has reacted with water to form ammonium ions and hydroxide ions:



7. Since ammonia solution is alkaline, it reacts with acids to form salts. For example, with nitric acid it forms ammonium nitrate. Ammonium nitrate is important fertilizer.



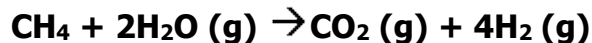
8. It can be used to test for copper (ii) compounds. If a solution contains copper (ii) compounds it will react with ammonia solution to give blue precipitate of copper(ii)hydroxide. When more ammonia solution is added the precipitate will dissolve giving deep blue solution. The blue solution proves that copper (ii) compound is present.

Making ammonia in industry

In industry, ammonia is made from nitrogen and hydrogen in a process called the **Haber process**.

First, you must get the two gases:

- Hydrogen is made from methane (North Sea gas) and steam.



- Nitrogen is obtained by burning hydrogen in air to remove oxygen. When the water vapor that forms in the reaction is cooled and condensed, the remaining gas is mainly nitrogen.

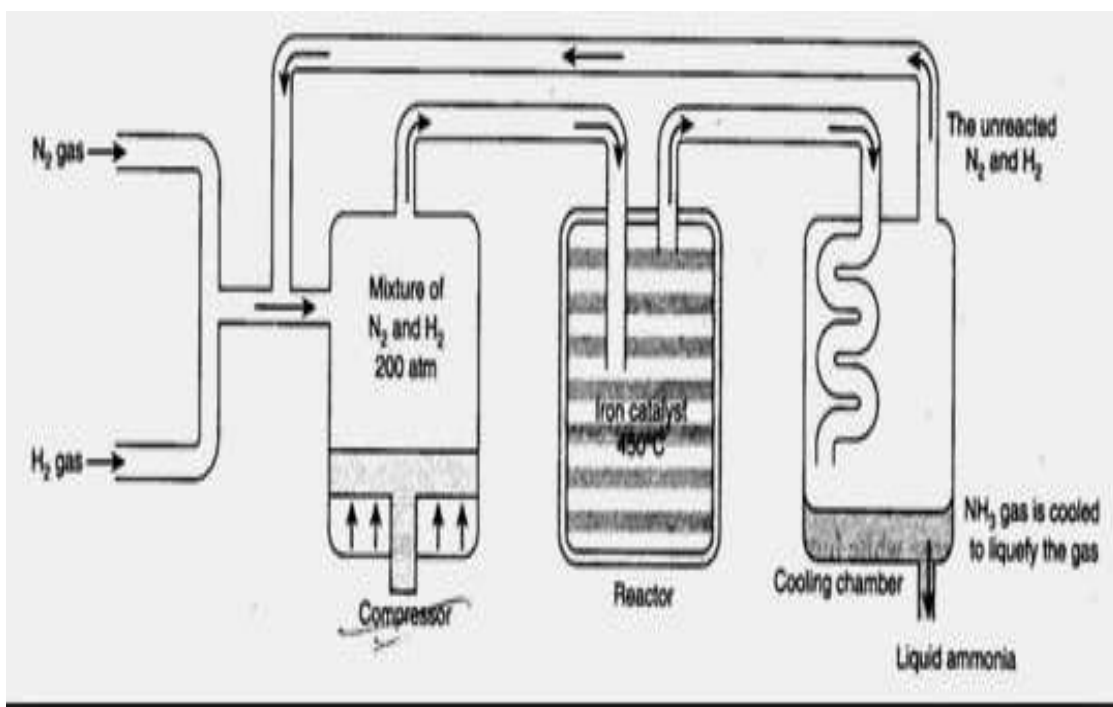
The reaction between hydrogen and nitrogen is reversible. At a certain point, ammonia breaks down at the same rate as it forms. So conditions must be chosen to give the highest possible yield for the lowest cost. These conditions are:

- ✓ High pressure.
- ✓ Ammonia removed, so that more will form.

- ✓ Moderate heat and a catalyst, to help the reaction reach equilibrium quickly.

These are the steps in Haber process:

1. The two gases are mixed, and the mixture is scrubbed to get rid of impurities.
2. It is compressed to a pressure of about 200 atmospheres.
3. Then it goes to the converter. This is a round containing bed of hot iron. The iron is a catalyst for the reaction:
$$\text{N}_2 (\text{g}) + 3\text{H}_2 (\text{g}) \rightarrow 3\text{NH}_3 (\text{g})$$
4. A mixture of all three gases leaves the converter. It is cooled until the ammonia condenses. Then the hydrogen and nitrogen are pumped back to the converter for another chance to react.
5. The ammonia is run into tanks and stored as a liquid under pressure.



Uses of ammonia

1. It is used to make fertilizers such as ammonium nitrate and ammonium sulphate.

2. It is used to make household cleaners, dyes, explosives, and nylon.
3. A lot is used to make nitric acid.

Fertilizers

Plants need carbon dioxide light and water for photosynthesis but they also need

- Nitrogen to make proteins for growth and root development
- Potassium to promote growth
- Phosphorus to help leaves develop

They also need smaller amounts of many other elements including calcium iron and sulphur they obtain these things from compounds in soil. But when the same crops are grown in the same soil year after year the supplies of compounds in the soil get exhausted and crops suffer. That is why fertilizers come from.

Fertilizers are added to the soil to replace the elements used up by plants to help crops grow. Animal manure is natural or organic fertilizer.

Artificial fertilizers are compounds such as

- Ammonium nitrate NH_4NO_3
- Ammonium phosphate $(\text{NH}_4)_3\text{PO}_3$
- Potassium chloride KCl

There are two main types of fertilizers

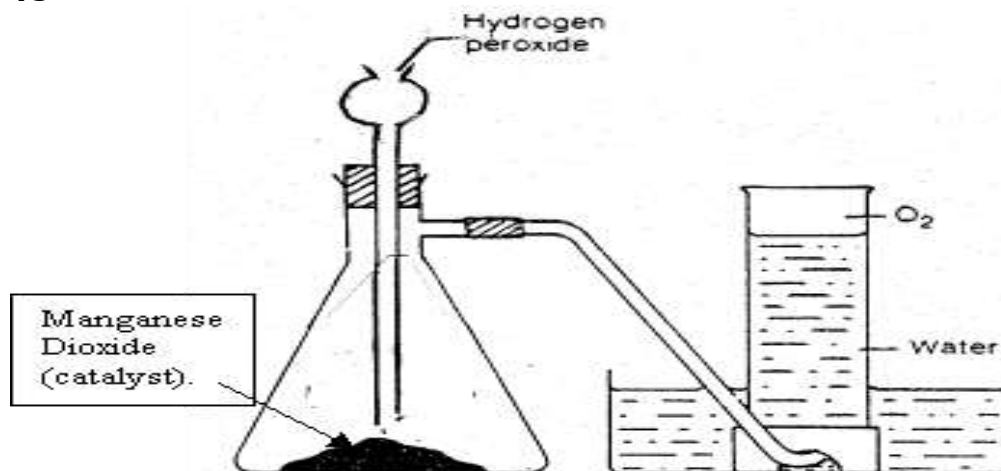
- Straight N fertilizers for farmers who want only nitrogen
- NPK fertilizer which are usually a mixture of Ammonium nitrate Ammonium phosphate and Potassium chloride

Problems of fertilizer

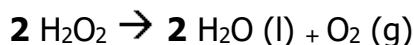
- Nitrates promote the growth of tiny water plant called algae. Some algae are poisonous to fish and humans
- Excess nitrate in the river may get into drinking water and increases the risk of blue baby syndrome

Oxygen

Making oxygen in the lab:



Hydrogen peroxide is a colorless liquid, with the formula H_2O_2 . It decomposes to water and oxygen:

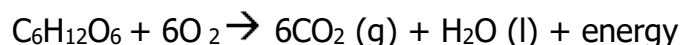


This reaction can be used to prepare oxygen in the lab. But it is very slow, so black manganese (V) oxide is added as a catalyst.

The properties of oxygen

1. It is a clear colorless gas with no smell.
2. It is only slightly soluble in water.
3. It is very reactive. It reacts with a great many substances to produce oxides, and the reaction usually gives out a lot of energy. For us, its two most important reactions are **respiration** and the **combustion of fuels**. These are very similar.

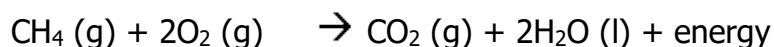
Respiration: this is the process that keeps us alive. During respiration, oxygen reacts with glucose in our bodies. The reaction produces carbon dioxide, water, and the energy we need:



The carbon dioxide and water from respiration pass from the cells back into the blood. The blood carries them to the lungs, and we breathe them out.

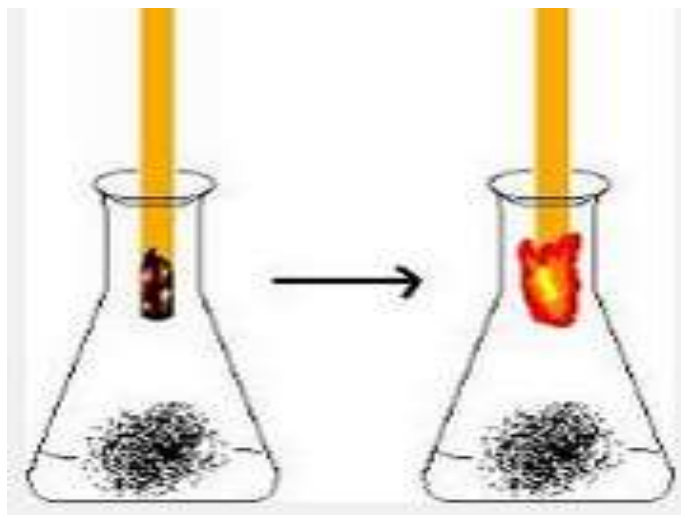
Combustion of fuels: fuels are substances we burn to get energy—usually in the form of heat. The burning needs oxygen. For example, methane burns like these:

Methane + oxygen $\rightarrow$ carbon dioxide + water vapor + energy



Test for oxygen: things burn much faster in pure oxygen than in air. The reason is that the oxygen in air is diluted by nitrogen and other gases. This gives a way to test a gas, to see if it is oxygen.

1. A wooden splint is lit. Then the flame is blown out. The splint keeps on glowing, because the wood is reacting with oxygen.
2. The glowing splint is plunged into the unknown gas.
3. If the gas is oxygen, the splint immediately bursts into flame.





Oxides

Oxygen directly combines with all other elements (except noble gases and unreactive metals like gold) to form compounds known as oxides.

According to their chemical properties, oxides are categorized into five groups. These are:

1. Acidic oxides.
2. Basic oxides.
3. Amphoteric oxides.
4. Neutral oxides.
5. Peroxides.

1. Acidic oxides.

Non—metals react with oxygen to form acidic oxides. These acidic oxides dissolve in water to form acids.

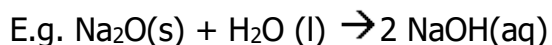
Acidic oxide + water $\rightarrow$ acid

e.g. $\text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{l}) \rightarrow \text{H}_2\text{CO}_3 (\text{aq})$

2. Basic oxides.

Metals react with oxygen to form basic oxides. These basic oxides dissolve in water to form bases.

Basic oxide + water $\rightarrow$ base



3. Amphoteric oxides.

Some oxides of metals show acidic properties as well as basic properties and they are called amphoteric oxides.

They act as bases when they react with acids, and they also behave as acids when they react with bases.

An example of amphoteric oxide is Zinc oxide.

Amphoteric oxides are oxides that form salts when they react with both acids and bases.

4. Neutral oxides

These as their name suggests give neither acid nor alkaline (basic) solution when dissolved in water. They are not capable of neutralizing acid or base. Examples are pure water (H_2O), carbon monoxide(CO), Nitrogen monoxide(NO) and dinitrogen oxide(N_2O).

5. Peroxides

Peroxides have more oxygen atoms per molecule than the normal oxides. In peroxides, the oxygen is in the -1 oxidation state. Some examples of peroxides are hydrogen peroxide (H_2O_2), sodium peroxide (Na_2O_2) and barium peroxide (BaO_2).

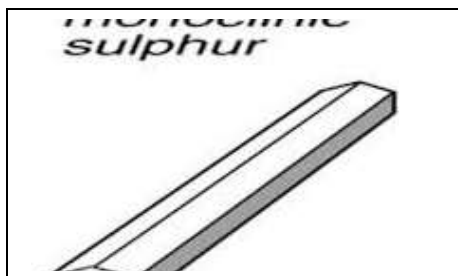
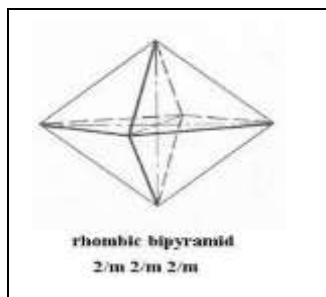
Sulphur.

Sulphur is the sixteen's most abundant element in the earth's crust

- Sulphur occurs as the element in huge underground sulphur beds.
- It is also found combined with metals in many metal ores. For example, in pyrite(FeS_2) and galena(PbS).
- It is found in natural gas as hydrogen sulphide(H_2S), and in crude oil as organic sulphur compounds.

Properties of sulphur

1. It is brittle yellow solid.
2. It is made up of crown-shaped molecules each with eight atoms.
3. It can take two different crystalline forms. Different forms existing at the same state are called ALLOTROPES. Rhombic and monoclinic are allotrope of



4. Because it is molecular, it has quite a low melting point (115°C). It melts easily in a Bunsen flame.
5. Like other non-metals, it doesn't conduct electricity.
6. Like most non-metals, it is insoluble in water.
7. It reacts with metals to form sulphides. For example, it reacts with iron to form iron(II)sulphide:

$$\text{Iron} + \text{sulphur} \rightarrow \text{iron sulphide}$$

$$\text{Fe(s)} + \text{S(s)} \rightarrow \text{FeS(s)}$$
8. It burns in oxygen to form sulphur dioxide:

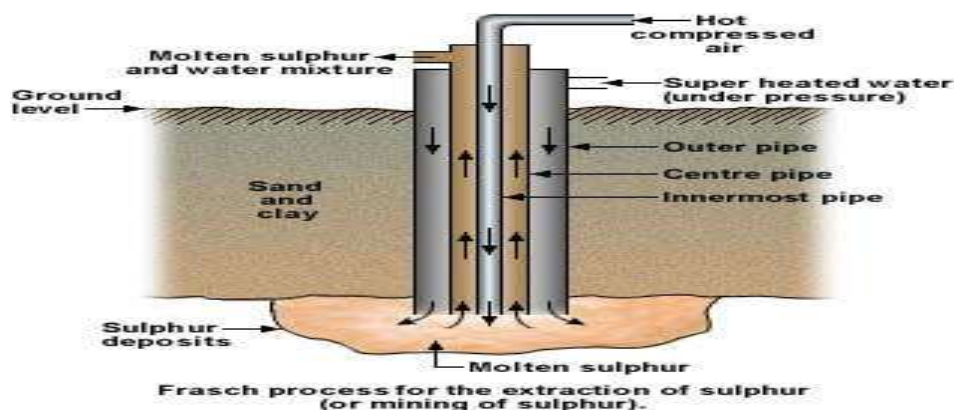
$$\text{Sulphur} + \text{oxygen} \rightarrow \text{sulphurdioxide}$$

$$\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)}$$

Extraction of Sulphur

The process of extracting sulphur from underground deposits is called **Frasch process**.

