NATIONAL TALENT SEARCH EXAMINATION 2025-26

(NTSE- 2025-26) STAGE -1& 2

State : ALL STATES **Paper : SAT**

UNIQUE SUCCESS IAS VISION

Max Marks : 100 **Time allowed : 120 mins**

NEW BATCH START (10 –SEP- 2025)

SCIENCE (PHYSICS , CHEMISTRY , BIOLOGY)

- **BASIC TO ADVANCE LEVEL ALL CONCEPT COVER**
- **ACCORDING TO SYLLABUS**

As you can see, this part covers **three subjects –**

- **Science**
  - Physics
  - Chemistry
  - Biology

The syllabus for each of these subjects is tabulated below.

Subject	No. of Questions	Marks
1. Science	40	40
2. Mathematics	20	20
3. Social Science	40	40
Total	100	100

SUBJECTWISE QUESTION WEIGHTAGE

- Physics - 13 QUESTIONS
- Chemistry- 13 QUESTIONS
- Biology - 14 QUESTIONS

SUBJECTWISE QUESTION WEIGHTAGE

- Physics - 13 QUESTIONS

Physics Syllabus

1. Magnetic effects of Electric Current
2. Electricity
3. Measurements
4. Motion
5. Force& laws of motion
6. Gravitation
7. Work and Energy
8. Light – Reflection&Refraction
9. The human eye
10. Sound
11. The Universe
12. Source of Energy

SUBJECTWISE QUESTION WEIGHTAGE

- Physics - 13 QUESTIONS

Chemistry Syllabus

1. Chemical reaction
2. Acid, Bases, Salt
3. Periodic Classification of Elements
4. Metals and Non-Metals
5. Carbon and its Compounds
6. Matter in our surroundings
7. Is matter around us pure
8. atoms and molecules
9. Structure of Atom

SUBJECTWISE QUESTION WEIGHTAGE

- ❖ Physics - 14 QUESTIONS

Biology Syllabus

1. The fundamental unit of life
2. Tissue
3. Improvement in food resources
4. Why do we fall ill
5. Natural resources
6. Diversity in Living Organisms
7. Life Processes
8. control & coordination
9. How do organisms Reproduced
10. Heredity & Evolution
11. Our Environment& Some Common Diseases
12. Management of natural resources
13. Soil Micro-organisms
14. Plant and Animal Nutrition & Human Body

SUBJECTWISE QUESTION WEIGHTAGE

- Physics - 13 QUESTIONS

Chemistry Syllabus

Chemical reaction and equation

Index in chapter study

1. Chemical reaction
2. Chemical equation& example
3. characteristics of chemical reactions
   - A. evolution of a gas
   - B. formation of a precipitate
   - C. change in colour
   - D. change in temperature
   - E. change in state
4. Balanced & unbalanced chemical equation
5. Implications of a balanced chemical equation
6. Types of chemical reactions :

A. Combination reactions
B. Decomposition reactions
C. Displacement reactions
D. Double displacement reactions
E. Precipitation

F. Neutralisation reactions

G. Reversible reactions

H. Irreversible reactions

I. Oxidation reactions

J. Reduction reactions

K. Redox reaction

L. Effect of oxidation reaction in everyday life

Corrosion of metals

Rancidity of food

1.Chemical reaction

- Chemical reactions are the processes in which new substances with new properties are formed.
- During a chemical reaction, atoms of one element do not change into those of another element.
- Only a rearrangement of atoms takes place in a chemical reaction.

2.Chemical equation & example

Magnesium (As ribbon) + Oxygen (From air) → Magnesium oxide (White powder)

We will now discuss reactants and products of a chemical reaction.

(i) The substances which take part in a chemical reaction are called reactants.

(ii) The new substances produced as a result of chemical reaction are called products.

Represented by the equation

$$2Mg + O_2 \rightarrow 2Mg$$

Magnesium (As ribbon) + Oxygen (From air) → Magnesium oxide (White powder)

- ❖ **When a magnesium ribbon is heated, it burns in air with a dazzling white flame to form a white powder called magnesium oxide.**

Actually, on heating, magnesium combines with oxygen present in air to form magnesium oxide.

- ❖ **The burning of magnesium in air to form magnesium oxide is an example of a chemical reaction,**

In this chemical reaction

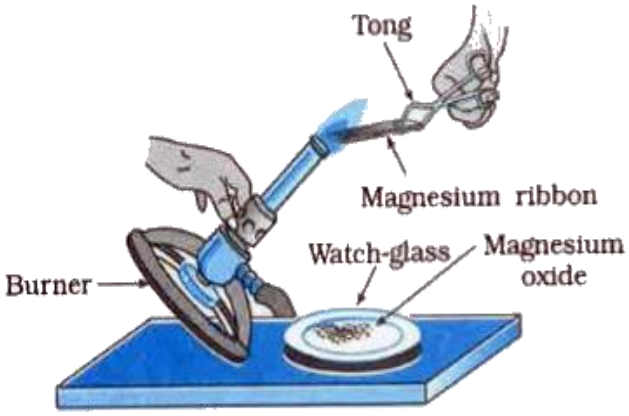
- ➢ **there are two reactants 'magnesium and oxygen**
- ➢ **but only one product 'magnesium oxide**

some important facts

The magnesium ribbon which we use usually has a coating of 'magnesium oxide' on its surface which is formed by the slow action of oxygen of air on it.

- ➢ **So, before burning in air, the magnesium ribbon is cleaned by rubbing with a sand paper.**
- ➢ **This is done to remove the protective layer of magnesium oxide from the surface of magnesium ribbon.**
- ➢ **so that it may readily combine with the oxygen of air on heating.**

Figure 1. When magnesium burns in air, it combines with the oxygen of air to form magnesium oxide.



Another point to be noted is that the dazzling (very bright) white light given out during the burning of magnesium ribbon is harmful to the eyes.

Hold the burning magnesium ribbon over a watch glass so that the magnesium oxide powder being formed collects in the watch glass [see Figure 1(a)].

A large number of chemical reactions keep on occurring in our daily life

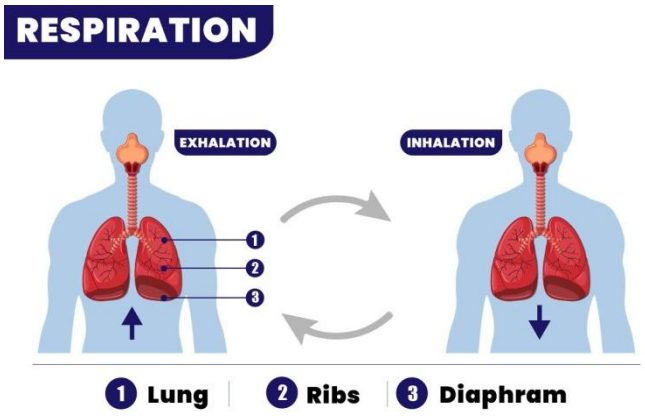
1. **Souring of milk (when left at room temperature during summer),**

2. **Formation of curd from milk,**

3. Cooking of food,

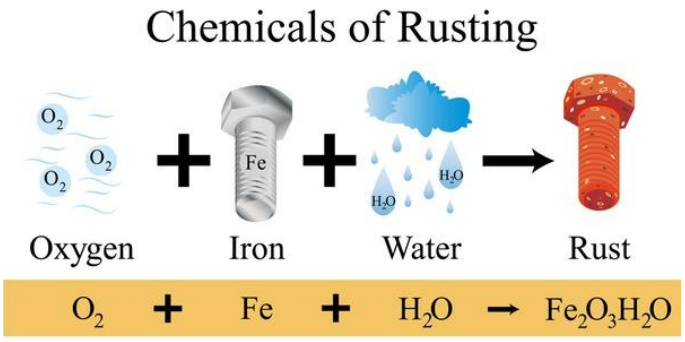
4. Digestion of food in our body,

5. Process of respiration,

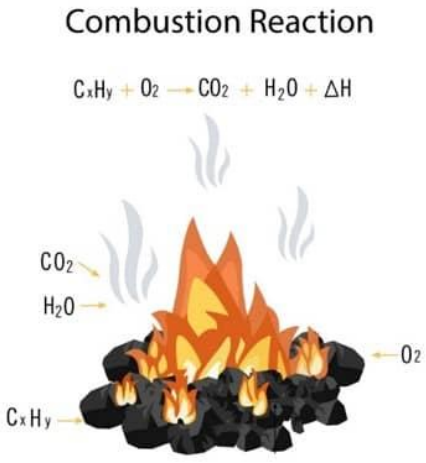


6. Fermentation of grapes,

7. Rusting of iron (when left exposed to humid atmosphere),



8. Burning of fuels (like wood, coal, kerosene, petrol and LPG)



9. Burning of candle wax

10. Ripening of fruits,

are all chemical changes which involve chemical reactions.