In the Frasch process, superheated water is pumped into the sulfur deposit; the sulfur melts and is extracted. The Frasch process is able to produce high purity sulfur.



In the Frasch process, three concentric tubes are introduced into the sulfur deposit. Superheated water (165°C , 2.5-3 MPa) is injected into the deposit via the outermost tube. Sulfur (m.p. 115°C) melts and flows into the middle tube. Water pressure alone is unable to force the sulfur into the surface due to the molten sulfur's greater density, so hot air is introduced via the innermost tube to froth the sulfur, making it less dense, and pushing it to the surface.

The Frasch process can be used for deposits 50–800 meters deep. 3-38 cubic meters of superheated water are required to produce every tonne of sulfur, and the associated energy cost is significant.

Uses of sulphur

1. Around 90% of sulphur is used to make sulphuric acid.
2. Some is added to rubber, for example for car tyres, to toughen it. This is called vulcanization of rubber.

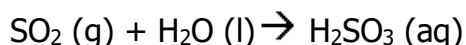
3. Some is used to make drugs, pesticides, matches, and paper.
4. Some is added to cement to make sulphur concrete. Unlike ordinary concrete, this is not attacked by acid. So it is used for walls and floors in plants where acid is used.

Sulphur dioxide

Sulphur dioxide is formed when sulphur burns in air. Its formula is **SO₂**

Properties of sulphur dioxide

1. It is a colorless gas, with a strong, choking smell.
2. It is heavier than air.
3. It is soluble in water. The solution is acidic because the gas reacts with water to form sulphurous acid, H₂SO₃.



Sulphurdioxide is, therefore, an acidic oxide. The acid easily decomposes again to sulphur dioxide and water.

4. It acts as bleach when it is damp or in solution. Some colored things lose color when they lose oxygen—that is, when they are reduced. Sulphur dioxide bleaches them by reducing them.
5. When it escapes into the air from engine exhaust and factory chimneys, it causes air pollution. It attacks the breathing system in humans and other animals. It dissolves in rain to give acid rain. Acid rain damages buildings, metal works, and plants.

Uses of sulphur dioxide

1. It is used to bleach wool, silk, and wood pulp for making paper.
2. It is used as a sterilizing agent in making soft drinks and jam and in drying fruit. It stops the growth of bacteria and moulds.

3. Its main use is in the manufacture of sulphuric acid.

Sulphuric acid, H_2SO_4

Sulphuric acid is the most manufactured each year than any other chemical.

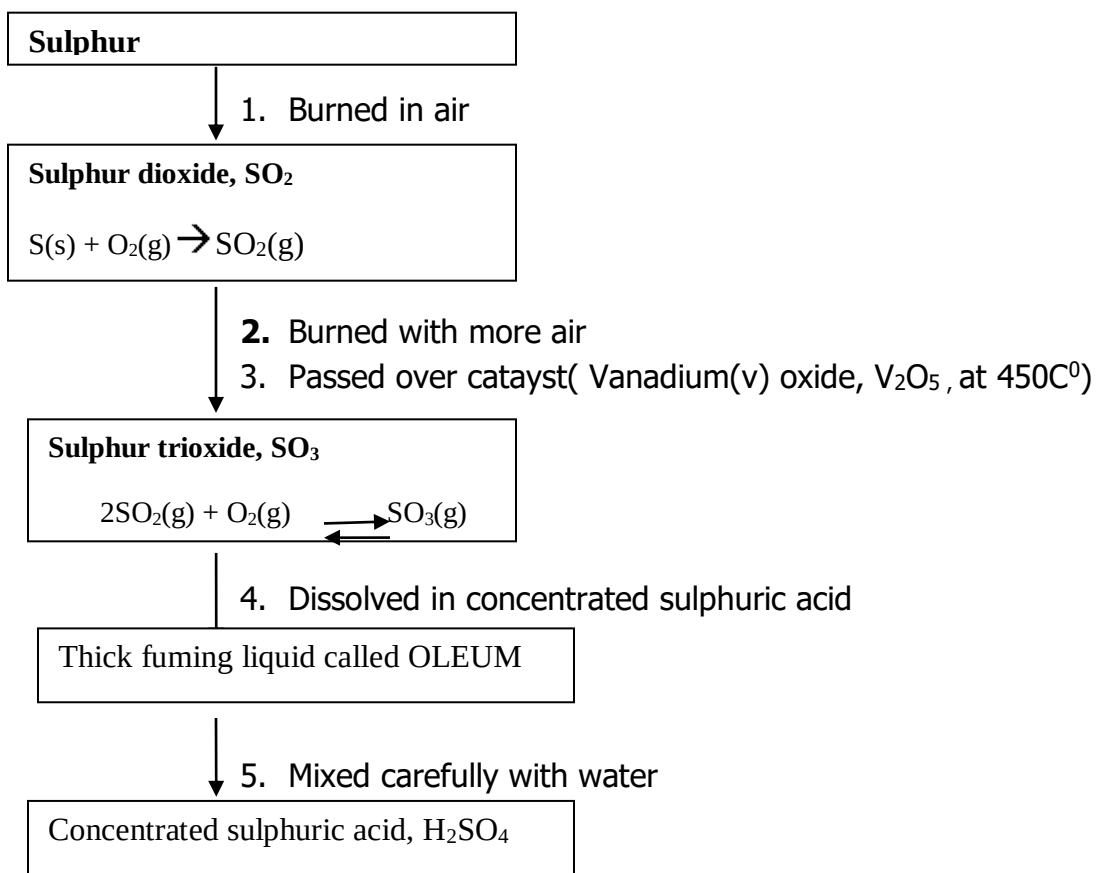
How sulphuric acid is made

The process of manufacturing sulphuric acid in industry is called **Contact process**.

The raw materials for the contact process are:

- Sulphur or sulphur dioxide
- Air
- Water

Starting with sulphur, the steps in the contact process are:



Uses of sulphuric acid

Nearly every industry uses some sulphuric acid: it is the cheapest acid to buy.

The main uses of sulphuric acid are in the manufacture of:

1. Fertilizers.
 2. Paints, pigments, and dyestuffs.
 3. Fibers and plastics.
 4. Soaps and detergents.
- It is also used as the acid in car batteries.

The properties of sulphuric acid

When concentrated, sulphuric acid has these properties:

1. It is a colorless oily liquid.
2. It is a dehydrating agent. It can remove water. It will dehydrate sugar, paper and wood. These are all made of carbon, hydrogen, and oxygen. The acid removes the hydrogen and oxygen as water, leaving carbon behind. For example:



It will also dehydrate flesh in the same way, so it's very dangerous.

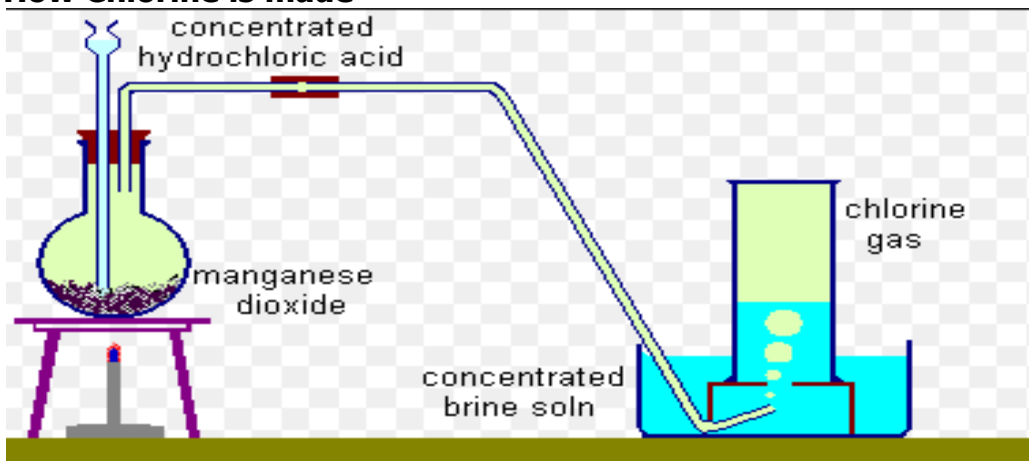
When it is dilute, it has typical acid properties including the following:

1. It turns blue litmus red.
2. It reacts with metals to give hydrogen, and salts called sulphates.
3. It will react with metal oxides and hydroxides to give sulphates and water.
4. It reacts with carbonates to give sulphates, water and carbon dioxide

Chlorine

Chlorine is very reactive so it's never found as the free element in earth's crust. It occurs mainly as sodium chloride or rock salt.

How Chlorine is made



In the lab: chlorine is made in the lab by oxidizing concentrated hydrochloric acid. The oxidizing agent is potassium manganate(IV). the apparatus is shown above.

The acid is dripped onto the potassium manganate(IV) while the mixture is heated gently.

Since the chlorine is denser than air, it is collected by **upward displacement of air**.

In industry: chlorine is made in industry by the electrolysis of molten sodium chloride or brine.

The properties of chlorine

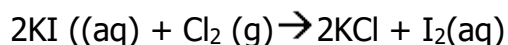
1. It is a greenish-yellow poisonous gas, with a choking smell.
2. It is heavier than air

3. It is soluble in water. The solution is called **chlorine water**. It is acidic because chlorine reacts with water to form two acids:
$$\text{Cl}_2 (\text{g}) + \text{H}_2\text{O} (\text{l}) \rightarrow \text{HCl}(\text{aq}) + \text{HOCl}(\text{aq})$$

The hypochlorous acid slowly decomposes, giving off oxygen:
$$2\text{HOCl} (\text{aq}) \rightarrow 2\text{HCl}(\text{aq}) + \text{O}_2(\text{g}) .$$
4. Chlorine water acts as bleach. This is because the hypochlorous acid can lose its oxygen to other substances—it oxidizes them. Some colored substances lose their color when oxidized.
5. Like other bleaches, chlorine water also acts as a sterilizing agent. It kills bacteria and other germs.
6. Hydrogen burns in chlorine to form hydrogen chloride. The reaction can be explosive:
$$\text{H}_2 (\text{g}) + \text{Cl}_2 (\text{g}) \rightarrow 2\text{HCl} (\text{g})$$
7. Chlorine also combines with most metals, forming metal chlorides. For example, when it is passed over heated the aluminum glows white hot and turns into aluminum chloride:
$$2\text{Al}(\text{s}) + 3\text{Cl}_2(\text{g}) \rightarrow 2\text{AlCl}_3(\text{s})$$
8. Chlorine is more reactive than halogens below it in group 7. So it will displace them from solutions of their compounds.

(A more reactive halogen will displace a less reactive halogen from solutions of its compounds.)

So colorless potassium iodide solution turns red -orange when chlorine is bubbled through it, because iodine is displaced.



Uses of chlorine

1. To make the plastic polyvinyl chloride(PVC).
2. To make hydrochloric acid.

3. To make solvents such as tetrachloromethane (for dry cleaning).
4. In making bleaches, disinfectants, and insecticides.
5. To sterilize drinking water and water in swimming pools.

Some compounds of chlorine

1. Hydrogen chloride

Hydrogen chloride is made in industry by burning hydrogen in chlorine. The gas has these properties:

- a. It is heavier than air.
- b. It has a choking smell, and it irritates the eyes and lungs.
- c. It dissolves very easily in water, forming hydrochloric acid.

2. Hydrochloric acid

Hydrochloric acid is made in industry by dissolving hydrogen chloride gas in water. It has typical acid properties including:

- a. It reacts with metals to give hydrogen and salts called chlorides.
- b. It reacts with metal oxides and hydroxides to form chlorides and water.
- c. It reacts with carbonates forming chlorides, water and carbon dioxide.

Bromine and Iodine

Like chlorine, bromine and iodine belong to group 7 of the periodic table. They have similar properties to chlorine, but are less reactive.

1. Like chlorine, they react with metals to gain electrons and form negative ions (Cl^- , Br^- and I^-). These ions are called halide ions.
2. They directly combine with hydrogen to form the colorless gases hydrogen bromide (HBr) and hydrogen iodide (HI).
3. Like hydrogen chloride, these gases dissolve very easily in water to form strong acids: Hydrobromic acid and Hydroiodic acid.

4. A more reactive halogen will displace a less reactive halogen from a solution of its compound

Uses of bromine and iodine

- Bromine is used in making fuel additives dyes and pesticides
- Iodine is used in making dyes printing inks animal feeds and pharmaceuticals.
- Like silver chloride silver bromide and silver iodide are used to coat photographic films.

Chapter Review questions

1. Which if the following is an allotrope of sulphur
 - A. monoclinic
 - B. graphite**
 - C. sulphuric acid
 - D. sulphur dioxide
2. in which of the following hydrogen is not used
 - A. manufacture of ammonia
 - B. manufacture of hydrochloric acid
 - C. manufacture of margarine
 - D. coolant and heat exchange in different manufacturing proses
3. **which of the following compounds would not be use as fertilizer**
 - A. NH_4NO_3**
 - B. NaCl
 - C. KNO_3
 - D. K_3PO_4
4. THE FORMULA OF DINITROGEN OXIDE IS
 - A. NO
 - B. NO_2
 - C. N_2O
 - D. N_2O_4
5. the chemical formula of ammonium nitrate
 - A. NH_4NO_3
 - B. NH_4SO_3
 - C. $(\text{NH}_4)_3\text{PO}_4$
 - D. $(\text{NH}_4)_2\text{SO}_4$
6. THE ONLY NON-METAL THAT IS LIQIUD AT ROOM TEMPERATURE IS
 - A. Chlorine
 - B. Argon
 - C. Bromine
 - D. Sodium
7. Which of the elements possess allotropes
 - A. Nitrogen
 - B. Argon
 - C. Chlorine
 - D. Sulphur
8. Which of the following is **NOT** the conversion of nitrogen from air to nitrates in soil
 - a. Nitrifying-bacteria
 - b. Fertilizers
 - c. Lightning-flash

- d. Nitrogen-fixing bacteria
9. Acidic oxides would be formed by oxygen and
- A.** Metals **B.** non metals **C.** metalloids **D.** none of them

Part two answer all the following questions

1. Describe the chemical test of oxygen?
2. Consider the reactions to produce ammonia by Haber process
- $$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 \quad \Delta H = -92 \text{ KJ/MOL}$$
- A.** What does the sign mean?
- B.** Is the reaction is exothermic or endothermic?
3. In Haber process nitrogen and hydrogen are reacted over iron oxide catalyst to give ammonia gas
- A.** Explain how hydrogen gas in the process is obtained?
- B.** Write the equation for the formation of ammonia?
- C.** Explain how the following factors would affect the yield of ammonia
- D.** A. high pressure
- E.** Low temperature
4. Write three properties of each of the following
- A.** Hydrogen
- B.** Sulphur
- C.** Sulphuric acid
- D.** Chlorine
5. Sulphur dioxide dissolves in water to form an acid
- A.** What is the acid is called
- B.** Write an equation for the reaction?
6. Iodine in solution is red brown. Potassium iodide is colorless what would you see when chlorine is bubbled through potassium iodide solution
- A.** Write an equation for reaction?
- B.** What is being oxidized in this reaction?
- C.** What is being reduced?