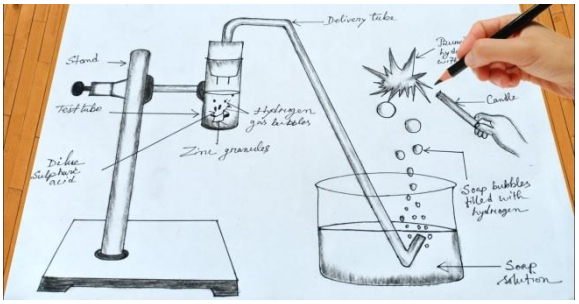
3.characteristics of chemical reactions

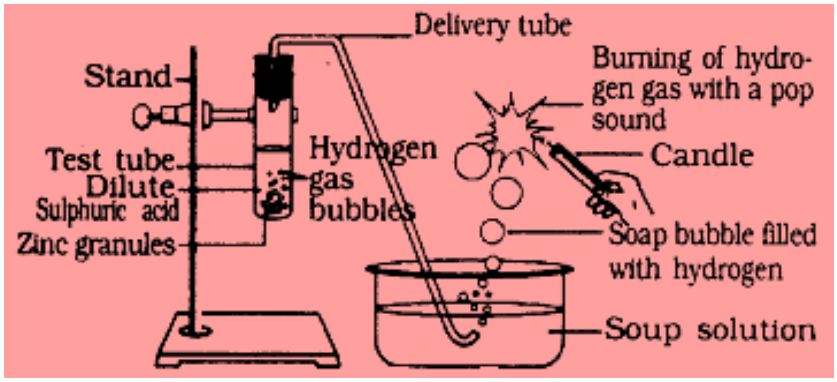
- In a chemical reaction, the substances known as reactants are converted into new substances called products.
- The conversion of reactants into products in a chemical reaction is often accompanied by some features which can be observed easily. The easily observable features (or changes) which take place as a result of chemical reaction are known as characteristics of chemical reactions .

1. evolution of a gas
2. formation of a precipitate
3. change in colour
4. change in temperature
5. change in state

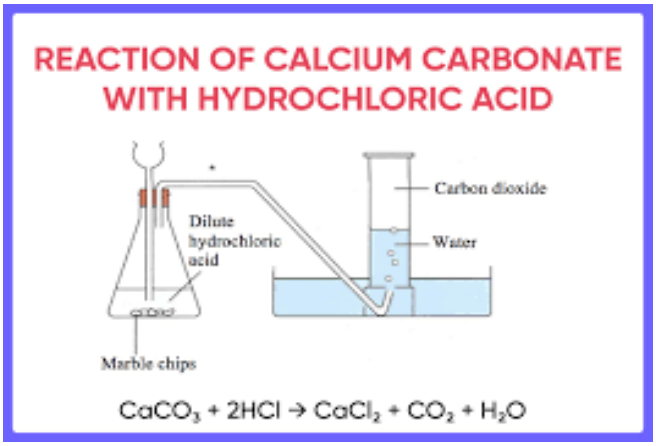
1.evolution of a gas

- when zinc granules react with dilute sulphuric acid, then bubbles of hydrogen gas are produced.
- So, the chemical reaction between zinc and dilute sulphuric acid is characterised by the evolution of hydrogen gas.





- **When magnesium reacts with a dilute acid (like dilute hydrochloric acid or dilute sulphuric acid), even then hydrogen gas is evolved**
- **When dilute hydrochloric acid is poured over sodium carbonate in a test-tube, then carbon dioxide gas is evolved.**



- **So, the chemical reaction between sodium carbonate and dilute hydrochloric acid is characterised by the evolution of carbon dioxide gas.**
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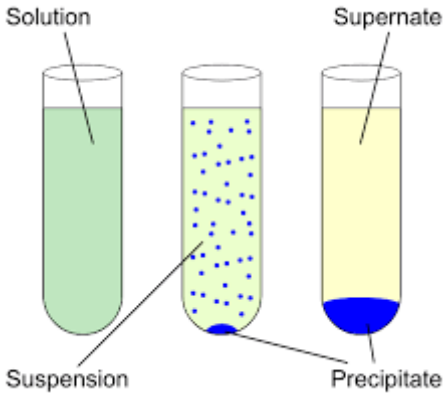
2.formation of a precipitate

- **A precipitate is a 'solid product' which separates out from the solution during a chemical reaction.**

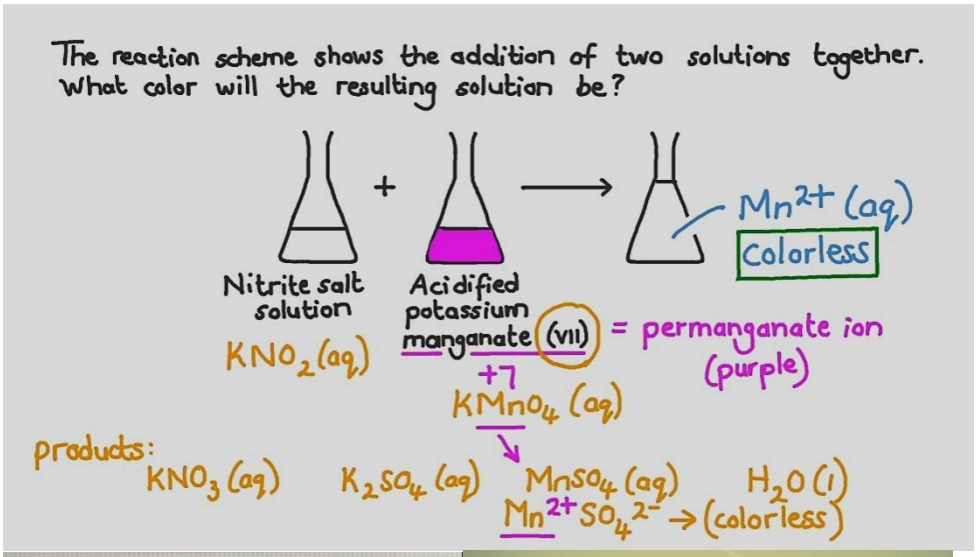


- **A precipitate can be formed by mixing aqueous solutions (water solutions) of reactants when one of the products is insoluble in water.**
- **A precipitate can also be formed by passing a gas into an aqueous solution of a substance (like passing carbon dioxide gas into lime water).**
- **the chemical reaction between potassium iodide and lead nitrate is characterised by the formation of a yellow precipitate of lead iodide.**

- **the chemical reaction between sulphuric acid and barium chloride solution is characterised by the formation of a white precipitate of barium sulphate.**

3.Change in colour

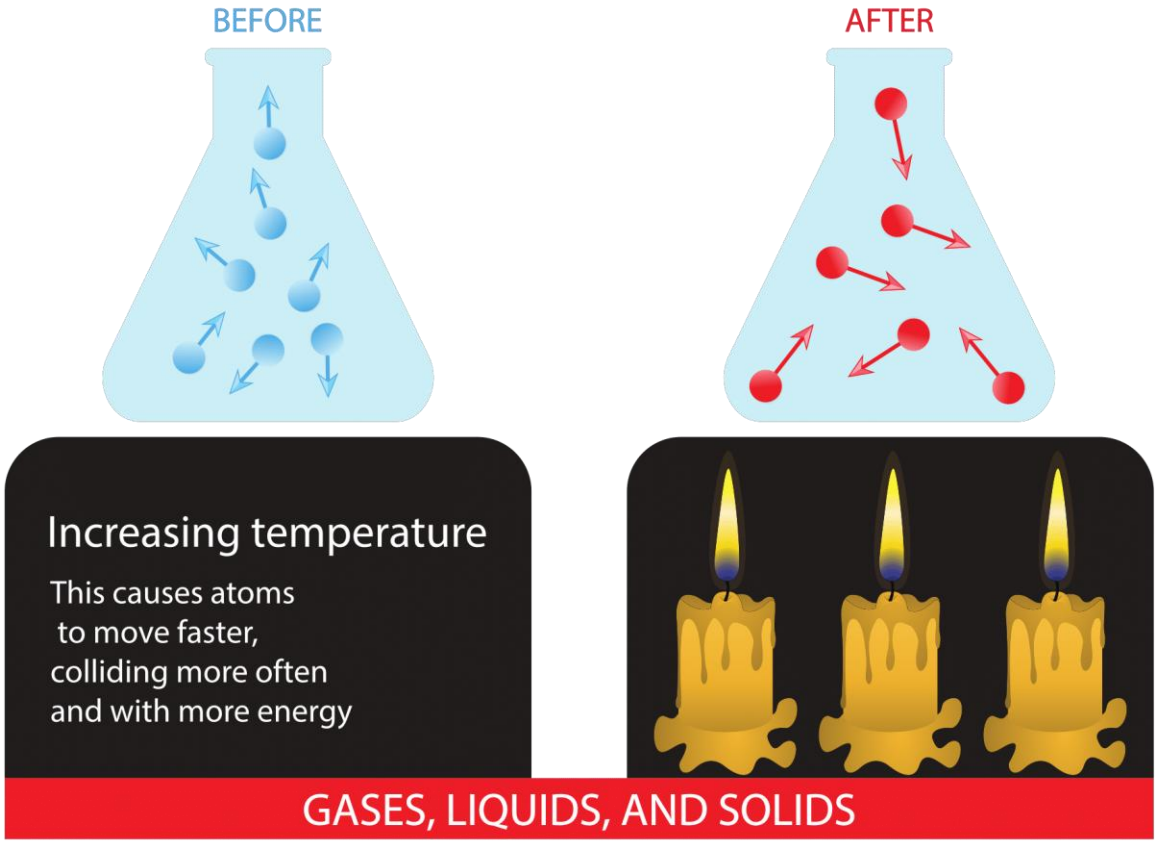
- **The chemical reaction between citric acid and purple coloured potassium permanganate solution is characterised by a change in colour from purple to colourless.**