7. What are halide ions?

Chapter five: Organic chemistry

Introduction

Organic chemistry is a branch of chemistry that involves the study of the structures, composition, and synthesis of carbon-containing compounds. There is such a vast number and variety of these compounds **that** a whole branch of chemistry is allocated to their study. Ranging from drugs to medicines to plastics, our modern world is totally dependent on organic compounds.

Organic chemistry is the branch of chemistry that deals with carbon compounds. Organic compounds: are carbon compounds found in living things, or based on compounds from living things. In their molecules carbon atoms always form the spine. There are thousands of organic compounds many more than all the inorganic compounds put together.

In order to make sense they are arranged in families of compounds with similar properties.

Alkanes

Alkanes are hydrocarbons (they contain only carbon and hydrogen) .

First four alkanes are methane, ethane, propane, and butane.

Alkanes are hydrocarbons which have single bonds between all the atoms.

They have the maximum number of hydrogen atoms for the number of carbon atoms present, and so alkanes are called saturated hydrocarbons.

Name	Formula	Structure	N. carbon	b. point
Methane	CH ₄	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$	1	-164 C°

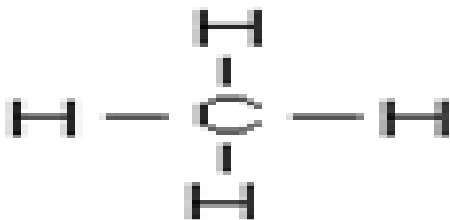
Ethane	C₂H₆	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$	2	-87 C°
Propane	C₃H₈	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$	3	-42C°
Butane	C₄H₁₀	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$	4	-0.5 C°

Note that

- The chain grows by one carbon atom each time.
- The family can be described by general formula **C_nH_{2n+2}** where "n" represents a number of carbons.
- A family of compounds like this, which fits a general formula and has similar chemical properties, is called a **homologous series**.
- *A group of organic compounds with similar properties is called a Homologues series.*

THINGS TO REMEMBER ABOUT ALKANES

1. They are found in oil and natural gases. Natural gases are mostly methane with small amounts of ethane propane and butane.
2. The longer the chain the higher the boiling point. So first four alkanes are gases at room temperature, the next twelve are liquids, and rests are solids.
3. In alkane molecule, each carbon atom forms four single covalent bonds like this



4. Alkanes burn well in a good supply of oxygen, forming carbon dioxide and water vapour and giving out heat. So they are used as fuels. For example burns like this:



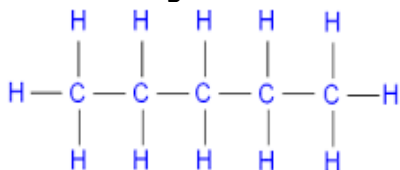
5. But if there isn't enough air, they undergo incomplete combustion when they burn, producing carbon monoxide and giving out less heat.
6. Alkanes also react with chlorine and bromine. For example:



The reaction needs light. It is called **substitution reaction**.

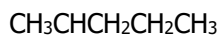
Naming alkanes

1. Find the longest continuous chain of carbon atoms.



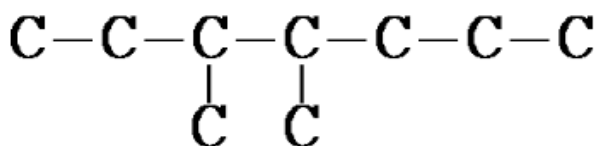
Pentane

2. Branched names are names in the same way. The name given to the longest continues carbon chain is then prefixed by the name of shorter side chains. The stems are used with the suffix -yl. Hence.....**CH₃** is methyl. The position of the alkyl group is indicated by a number. The carbon atoms in the longest carbon chain are numbered from one end of the chain. Numbering starts from the end that produces the lowest possible numbers for the side chains.



CH₃ **2 methyl pentane**

3. Each side chain must be included in the name. if there are several identical side chains the name is prefixed by (di ,tri and tetra)



3,4 –dimethyl heptanes

4. The numbers are separated by comma while number and a letter are separated by hyphen.

2,2,3- trimethyl octane

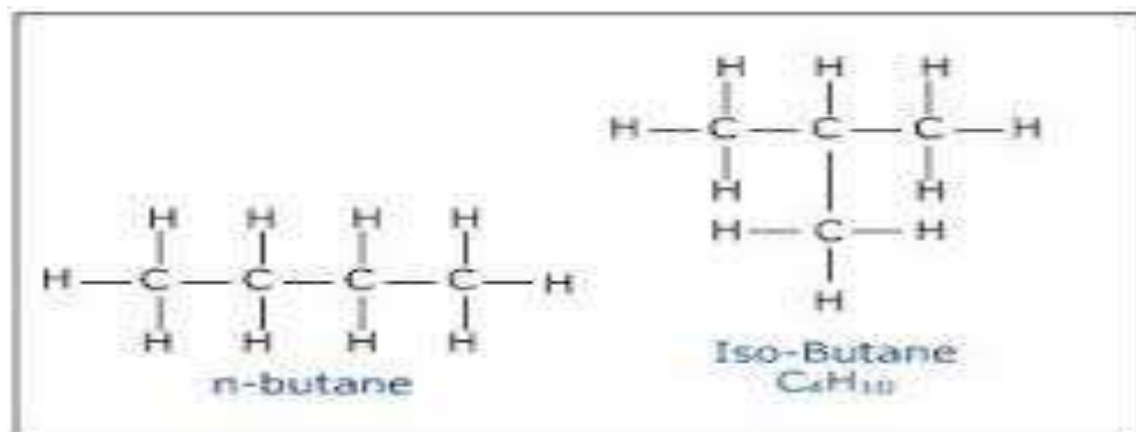
5. Where different alkyl groups are present they are placed in alphabetical order.

3-ethyl -2-methyl pentane

Isomers of alkanes

Isomers are compounds with the same formula but different structures.

Compare the structures of the two molecules below. Both of them have the formula C₄H₁₀. But they have different structures.



The formula for finding the number of isomers that an alkane molecule can form is $2^{n-4} + 1$.

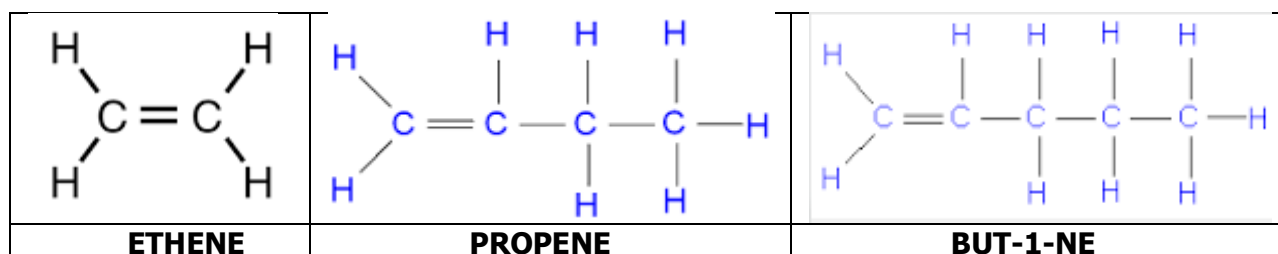
The more carbon atoms there are in a chain, the more isomers there will be. Since isomers have different structures they also have slightly different properties

- Have lower boiling point than straight chains. The branches make more difficult for the molecules to get close and attract each other.
- Burn little readily with oxygen

Alkenes

In this unit we look another family of hydrocarbons which are known ALKENES.

The first three alkenes are ETHENE PROPENE AND BUT-1-ENE.

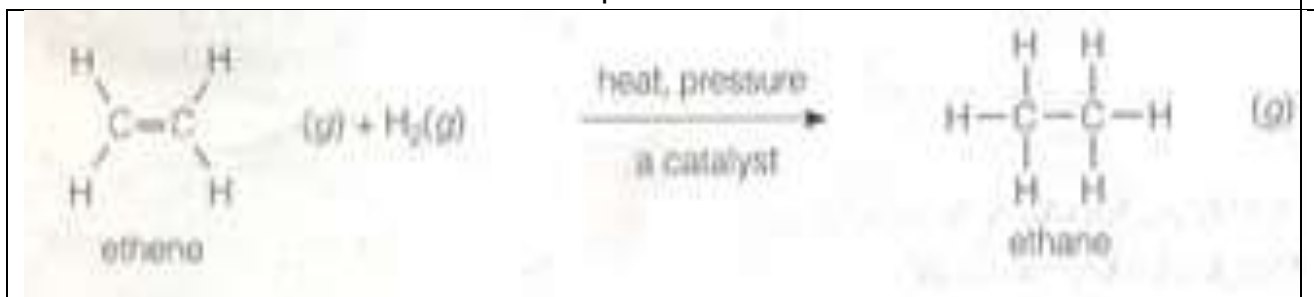


Note that the alkenes fit the general formula C_nH_{2n} . Like the alkanes they form a homologous series.

In alkenes molecule there is only one double bond between carbon atoms and the double bond allows them to add on other atoms for that reason alkenes are said to be un- saturated.

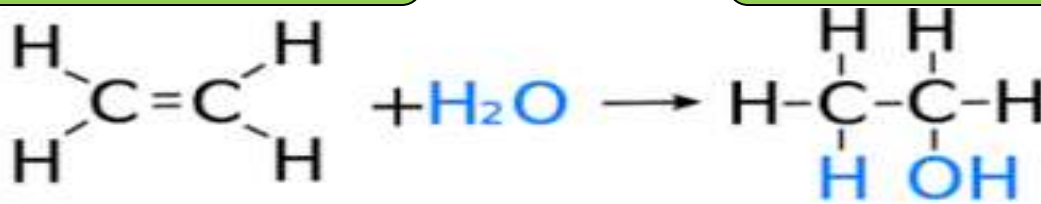
Properties of alkenes

1. They are made from alkanes from cracking.
2. In an alkene molecule contain one double between carbon atoms
3. Alkenes are much more reactive than alkanes. This is because the double can break and add on other atoms. For example



This is called addition reaction

4. Alkenes undergo a similar addition reaction with a water to form compounds called **ALCOHOLS**.



ethanol Ethane water

5. Alkenes molecules can also add on to each other to form long chain molecule called **POLYMERS**.

And the reaction is called polymerization reaction. For example ethane polymerizes to form poly ethane like this



6. Because the double bond allows them to add on more atoms alkenes are said to be unsaturated.
The alkanes don't have a double bond and cannot add more atoms so they are saturated
7. The alkenes are highly flammable and burn readily in air, forming carbon dioxide and water, For example, ethene burns as follows :



Test for un-saturation

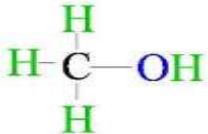
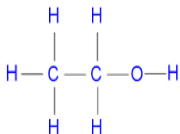
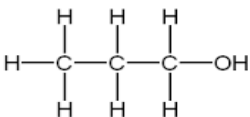
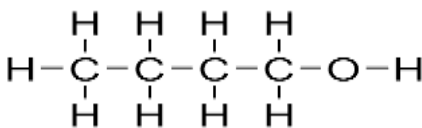
The carbon-carbon double bonds in unsaturated oils can be detected using the elements bromine or iodine. These elements react with the double bonds in the oils, and the more double bonds there are, the more bromine or iodine is used up.

You can check for unsaturated fats using a simple test with bromine water. The test is similar to one used to differentiate alkenes from alkanes.

Bromine water is a dilute solution of bromine, which is normally orange-brown in colour. It becomes colorless when shaken with alkenes, or with unsaturated fats. When shaken with alkanes or saturated fats, its colour remains the same.

Alcohols

The third family of organic compounds are alcohols. Alcohol contains the -OH group. This is called functional group because it determines their reactions. It is the reason why they all react in a similar way. This table shows the first four members

Name	Methanol	Ethanol	Propanol	Butanol
Formula	CH_3OH	$\text{C}_2\text{H}_5\text{OH}$	$\text{C}_3\text{H}_7\text{OH}$	$\text{C}_4\text{H}_9\text{OH}$
Structure of molecule				
Number of carbon atoms in chain	1	2	3	4
Boiling points	65°C	78°C	97°C	117°C

Ethanol is an important alcohol

Ethanol is the best known of all the alcohols

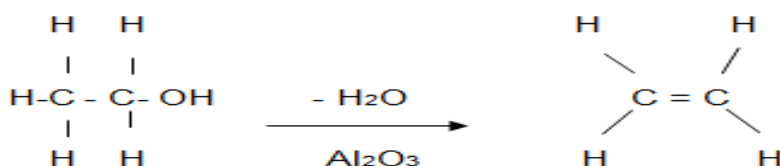
- It is the ingredient in alcoholic drink
- It is good solvents dissolving many compounds that are insoluble in water.
- It evaporates easily, that makes good solvents for use in things like glues and printing inks.
- It is the starting points of many chemicals. For example it is used in making sweet smelling liquids called esters.
- It is also growing in importance as a fuel.

Things to remember about ethanol

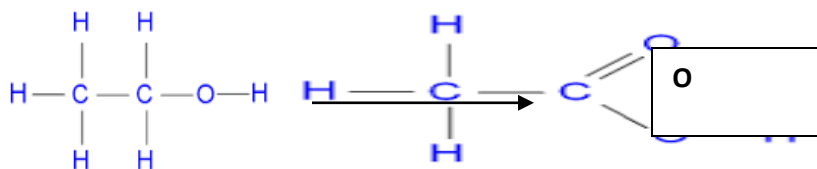
1. It is clear, colorless liquid that boils at 78°C .
2. It is miscible with water- it mixes completely with it.
3. It burns well in oxygen, giving out plenty of heat.