- **The chemical reaction between sulphur dioxide gas and acidified potassium dichromate solution is characterised by a change in colour from orange to green.**

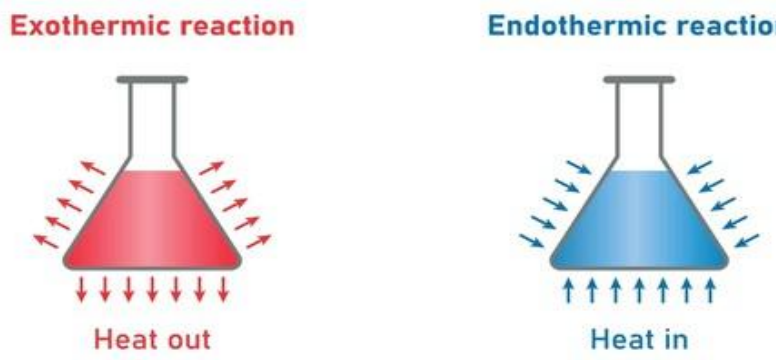
4.Change in temperature

- **We should know why temperature changes take place in chemical reactions.**
- **Chemical reactions often produce heat energy.**
- **When a chemical reaction produces heat energy,**
- **then the temperature of reaction mixture rises (or increases) and it becomes hot.**



- **exothermic reaction** releases energy into its surroundings, usually in the form of heat or light, causing the surroundings to warm up **is known as exothermic reaction.**

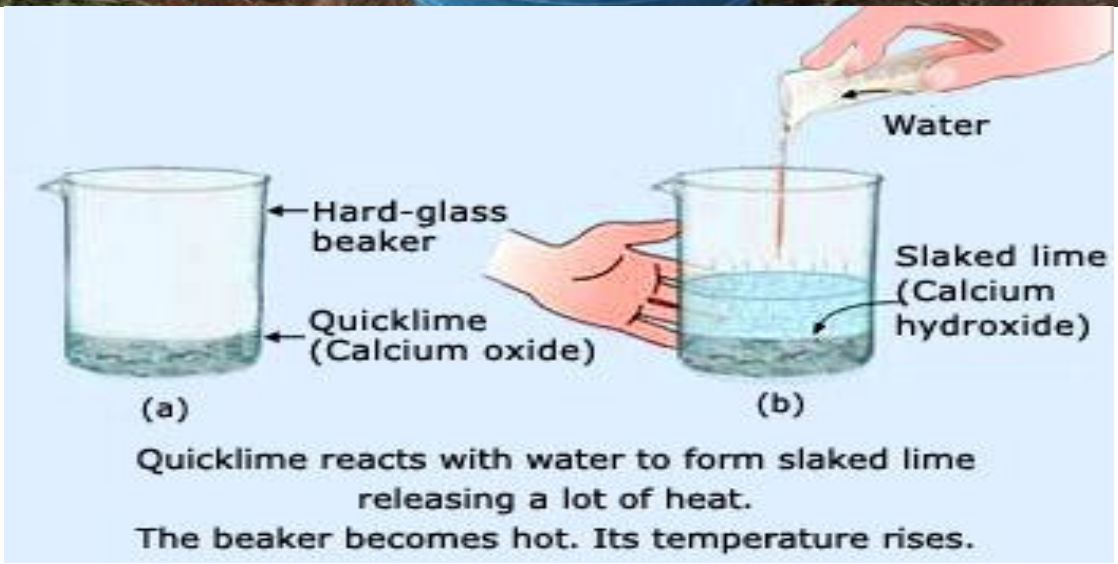
- **In some cases, however, chemical reactions absorb heat energy When a chemical reaction absorbs heat energy,**
- **then the temperature of reaction mixture falls (or decreases) and it becomes cold.**
- **endothermic reaction** chemical or physical process that absorbs energy (usually heat) from its surroundings, leading to a decrease in the surrounding temperature **is known as endothermic reaction.**



Example

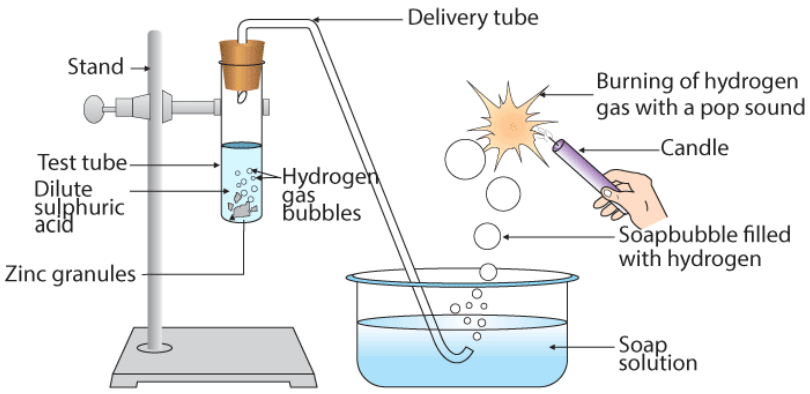
Another point to be noted is that the compound 'calciumoxide is known by two common names lime' as well as 'quicklime'.

And the compound 'calcium hydroxide' is known as 'slakedlime

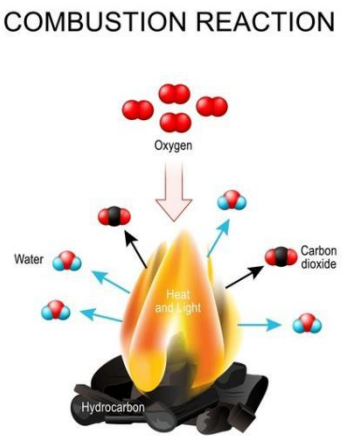


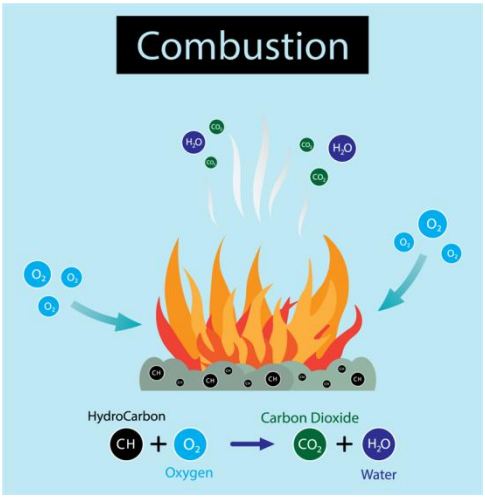
- **when quicklime reacts with water, then slaked lime is formed and a lot of heat energy is produced.**
- **This heat raises the temperature due to which the reaction mixture becomes hot.**
- **So, we can say that the chemical reaction between quicklime and water to form slaked lime is characterised by a change in temperature (which is rise in temperature).**
- **The reaction between quicklime and water to form slaked lime is an exothermic reaction (which means heat producing reaction).**

- **the chemical reaction between zinc granules and dilute sulphuric acid to produce hydrogen gas.**
- **If we touch the conical flask containing zinc granules and dilute sulphuric acid, it is found to be warm (which means that the temperature rises during this reaction).**



- **The chemical reaction in which carbon burns in air to form carbon dioxide also releases a lot of heat.**