4. Ethanol can be dehydrated to form ethene.



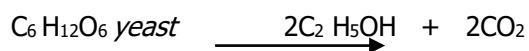
5. If ethanol is left standing in air it will be oxidized with the help of bacteria forming ethanoic acid.

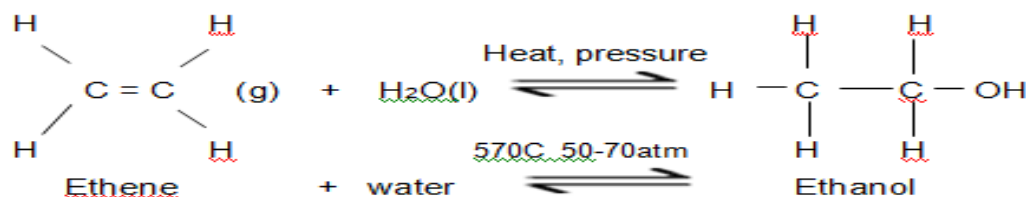


The manufacture of ethanol

Ethanol by fermentation:

Look what happens when you mix glucose and yeast in absence of air.



Ethanol from ethene:

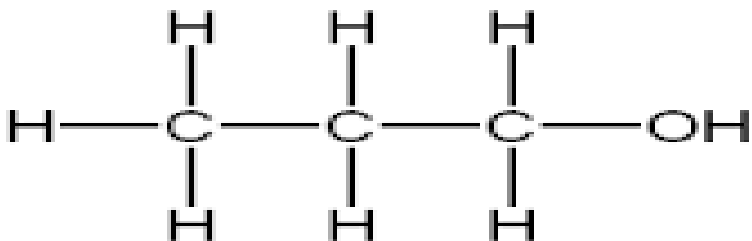
Ethanol fermentation	Ethanol from ethene
Advantages	advantages
It uses renewable resources	The reaction is fast
It is good way to use waste organic material	The process produces pure ethanol
Disadvantages	Disadvantages
Large volume of material is used to make 1 liter of ethanol	Ethene is made from oil which is non-renewable resource
Fractional distillation uses energy	A high percentage of ethene remains un-reacted
Fermentation is slow	Energy is needed to make steam

Types of alcohols

Alcohols are classified into three types according to the position of their functional group OH.

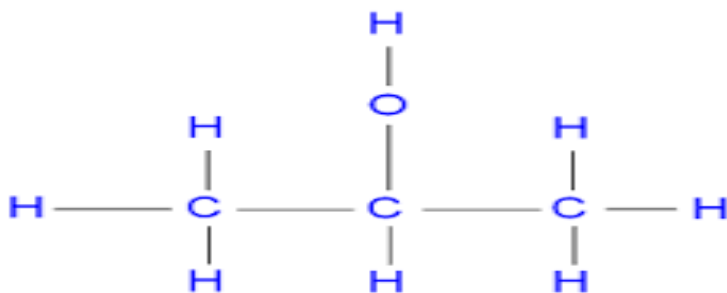
Primary alcohol

In primary alcohol the OH group is attached on carbon which is not attached to any other carbon atom for example



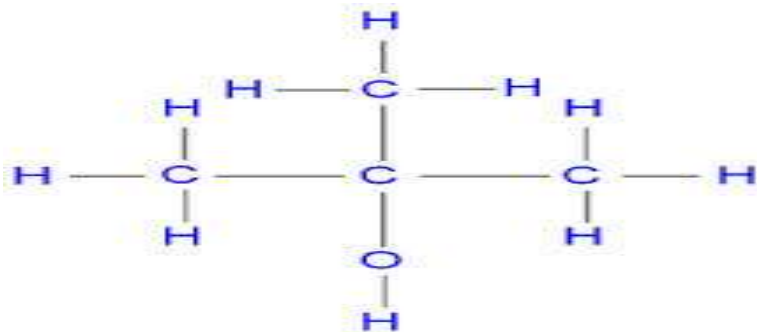
Secondary alcohol

In secondary alcohol the OH group is attached on carbon which on turn attached on on two other carbon atoms for example



Tertiary alcohols

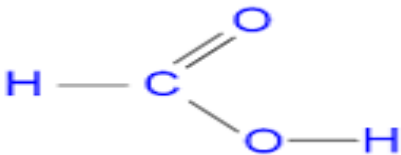
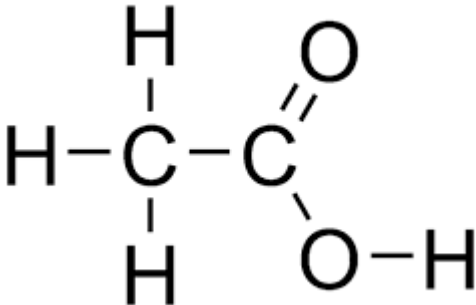
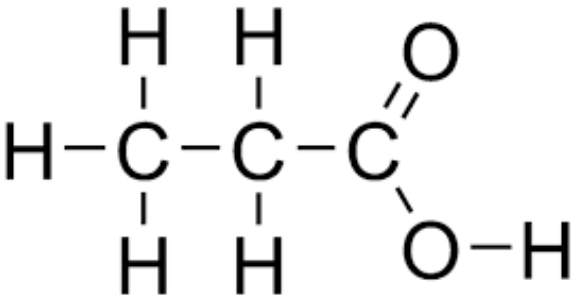
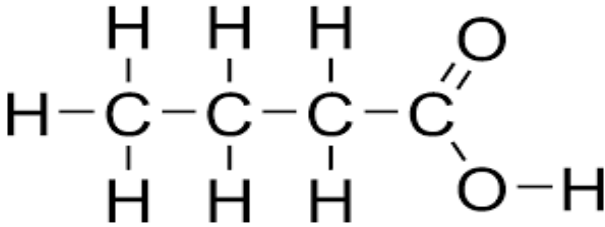
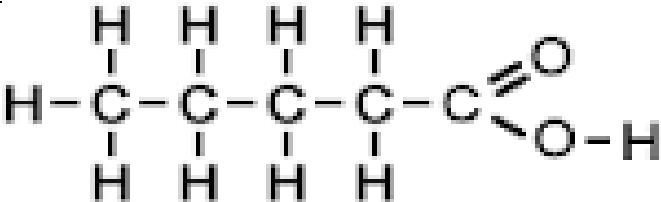
In tertiary alcohol the OH group is attached on carbon which on turn attached on three other carbon atom for example



Carboxylic acid and their compounds

Now we look another family of organic acids which are called carboxylic acid. All the members of this family have the functional group -COOH which are called the **CARBOXYL GROUP**.

Here are the first five members.

Name of the acid	Structure of the molecule	Formula	Carbon atoms in a chain
Methanoic acid		HCOOH	1
Ethanoic acid		CH ₃ COOH	2
Propanoic acid		CH ₃ CH ₂ COOH	3
Butanoic acid		CH ₃ CH ₂ CH ₂ COOH	4
Pentanoic acid		CH ₃ CH ₂ CH ₂ CH ₂ COOH	5

The general formula for carboxylic acids is **C_nH_{2n+1}COOH**. Do you agree?

.....

The properties of carboxylic acids

1. Their solutions affect the color of indicators just like in organic acids.
2. Like in organic acids their solutions contain **H⁺** ions. For example when ethanoic acid is dissolved in water it forms ions like this



3. Like in organic acids they react with metals. Alkali metal carbonates and metal hydrogen carbonates to form salts. For example



Ethanoic acid + sodium hydroxide sodium ethanoate + water

4. They react reversibly with alcohols to form sweet smelling ESTERS. Concentrated sulphuric acid is used as catalyst.



Ethanoic acid + ethanol ethyl ethanoate + water

MULTIPLE CHOICES questions

1. the diagram below shows two ways of preparing substance X



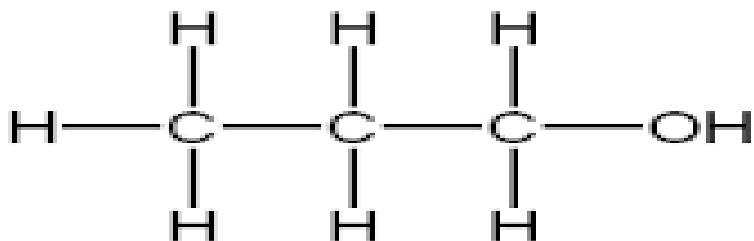
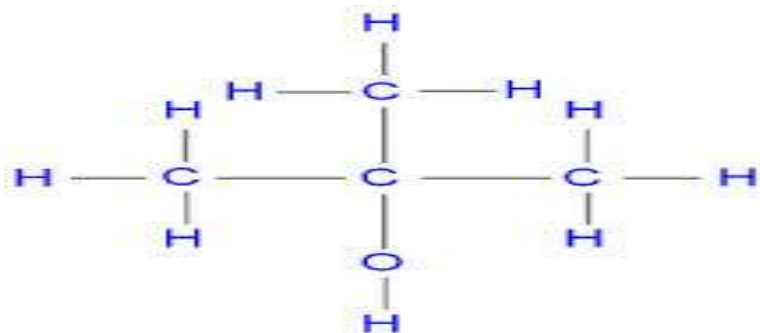
SUBSTANCE X IS

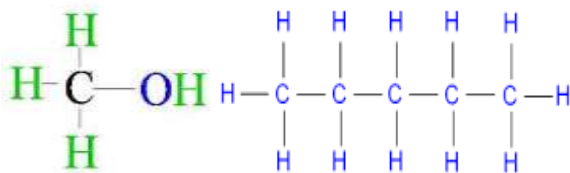
- b. CARBON DIOXIDE
 - c. ETHANOL
 - d. GLUCOSE
 - e. METHANE
2. Which of the following molecules are produced when ethene is hydrated with concentrated sulphuric acid?
 - A. C_2H_6
 - B. C_2H_5OH
 - C. CH_4
 - D. CO_2
3. HOW MANY DIFFERENT ISOMERS OF PENTANE HAS
 - A. 1
 - B. 2
 - C. 3
 - D. 4
4. Organic molecules with the same molecules but different structural formula are called
 - A. Allotropes
 - B. Isomers
 - C. Isotopes
 - D. Polymer
5. Which of the following is not hydrocarbon?
 - A. Methane
 - B. Ethane
 - C. Ethanol
 - D. Ethyl
6. Hydrocarbon compounds burn oxygen to give
 - A. Water and carbon dioxide
 - B. Water only
 - C. Hydrogen and carbon dioxide
 - D. Hydrogen only
7. Which of the following is tertiary alcohol
 - A. 3,3 -di methyl butanol
 - B. Butanol
 - C. 2-methyl propan-2-ol

- D. Pentan-2-ol
8. One general characteristics of homologous series is that the compound all have the same
- A. Molecular formula
B. Empirical formula
C. Structural formula D. general formula

Answer the following questions

1. What is the general formula of alcohols?
2. What is the most common reaction that alkenes made is called?
3. What is the most common reaction that alkanes made is called
4. The hydrocarbons methane ethane and propane are members of the same homologous series
 - A. What does mean hydrocarbon?
 - B. Name the homologous series these belong?
 - C. Name a raw material rich in hydrocarbons?
5. Name the following compounds





6. Explain why it is possible to distinguish ethene and ethane using bromine water?
7. Ethanoic acid is a carboxylic acid and had the formula CH_3COOH
 - A. Give the name and structural formula of carboxylic acid that has three carbon atoms in its molecule
 - B. Write an equation with ethanoic acid dissociates with water?
8. Ethanoic acid reacts with methanol to form ester
 - A. What conditions are necessary for this reaction to occur?
9. Write an equation for the combustion of
 - A. Ethanol
 - B. Methane
10. what is called the process grape juice is turned into vinegar
11. What is added to the grape juice for the reaction to occur?
12. Write the balanced chemical equation for the reaction between methane and chlorine?
Name the type of this reaction?
13. Explain why decane does not react with bromine water?
14. Draw the displayed formula of pentane?
15. The table below shows the formula of some organic compounds and their boiling point

Compound	Boiling point
C_3H_6	225
C_4H_8	266
C_6H_{12}	336

- A. State the homologous series of which the compounds belong?
- B. Give the molecular formula of the first three members of this family

16, complete the following table

Compound	Ethane	Ethene	Ethyne
Homologous series	Alkane	_____	_____
Molecular formula	_____	_____	_____
Structural	_____	_____	_____
Type of covalent bond	_____	Double bond	_____
Saturated /unsaturated	_____	_____	_____

Chapter 6:Rate of a reaction

Fast or slow

Some reactions are slow and some are fast. It is not always enough to know just that a reaction is fast or slow. It is needed to know exactly how fast that each reaction is going and how long it takes to complete—that is, the rate of a reaction

Rate: is a measure of fast or slow something is.

Rate is a measure of the change that happens in a single unit of time. Any suitable unit of a time can be used. Second, minute, hour or a day.



This plane has just flown 2000 kilometres in 1 hour. It flew at a rate of 2000 kilometres per hour.



This petrol pump can pump petrol at a rate of 50 litres per minute.



This machine can print newspapers at a rate of 10 copies per second.

Rate of a chemical reaction



When zinc is added to dilute sulphuric acid, they react together. The zinc disappears slowly, and a gas bubbles off.

In this example, the gas that forms is hydrogen.



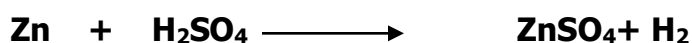
After a time, the bubbles of gas form less quickly. The reaction is slowing down.



Finally, no more bubbles appear. The reaction is over, because all the acid has been used up. Some zinc remains behind.

In this example the gas that forms is hydrogen. The equation for the reaction is

Zinc + sulphuric acid $\longrightarrow$ zinc sulphate + hydrogen



Both zinc and sulphuric acid get used up at the same time Zinc sulphate and hydrogen form.

You could measure the rate of reaction by measuring either

- ❖ The amount of zinc used up per minute
- ❖ The amount of sulphuric acid used up per minute
- ❖ The amount of zinc sulphate produced per minute
- ❖ The amount of hydrogen produced per minute

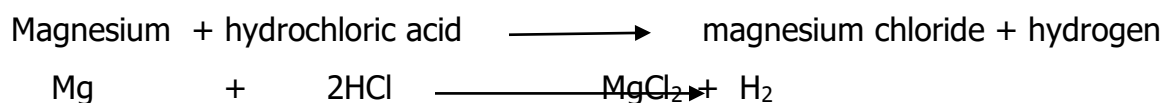
For this reaction it is easiest to measure the amount of hydrogen produced per minute. The hydrogen can be collected as it bubbles off and its volume can then be measured.