- **a chemical reaction in which heat energy is absorbed due to which the temperature falls.**
- **When barium hydroxide [Ba(OH),] is added to ammonium chloride (NH,CI) taken in a test-tube and mixed with a glass rod, then barium chloride, ammonia and water are formed.**
- **A lot of heat energy is absorbed during this reaction due to which the temperature of reaction mixture falls and the bottom of test tube becomes very cold.**

- **Thus, the chemical reaction between barium hydroxide and ammonium chloride to form barium chloride, ammonia and water is characterised by a change in temperature (which is fall in temperature).**

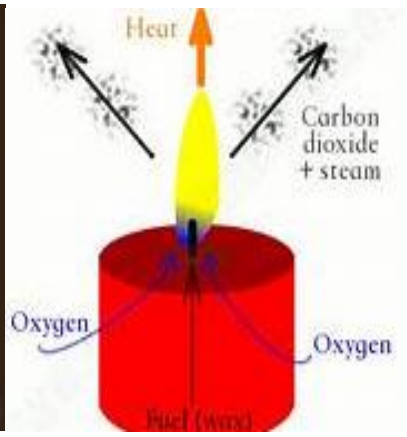
Chemical Equation

$$Ba(OH)_2 + NH_4Cl \longrightarrow Ba(Cl)_2 + H_2O + NH_3$$

- **It is an endothermic reaction (which means heat absorbing reaction).**

5.Change in state

- **when wax is burned (in the form of a wax candle), then water and carbon dioxide are formed.**



- **Now, wax is a solid, water is a liquid whereas carbon dioxide is a gas.**

- **This means that during the combustion reaction of wax, the physical state changes from solid to liquid and gas.**

- **the combustion reaction of candle wax is characterised by a change in state from solid to liquid and gas**

- because wax is a solid.
- water formed by the combustion of wax is a liquid at room temperature.
- whereas carbon dioxide produced by the combustion of wax is a gas.

Method Chemical Equations

1. Balanced chemical Equations

2. Unbalanced chemical Equations

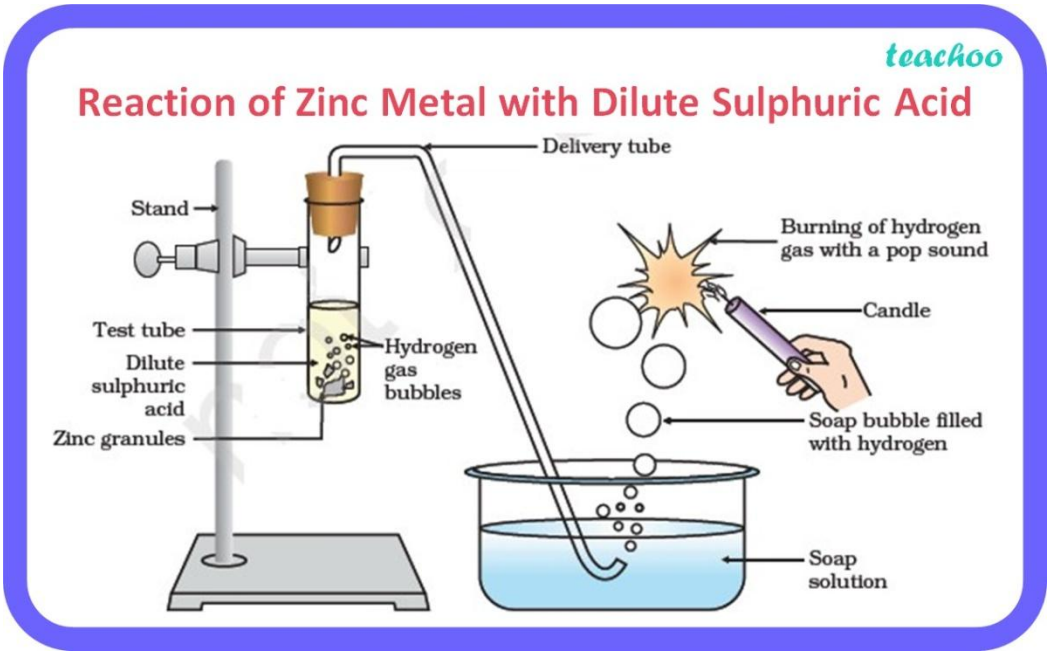
Method Chemical Equations

- The method of representing a chemical reaction with the help of symbols and formulae of the substances involved in it is known as a chemical equation.
- Zinc metal reacts with dilute sulphuric acid to form zinc sulphate and hydrogen gas.