In general, to find the rate of a reaction, you should measure:

- ❖ the amount of a reactant used up per unit of time or the amount of a product produced per unit of time

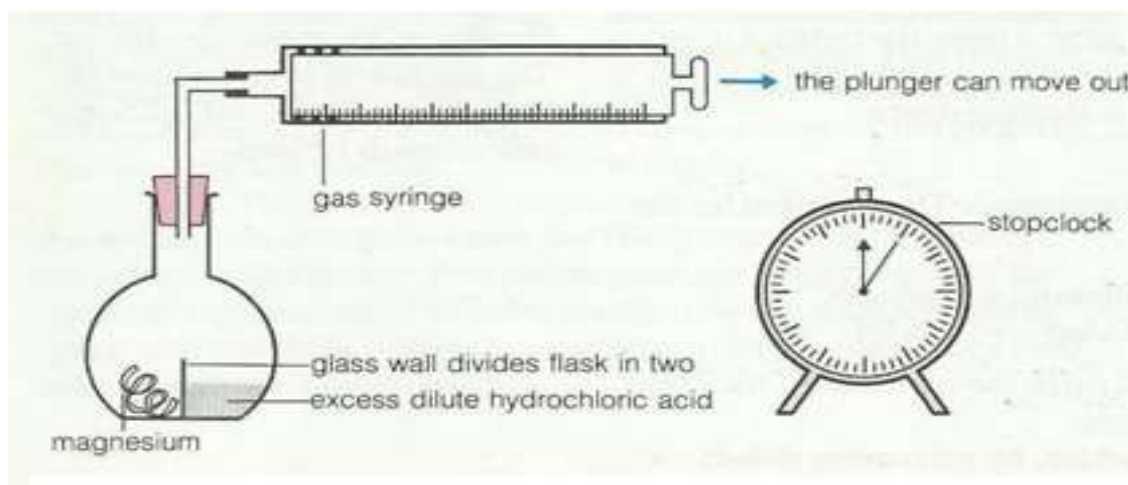
Measuring the rate of reaction

On the last page we saw that the rate of chemical is found by measuring the amount of reactant used up or the amount of product produced per unit of a time. For example the reaction between magnesium and excess dilute hydrochloric acid. Its equation is

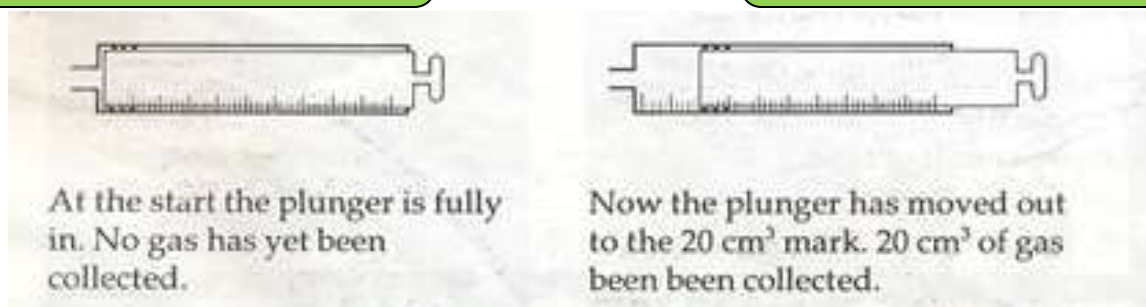


In this reaction hydrogen is the easiest substance to measure. This is because it is the only gas in this reaction. It bubbles off and can be collected in a gas syringe where its volume is measured.

The method.



Magnesium is cleaned with sand paper and put in one part of a conical flask. Dilute hydrochloric acid is put in other part. The flask is tipped up to let the two reactants mix and the clock is started at the same time. Hydrogen begins to bubble off. It rises up the flask and pushes its way into the gas syringe. The plunger is forced to move out.

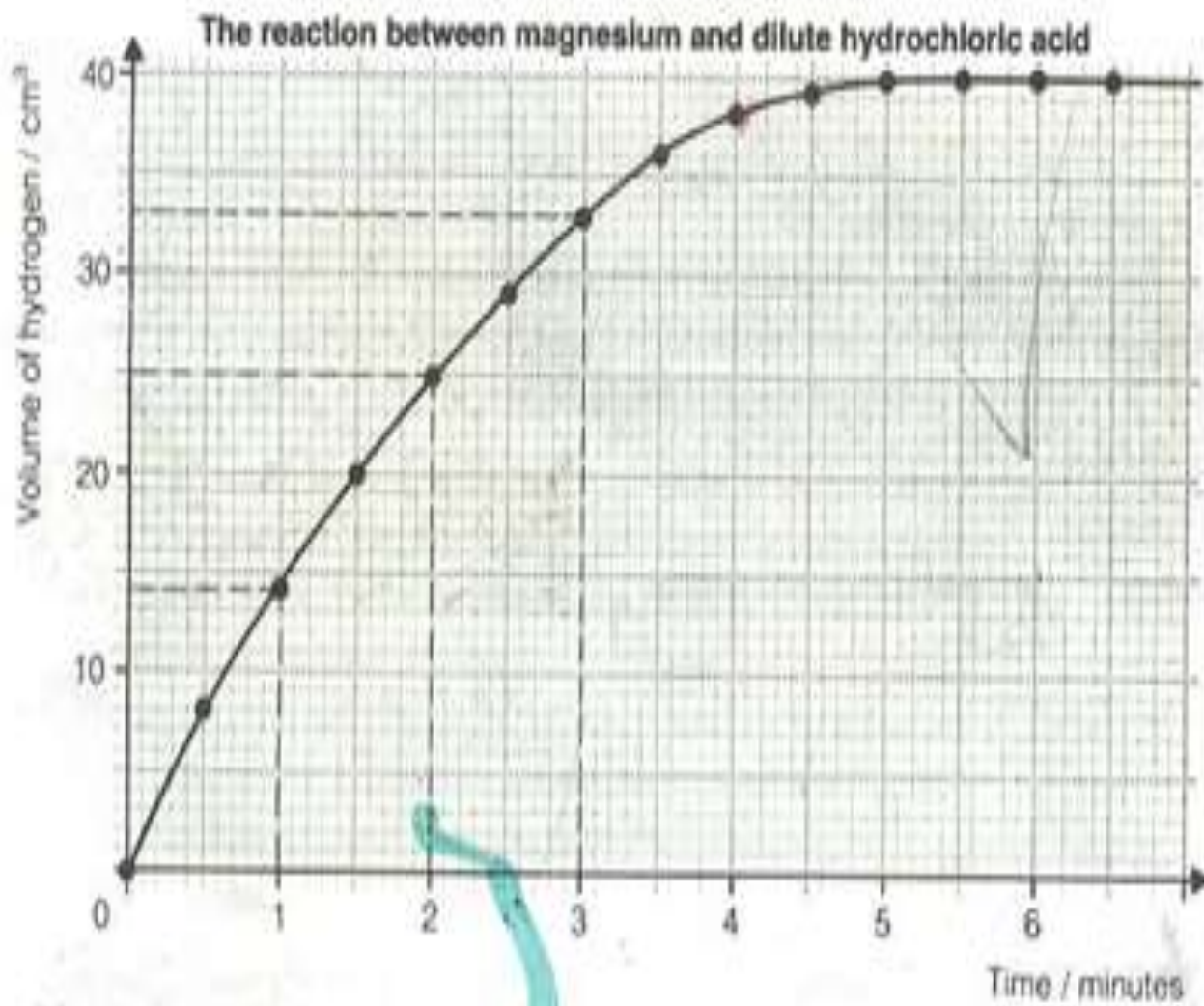


The volume of gas syringe is noted as intervals. For example at the end of each half minute.

The results

Time/minutes	0	$\frac{1}{2}$	1	$1\frac{1}{2}$	2	$2\frac{1}{2}$	3	$3\frac{1}{2}$	4	$4\frac{1}{2}$	5	$5\frac{1}{2}$	6	$6\frac{1}{2}$
Volume of hydrogen/cm ³	0	8	14	20	25	29	33	36	38	39	40	40	40	40

These results can be plotted with a graph.



Notice these things about the results.

1. In the first minute 14 cm^3 of hydrogen is produced. In the second minute only 11 cm^3 is produced ($25-14$). So the rate of for the second minute is 11 cm^3 of hydrogen per minute. The rate of a third minute is 8 cm^3 of hydrogen per minute. So we can see that the rate decreases as the time goes on.

The rate changes all through the reaction. It is greatest at the start but gets less as the reaction proceeds.

2. The reaction is fastest in the first minute and the curve is steepest then it gets less steep as the reaction gets slower.

The faster the reaction the steeper the curve.

1. After 5 minutes no more hydrogen is produced so the volume no longer changes. The reaction is over and the curve goes flat.

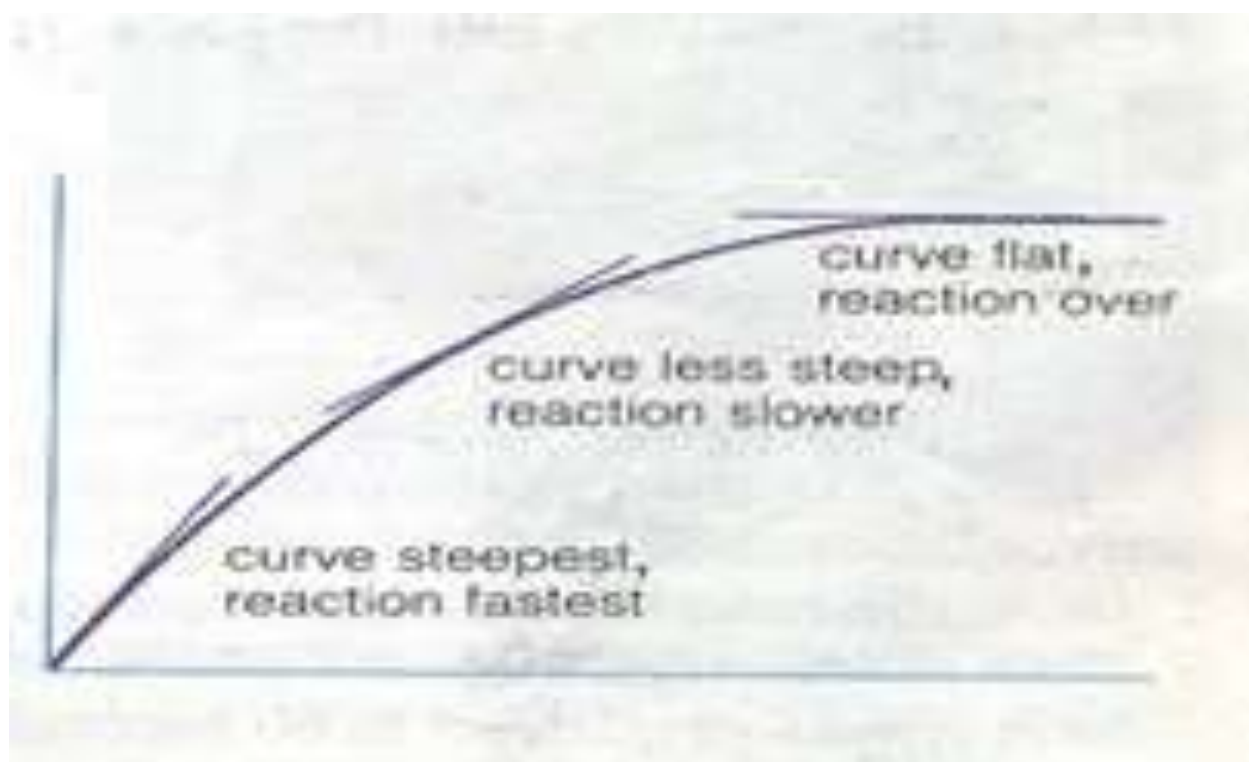
When the reaction is over the curve goes flat.

2. Altogether 40 cm^3 of hydrogen is produced in 5 minutes.

The average rate of reaction = total volume of hydrogen / total time taken

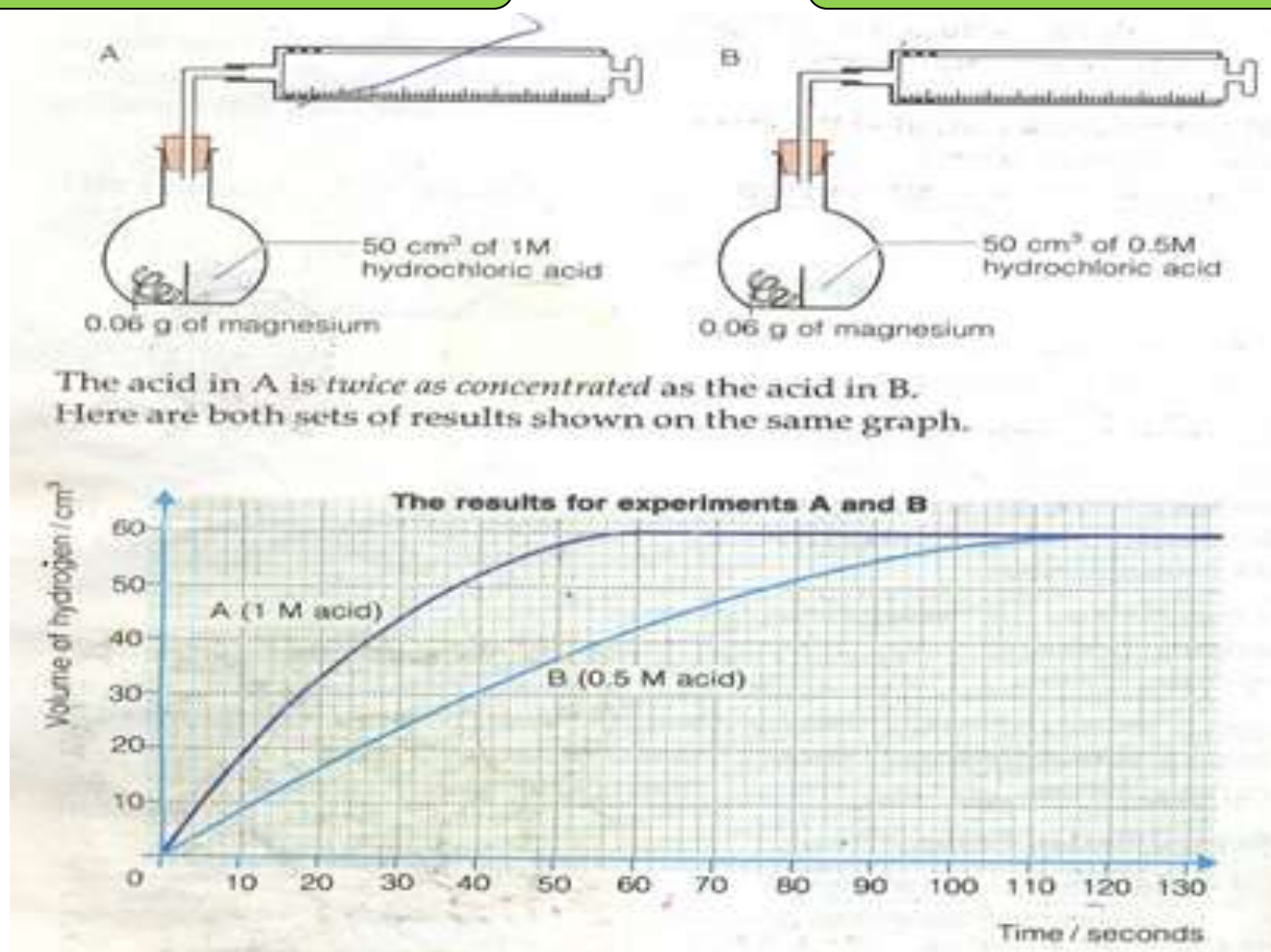
$40 \text{ cm}^3 / 5 \text{ minutes}$

8 cm^3 of hydrogen per minute.



The effect of concentration on rate of chemical reaction.

A reaction can be made to go faster or slower by changing the concentration of reactants. Suppose the experiment with magnesium and hydrochloric acid is repeated twice (A and B). every thing is kept the same each time except the concentration of acid.



Curve A is steeper than curve B. from this we can we can tell straight away that the reaction was faster than in A than B.

In A the reaction lasts 60 seconds and in B it lasts 120 seconds.

Both reactions produces 60 cm³ of hydrogen. In A it was produced in 60 seconds so the average rate was 1cm³ of hydrogen per second. In B it was produced in 120 seconds so the evarage rate was 0.5 cm³ per second.

The rate rate in A was twice the everage rate in B.

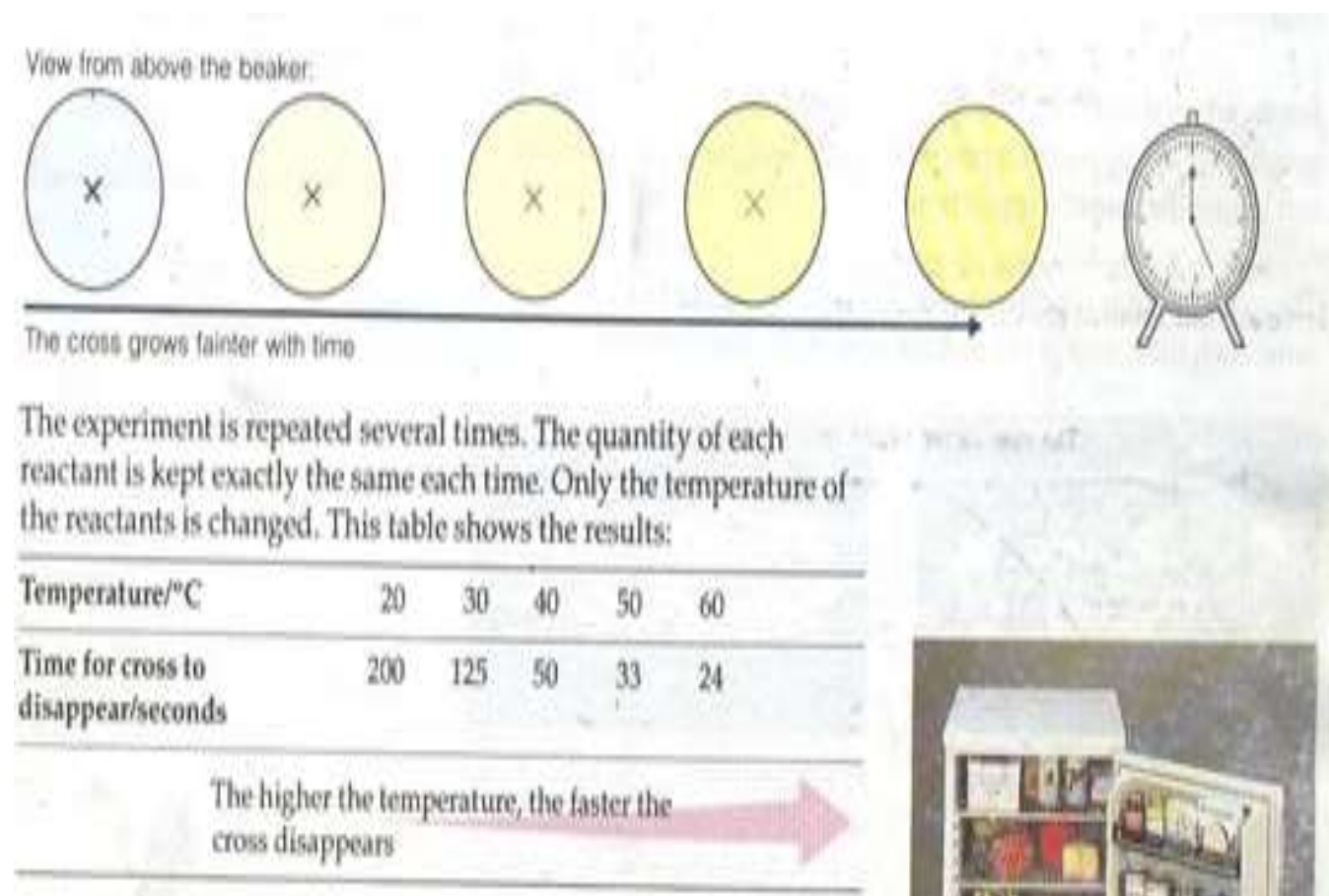
A reaction goes faster when the concentraraion of a reactant is increased.

The effect of temperature

A reaction can also be made to go faster or slower by changing the temperature of the reactants.

This time a different reaction is used. when dilute hydrochloric acid is mixed with sodium thiosulphate solution a fine yellow precipitate of sulphur forms. The rate can be followed like this.

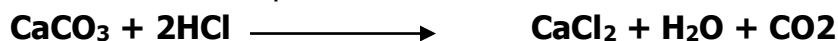
A cross is marked on a piece of paper. A beaker containing some sodium thiosulphate solution is put on top of the paper. The cross should be easy to see through the solution. hydrochloric acid is added quickly and clock is started at the same time. The cross grows fainter as the precipitate forms. The clock is stopped the moment the cross can be no longer seen.



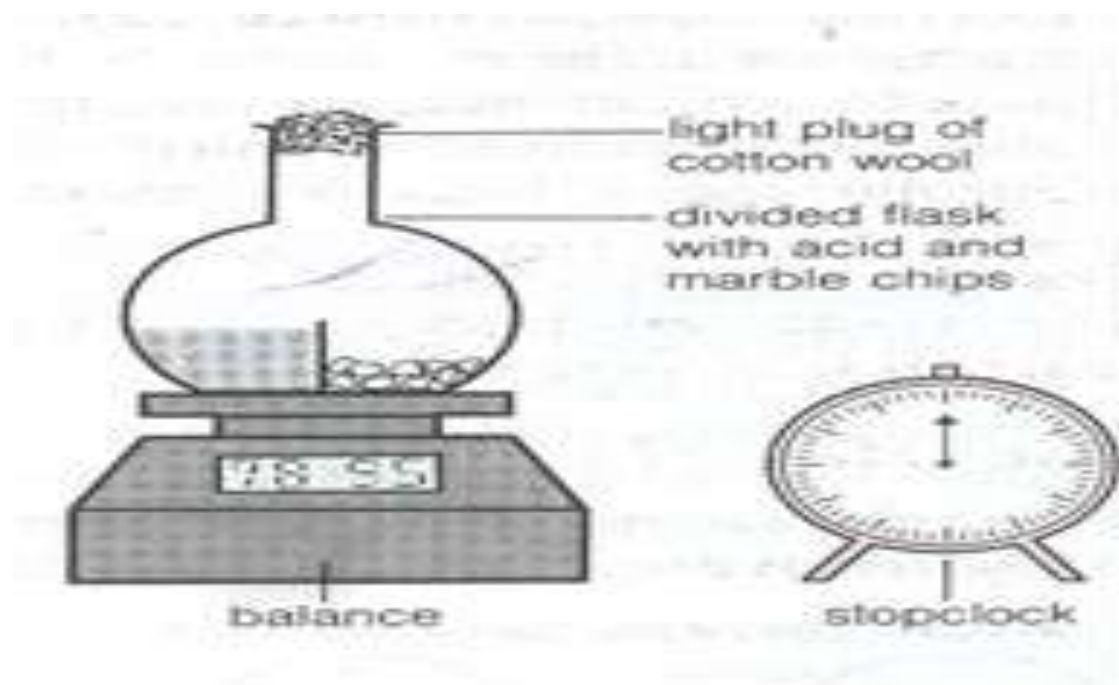
(A reaction goes faster when the temperature is raised. When the temperature is increased by 10 °C the rate doubles.)

The effect of surface area

In many reactions one of the reactants is a solid. The reaction between hydrochloric acid and calcium carbonate (marble chips) is an example. Carbon dioxide is an example.



The rate can be measured using apparatus as below



The method

Marble chips and acid are placed in the flask which is then plugged with cotton wool. This prevents any liquid from splashing out during the reaction. Next the flask is weighed. Then it is tipped up to let the reactants mix and clock is started at the same time. The mass is noted at intervals until the reaction is complete.

Since carbon dioxide can escape through the cotton wool the flask gets lighter as the reaction proceeds. So by weighing the flask you can follow the rate of the reaction.

The experiment is repeated twice. Everything is kept exactly the same each time except the surface area of marble chips.

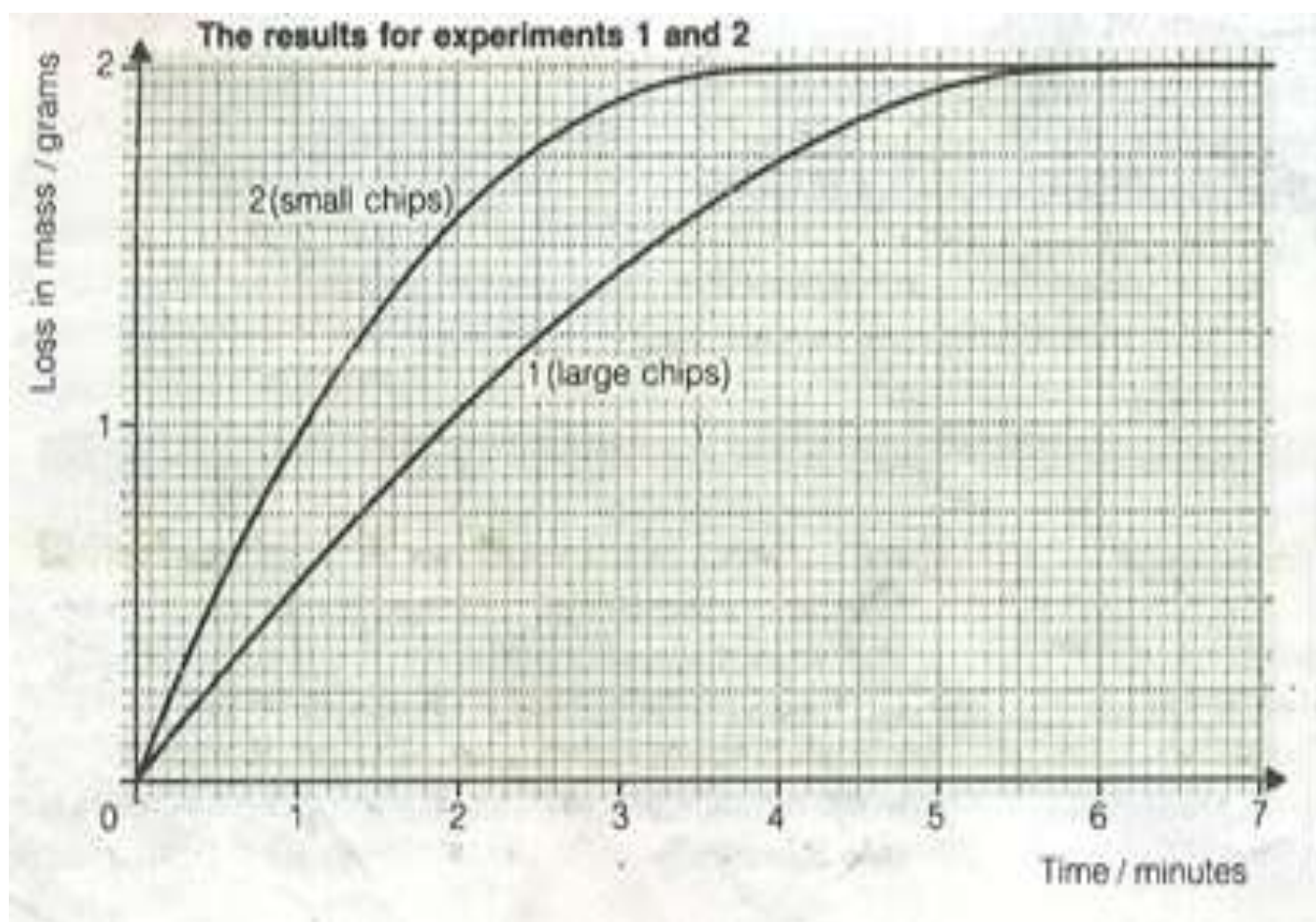


For experiment 1, large chips of marble are used. The surface area is the total area of the surface of these chips.



For experiment 2, the same *mass* of marble is used. But this time it is in small chips, so its surface area is greater.

The results of the two experiments are plotted below



You could notice these things about the result.

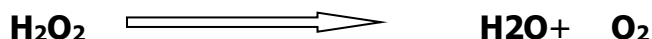
1. Curve 2 is steeper than curve 1. This shown immediately that the reaction is faster for the small chips.
2. In both experiments the final loss in mass is 2 grams. In other words 2 grams of carbon dioxide is produced each time.
3. For the small chips the reaction is complete in 4 minute. For the large it lasts 6 minute.

(The rate of reaction increases when the surface area of a solid reactant is increased.)

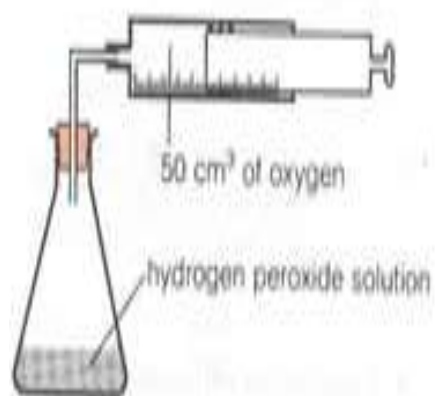
The effect of catalyst

Hydrogen peroxide is clear colourless liquid with the formula H_2O_2 . It can decompose to water and oxygen.

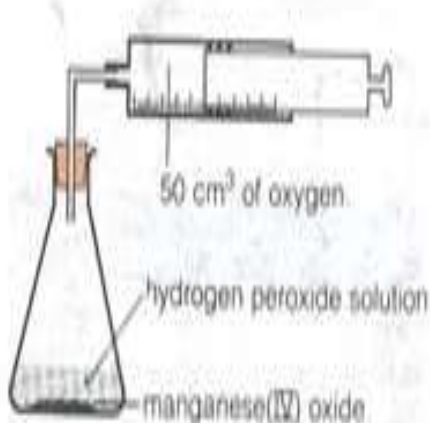
Hydrogen peroxide $\longrightarrow$ water + oxygen



The rate of reaction can be followed by collecting the oxygen.



The reaction is in fact very slow. It could take 500 days to collect 50 cm³ of oxygen.



However, if 1 gram of black manganese(IV) oxide is added, the reaction goes much faster. 50 cm³ of oxygen is produced in a few minutes.



After the reaction, the black powder is removed by filtering. It is then dried, weighed and tested. It is still manganese(IV) oxide, and still weighs 1 gram.

The manganese (iv) oxide speeds up the reaction without being used up itself. It is called **catalyst** for the reaction.

(A catalyst is substance that changes the rate of reaction but remains chemically unchanged itself.)

Explaining rates (Collision theory)

Magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen

In order for the magnesium and acid particles react together

- They must collide to each other
- The collision must have enough energy

(The rate of reaction depends on how many successful collisions there are in a given unit of time.)

Why rate increases with Concentrated reactants have more particles in it. Therefore, more chance of successful collision will occur—that is, the faster the reaction.

Why rate increases with Temperature: when the reacting substances are heated, the particles take in energy; they move faster, which means they collide more often and the rate goes up. For the same reason reaction that depends on the light energy speed up in stronger light.

Why rate increases with Surface Area: when the reactants are powdered, many more atoms are exposed. So there is a greater chance of successful collision.

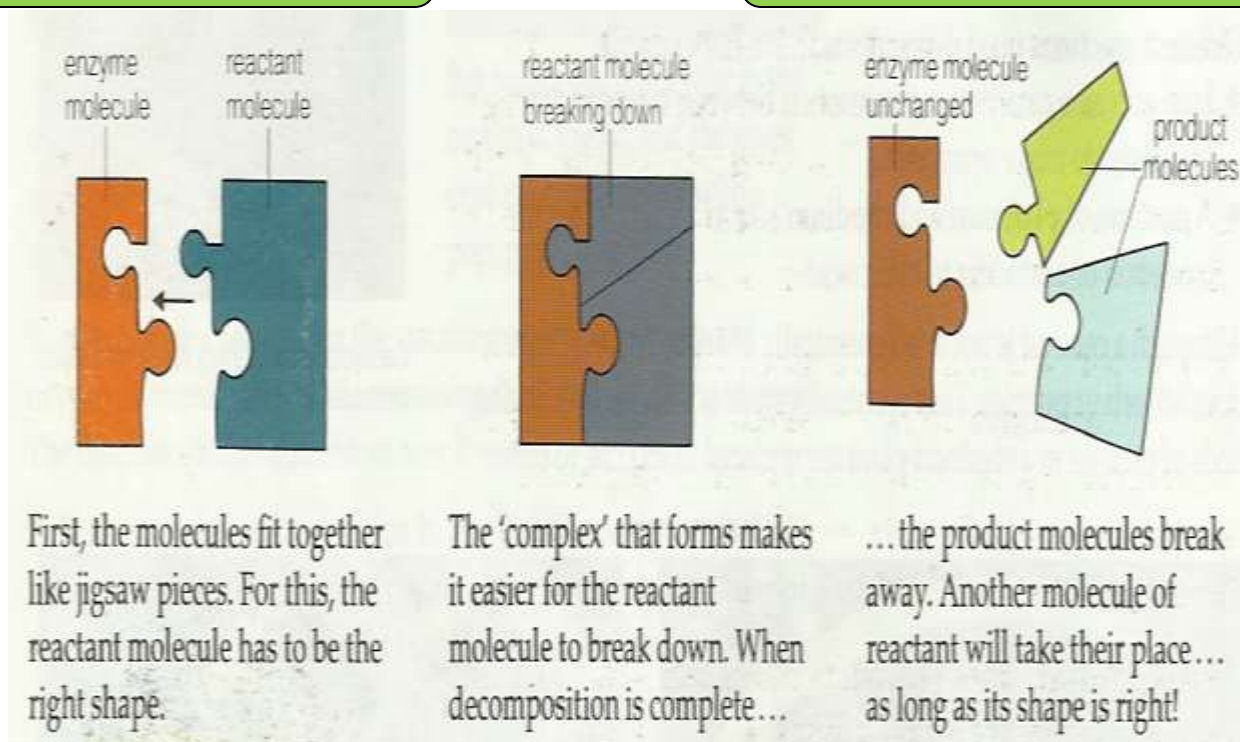
Why Catalyst increases the rate: in a presence of catalyst, a collision needs less energy in order be successful, so the reaction goes faster. Catalysts are very important in industries because they speed up reaction even at low temperature—that is, less fuel are needed so money is saved. Catalyst lowers the activation energy for the reaction. Different reactions need different catalysts

(Catalyst lowers the energy needed for a reaction takes place.)

Enzymes

Enzymes are proteins produced by living organisms. They are catalysts for biological reactions.

How enzymes work



Note: since the shape is so important an enzyme will catalyze only specific reaction where the fit is right.

An enzyme works best between about 25°C and 45°C . At higher temperature it loses its shape and stops working. It becomes denatured. At low temperature it becomes inactive.

Uses of enzymes:

- ❖ Making bread.
- ❖ Making yoghurt, cheese, and chocolate.
- ❖ Making beer and wine.
- ❖ Biological detergents.

Reversible reactions

When you heat blue crystals of copper (II) sulphate, they break down into anhydrous copper (II) sulphate, a white powder. The reaction is easy to reverse: just add water. The white powder turns blue again. In fact this reaction is used for test for water.



The above reaction is reversible: it can go either direction.

Reaction one is forward and reaction two is back ward. We use the symbol $\rightleftharpoons$ to show that the reaction is reversible.

Thermal decomposition of ammonium chloride



Reversible reactions and dynamic equilibrium

Lead (II) chloride is not very soluble. 1.1g forms saturated solution in 100g of water: if you add more lead chloride it will sink to the bottom.



With stopper on, the flask becomes a closed system. Ions move continually from the solid to the solution and from solution to the solid. This happens at the same rate in both directions so no extra lead (ii) chloride is dissolving. We can describe the situation like this.



Inside the flask the solid and its saturated solution have reached a state of dynamic equilibrium.

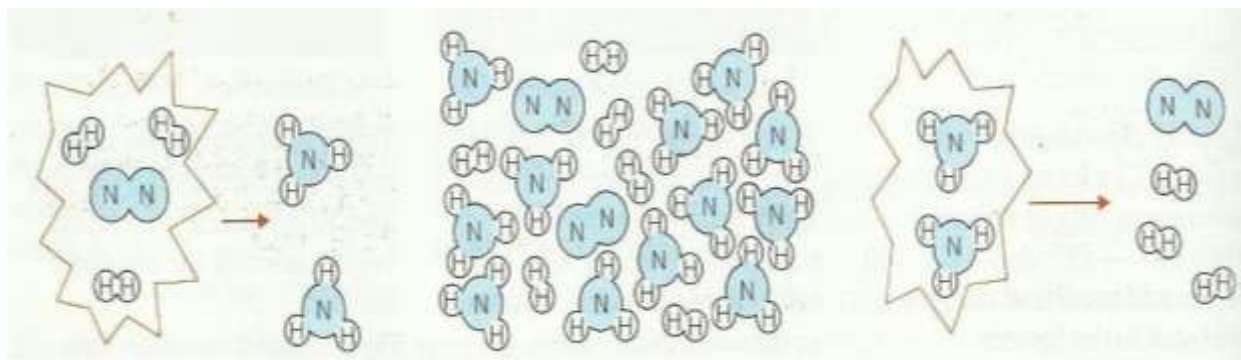
Equilibrium means there is no overall change. But dynamic means that change is still taking place continuously.

(In a closed system a reversible reaction reaches a state of dynamic equilibrium. The forward and back ward reactions take place at the same rate, so no overall change takes place.)

Shifting the equilibrium

The manufacture of ammonia

Imagine you run a factory that makes ammonia from nitrogen and hydrogen. As you saw, the reaction is reversible.



Three molecules of hydrogen react with one of nitrogen to form two ammonia. So if you mix the right amount of nitrogen and hydrogen.

will it turn all ammonia molecules? No! Once certain ammonia is formed the system reaches dynamic equilibrium.

Every time, two ammonia molecules form, another two break down into nitrogen and hydrogen. So the level of ammonia remains unchanged.

But more ammonia that forms the better for you profits. So can you increase the product?

The idea called le chatelier principle will help us.

(When a reversible reaction is in equilibrium and you make a change, it will do what it can to oppose that change)

You cannot make a reversible reaction go to completion. It always ends up in state of equilibrium. But by changing the conditions you can shift equilibrium to the right and obtain more products.

1. Increasing the temperature

Heat speeds up any reaction Forward reaction is exothermic backward reaction is endothermic. If you heat the equilibrium mixture it will act to oppose the change, that is, more ammonia will break down in order to use up the heat. Therefore, moderate temperature is recommended

2. Increasing the pressure

Pressure is caused by collisions between gas molecules and the wall of the container, if you apply more pressure, the equilibrium mixture will act to oppose the change more ammonia will form in order to reduce the number of molecules

3. Removing the ammonia

When you remove the ammonia, nitrogen and hydrogen react to restore the equilibrium; suppose you cool the mixture ammonia condenses first, so you can run it off as a liquid, and then warm the remaining nitrogen and hydrogen again.

4. Adding a catalyst

Iron acts as a catalyst for this reaction but note that it speeds up the forward and back reactions equally although the reaction reaches equilibrium faster, the equilibrium position does not change.

Choosing optimum condition

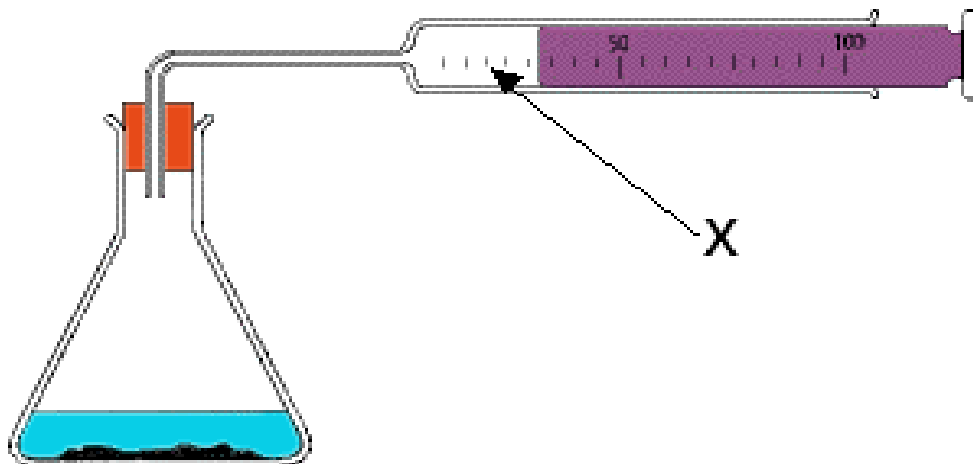
1. Use high pressure
2. Remove the ammonia to shift the equilibrium to the right
3. Use catalyst to reach equilibrium quickly
4. Use moderate temperature

Chapter Review questions

1. Which of the following statements best describes the action of catalyst in a reaction
 - A. It lowers the energy of reactants
 - B. It lowers the energy of the product
 - C. It lowers the activation energy

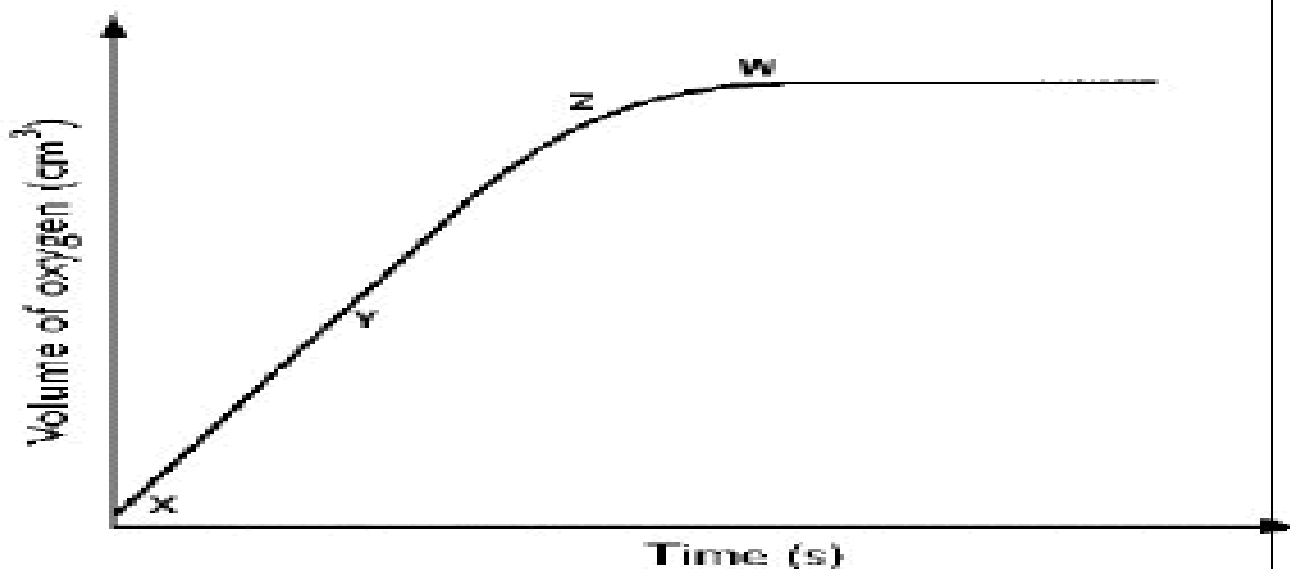
- D. It lowers the rate of the reaction
2. Which statement is true about chemical reaction in equilibrium
- A. The forward and backward reaction proceed at equal rate
 - B. The forward and backward reactions have stopped
 - C. The concentration of the reactants and products are equal
 - D. The forward reaction is exothermic
3. Which of the following of factors does not affect the rate of a chemical reaction
- A. Colour of the reactants
 - B. Temperature of the reactants
 - C. Particle size
 - D. Nature of the reactants
4. If the temperature of a reaction is decreased, what effect will this have on the rate of reaction
- A. The rate of the reaction decrease
 - B. The reaction will stop
 - C. It will have no effect
 - D. The rate of reaction will decrease
5. What effect does doubling the concentration of one of the reactants have on the rate of reaction
- A. It decrease the rate of the reaction
 - B. It has no effect
 - C. The rate of the reaction is halved
 - D. It increases the rate of reaction
6. For the reaction between calcium carbonate and dilute hydrochloric acid
- $$\text{CaCO}_3 + \text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$$
- which one of the following factors will sped up the reaction?
- A. Use smaller pieces of calcium carbonate
 - B. Use more dilute acid
 - C. Use a bigger conical flask
 - D. Lower the temperature of the hydrochloric acid

7. In a reaction between two gases, which one of the following factors will **not** affect the rate of reaction?
- A. The temperature of the gases
 - B. The colour of the gases
 - C. The volume of the container that the gases react in
 - D. The pressure of the gases
8. A substance that increases rate of reaction without being changed itself is called
- A. Enzyme
 - B. Allotrope
 - C. Isotopes
 - D. Catalyst
9. When **powdered** calcium carbonate is added to a solution of hydrochloric acid, it reacts completely within seconds. When the same mass of calcium carbonate **chips** are added to a similar amount of hydrochloric acid, the reaction takes several minutes to go to completion. Why?
- A. The powder has a greater surface area.
 - B. An acid will react faster with a powder
 - C. The calcium carbonate chips are less soluble than the powder.
 - D. There are more impurities in the calcium carbonate chips
10. The diagram shows the apparatus for studying the rate of reaction for the decomposition of hydrogen peroxide. Name the gas labeled **X** in the syringe.



- A. Oxygen
- B. carbon dioxide
- C. Water vapor
- D. Hydrogen

11. The diagram shows the volume of oxygen gas collected over time for the decomposition of hydrogen peroxide. At which point on the graph has the reaction stopped?



- A. W
B. Y
C. X
D. Z
12. Why does the rate of reaction increase when powdered calcium carbonate is used instead of marble chips
- A. The powdered calcium carbonate acts as a catalyst
B. There is an increase of the particle size of the calcium carbonate
C. There is an increase of the concentration of the calcium carbonate
D. There is an increase of the surface area of the calcium carbonate
13. If the temperature of a reaction is increased by 10°C, the reaction rate will be
- A. Four times faster
B. Double
C. Ten times faster
D. Halved

14. The diagram shown for measuring the rate of reaction between hydrochloric acid and marble chips, what is the purpose of the cotton wool?



- A. To prevent air getting into the conical flask
- B. To prevent heat escaping from the flask
- C. To prevent the contents of the flask from splashing out
- D. To prevent hydrogen gas escaping

Glossary

Acid

A sour substance which can attack metal, clothing or skin, The chemical opposite of an alkali. When dissolved in water, its solution has a pH number less than 7. Acids are proton (H^+ ion) donors.

Activation Energy

The minimum amount of energy needed to start a reaction.

Alkali

A soluble base with a pH greater than 7. Makes OH^- ions in water.

Alkali Metal

Elements in group 1 of the periodic table e.g. Lithium.

Alkane

Saturated hydrocarbon with the general formula C_nH_{2n+2} e.g. Methane (CH_4).

Alkene

Unsaturated hydrocarbon which contains a carbon-carbon double bond. Its general formula is C_nH_{2n} e.g. Ethene (C_2H_4).

Alloy

A mixture of metals (and sometimes non metals). E.g. brass = copper and zinc.

Aluminium

A low density corrosion-resistant metal used in many alloys, including those used in the aircraft industry.

Anhydrous

A substance which does not contain water.

Aqueous Solution

The mixture made by adding a soluble substance to water.

Atmosphere

The relatively thin layer of gases that surround planet Earth.

Atom

The smallest part of an element that can still be recognised as that element.

Atomic Number

The number of protons (which equals the number of electrons) in an atom. It is sometimes called proton number.

Base

The oxide, hydroxide or carbonate of a metal that will react with an acid, forming a salt as one of the products. Bases are proton acceptors.

Blast Furnace

The huge reaction vessels used in industry to extract iron from its ore.

Bond Energy

The energy needed to break a particular chemical bond.

Burette

A long glass tube with a tap at one end and markings to show volumes of liquid, used to add precisely known amounts of liquids to a solution in a conical flask below it.

Calcium Carbonate

The main compound found in limestone. It is a white solid whose formula is CaCO_3 .

Calcium Hydroxide

A white solid made by reacting calcium oxide with water. It can be used as a cheap alkali in industry.

Calcium Oxide

A white solid made by heating limestone strongly, e.g. in a lime kiln.

Carbon Monoxide

A toxic gas whose formula is CO .

Carbon Steel

Alloy of iron containing controlled small amounts of carbon.

Cast Iron

The impure iron taken directly from a blast furnace.

Catalyst

A substance that speeds up a chemical reaction but remains chemically unchanged itself at the end of the reaction.

Catalytic Converter

Fitted to exhausts of vehicles to reduce pollutants released.

Collision Theory

An explanation of chemical reactions in terms of particles colliding with sufficient energy for a reaction to take place.

Compound

A substance made when two or more elements are chemically bonded together. E.g. water is made from hydrogen and oxygen (H_2O).

Copper Rich Ore

Rock that contains a high proportion of a copper compound.

Core

The centre of the Earth.

Covalent Bond

The attraction between 2 atoms that share one or more pairs of electrons.

Covalent Bonding

The attraction between 2 atoms that share one or more pairs of electrons.

Cracking

The reaction used in the oil industry to break down large hydrocarbons into smaller, more useful ones. This occurs when the hydrocarbon vapour is either passed over a hot catalyst or mixed with steam and heating to a high temperature.

Crust

The outer solid layer of the Earth.

Delocalised Electron

Bonding electron that is no longer associated with any one particular atom.

Directly Proportional

A relationship that, when drawn on a line graph, shows a positive linear relationship that crosses through the origin.

Displace

When one element takes the place of another in a compound. E.g. iron and copper sulfate --> copper and iron sulfate.

Distillation

Separation of a liquid from a mixture by evaporation followed by condensation.

Electrolysis

The breakdown of a substance containing ions by electricity.

Electrolyte

A liquid, containing free moving ions, that is broken down by electricity in the process of electrolysis.

Electron

A tiny particle with a negative charge. Electrons orbit the nucleus in atoms or ions.

Electronic structure

A set of numbers to show the arrangement of electrons in their shells. E.g. Potassiums structure is 2.8.8.1.

Electrolysis

the passage of an electric current through an electrolyte with subsequent migration of positively and negatively charged ions to the negative and positive electrodes.

Electroplating

The process of depositing a thin layer of metal on an object during electrolysis.

Element

A substance made up of only 1 type of atom. An element cannot be broken down chemically into any simpler substances.

Empirical Formula

The simplest ratio of elements in a compound.

Endothermic

A reaction that takes in energy from the surroundings.

End point

The point in a titration where the reaction is complete and titration should stop.

Equilibrium

The point in a reversible reaction in which the forward and backwards rates of reaction are the same. Therefore the amounts of substances present in the reacting mixture remain constant.

Ethene

An alkene with the formula C_2H_4 .

Evidence

Data which has been shown to be valid.

Exothermic

A reaction that gives out energy into the surroundings.

Fermentation

The reaction in which the enzymes in yeast turn glucose into ethanol and carbon dioxide.

Flammable

Easily ignited and capable of burning readily.

Fraction

Hydrocarbons with similar boiling points separated from crude oil.

Fractional Distillation

Separation of many substances in a mixture by heating, using the fact that they have different boiling points.

Functional Group

An atom or group of atoms that give organic compounds their characteristic reactions.

Gas

A state of matter with no definite shape or volume

Group

All the elements in each column (labelled 1 to 7 and 0) down the periodic table.

Half Equation

An equation involving either an oxidation or a reduction process. The electrons lost or gained in this process are shown in the equation.

Hardening

The process of reacting plant oils with hydrogen to raise their melting point. This is used to make spreadable margarine.

Hard Water

water that contains salts (as calcium and magnesium ions) that limit the formation of lather with soap.

Hazard

a source of danger

High Alloy Steel

Expensive alloy of iron mixed with relatively large proportions of other metals e.g. stainless steel which contains nickel and chromium.

Homologous Series

Series of organic compounds having the same functional group but with each successive member differing by CH_2 .

Hydrated

Crystalline and containing water molecules.

Hydration

the process of combining with water.

Hydrocarbon

an organic compound containing only carbon and hydrogen.

Hydrogenated Oil

hydrogen is added to vegetable oils to change the oil from liquid to solid.

Hydrophilic

attracts water molecules.

Hydrophobic

water hating.

Hypothesis

a proposal intended to explain certain facts or observations.

Incomplete Combustion

Combustion in which not enough oxygen is supplied to completely burn the fuel. Carbon monoxide is a common product.

Inert

Unreactive.

Intermolecular Forces

the forces of attraction between molecules.

Ion

a particle that is electrically charged (positive or negative).

Ionic Bond

The electrostatic force of attraction between oppositely charged ions.

Ionic Bonding

The electrostatic force of attraction between oppositely charged ions.

Isotope

one of two or more atoms with the same atomic number but with different numbers of neutrons

Limewater

A solution of calcium hydroxide, used to test for carbon dioxide.

Liquid

a substance in the fluid state of matter having no fixed shape but a fixed volume.

Low Alloy Steel

an alloy of iron and carbon, with other elements added for increased strength.

Macromolecule

any very large complex molecule.

Mass Number

the sum of the number of neutrons and protons in an atomic nucleus.

Mixture

(chemistry) a substance consisting of two or more substances mixed together (not in fixed proportions and not with chemical bonding).

Mole

The amount of substance in the relative atomic or formula mass of a substance in grams.

Molecular Formula

a chemical formula that shows the number and kinds of atoms in a molecule, but not the arrangement of the atoms.

Molecule

two or more atoms held together by covalent bonds.

Monitor

To watch closely and frequently; to observe and make notes.

Monomers

small unit that can join together with other small units to form polymers.

Neutral

pH 7.

Neutralisation

A chemical reaction in which an acid and a base interact with the formation of a salt.

Neutron

A subatomic particle that has no charge and that is found in the nucleus of an atom.

Nitrogen Dioxide

a foul-smelling reddish brown gas, produced primarily by the incomplete combustion of fuel, that contributes to smog and acidic deposition.

Non Renewable

A resource that cannot be remade quickly or cannot be remade at all.

Nucleus

the positively charged dense center of an atom.

Ore

a mineral or rock that contains a useful substance that can be mined for profit.

Oxidation

A chemical reaction that involves the loss of an electron. Oxidation often involves the addition of oxygen and the loss of hydrogen ions.

Percentage Yield

The ratio of the actual yield to the theoretical yield, multiplied by 100.

Periodic Table

A table that shows the elements, their atomic number, symbol, and average atomic mass; elements with similar chemical properties are grouped together.

pH Scale

a range of numbers used to describe how acidic or basic a solution is; ranges from 0 (most acidic) to 14 (most basic).

Pipette

A glass or transparent plastic tube used to accurately measure small amounts of liquid.

Polymer

large molecule formed when many smaller molecules bond together.

Polymerisation

a chemical reaction joining monomers in long chains to form a polymer.

Precipitate

a solid that forms and settles out of a liquid mixture.

Precise

sharply defined and exact, reduces random error.

Precision

the quality of being reproducible in amount or performance.

Prediction

a statement made in advance that expresses the results that will be obtained from testing a hypothesis if the hypothesis is supported; the expected outcome if a hypothesis is accurate.

Product

a chemical substance formed as a result of a chemical reaction.

Propene

C_3H_6 .

Reactant

a chemical substance that is present at the start of a chemical reaction.

Reactivity Series

The group one metals series reactivity are as follows: the most reactive metals are at the bottom of group 1, and the less reactive at the top (e.g 1) potassium 2) Sodium 3) Lithium 4) Calcium 5) Magnesium 6) Zinc 7) Iron/lead 8) Copper/Silver/Gold).

Reduction

any process in which electrons are added to an atom or ion (as by removing oxygen or adding hydrogen).

Relationship

how one thing is connected to another.

Relative Atomic Mass

The weighted mean of the relative isotopic masses of an element on the scale where carbon-12 is exactly 12.

Relative Formula Mass

The sum of the relative atomic masses of the elements as given in the formula for any non-molecular compound.

Reversible Reaction

A chemical reaction in which the products re-form the original reactants.

Salt

a compound formed by replacing hydrogen in an acid by a metal (or a radical that acts like a metal).

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Saturated Hydrocarbon

A hydrocarbon in which all the bonds between carbon atoms are single bonds.

Shell

a group of electrons in an atom that orbit at roughly the same distance from the nucleus.

Solid

the state in which a substance has no tendency to flow under moderate stress.

State Symbol

Symbol used in equations to show whether something is solid, liquid, gas or in solution in water.

Symbol Equation

an equation where the chemical symbols are put in place and not used by words.

Thermal Decomposition

The breaking up of a chemical substance with heat into at least two chemical substances.

Titration

process in which a solution of known concentration is used to determine the concentration of another solution.

Transition Element

elements in Groups 3 through 12 of the periodic table; occur in nature as uncombined elements and include the iron triad and coinage metals.

Transition Metal

An Element in group 3 through 12 of the periodic table.

Universal Indicator

An indicator with a different colour for each pH value.

Unsaturated Hydrocarbon

hydrocarbon with one or more double or triple bonds.

Unsaturated Oil

has one or more double bonds between the carbons formed by the removal of an hydrogen atom from the carbon skeleton. It is liquid at room temperature.

Word Equation

aAn equation in which the reactants and products in a chemical reaction are represented by words.

Yield

An amount of a product.

- 1) Lawrie Ryan **AQA GCSE Chemistry 3rd edition**,
oxford University publisher, 2016.
- 2) *Edexcel* International **GCSE Chemistry, online edition**,
2009