Zinc+ Sulphuric acid ⟶ Zinc sulphate + Hydrogen

$$Zn + H_2SO_4 \longrightarrow ZnSo_4 + H_2$$

Reactants Products



- **The substances which combine or react are known as reactants.**
- **The reactants are always written on the left hand side in an equation with a plus sign(+) between them.**

- **The new substances produced in a reaction are known as products.**
- **The products are always written on the right hand side in an equation with a plus sign (+) between them.**
- **The arrow sign (→) pointing towards the right hand side is put between the reactants and products.**

1.Balanced chemical Equations

1. A balanced chemical equation has an equal number of atoms of different elements in the reactants and products.

- **A balanced equation has an equal number of atoms of the different elements on both the sides.**

- **Zinc reacts with dilute sulphuric acid to give zinc sulphate and hydrogen. This can be written in equation form as:**

$$Zn + H_2SO_4 \rightarrow ZnSO4 + H_2$$

➢ **Let us count the number of atoms of all the elements in the reactants and products separately.**

	In reactants	in products
No. of Zn atoms:	1	1
No. of H atoms:	2	2
No. of S atoms:	1	1
No. of O atoms:	4	4

➢ **there is an equal number of atoms of different elements in the reactants and products, so the above chemical equation is a balanced equation.**

2. Unbalanced chemical Equations

1. **An unbalanced chemical equation has an unequal number of atoms of one or more elements in the reactants and products.**

❖ **An unbalanced chemical equation has an unequal number of atoms of one or more elements on its two sides.**

Example

Hydrogen reacts with oxygen to form water.

$$H_2 + O_2 \rightarrow H_2O$$

- **Let us count the number of hydrogen atoms and oxygen atoms in the reactants as well as product**

	In reactant	In product
No. of H atoms:	**2**	**2**
No. of O atoms:	**2**	**1**

- **the number of atoms of various elements in reactants and products is unequal,**

According law of conservation law of mass

- **matter can neither be created nor destroyed in a chemical reaction.**
- **This means that the total mass of all the reactants must be equal to the total mass of the products.**

Notes

- **The chemical equations are balanced to satisfy the law of conservation of mass in chemical reactions.**

- **The reaction between hydrogen and oxygen to form water cannot be written as:**

$$H_2 + O \rightarrow H_2O$$

- because oxygen occurs in the form of O_2 molecules and not as atoms O.

Notes

- we should never change the formula of an element or a compound to balance an equation.

Notes

- The elements which exist as diatomic molecules are oxygen, O_2, hydrogen, H_2, nitrogen, N_2, fluorine, F_2, chlorine, Cl_2, bromine, Br_2 and iodine, I_2

Balancing of Chemical Equations

- The process of making the number of different types of atoms equal on both the sides of an equation is called balancing of equation,
- The simple equations are balanced by hit and trial method.

EXAMPLE

- Hydrogen burns in oxygen to form water.

$$H_2 + O_2 \rightarrow H_2O$$

Step-01

- In this reaction H_2 and O_2 are reactants
- Whereas H_2O is the product.
- Let us count the number of hydrogen atoms and oxygen atoms in reactants and product.

	In reactants	In product
No. of H atoms:	2	2
No. of O atoms:	2	1

The number of the hydrogen atoms equal both sides (2 each) , but the number of oxygen atoms on the left side but only 1 oxygen atom on right side .To have 2 oxygen atoms on the right side, we multiply H_2O by 2 and write $2H_2O$ so that

$$H_2 + O_2 \longrightarrow 2H_2O$$

Step-02

Let us count the number of various atoms on both the sides again

	In reactants	In product
No. of H atoms:	2	2
No. of O atoms:	2	4

Though the number of oxygen atoms has become equal (2 on both sides),

but the number of hydrogen atoms has now become unequal.

There are 2 hydrogen atoms on the left side but 4 hydrogen atoms on the right side.

To have 4 hydrogen atoms on the left side, we multiply H_2 by 2 and write $2H_2$, so that:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Step- 03

Let us count the number of various atoms on both the sides again

	In reactants	In product
No. of H atoms:	4	4
No. of O atoms:	2	2

This chemical equation contains an equal number of atoms of hydrogen and oxygen on both the sides, so this is a balanced equation.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

To Make Equations More Informative

The chemical equations can be made more informative in three ways:

1. **By indicating the "physical states" of the reactants and products.**
2. **By indicating the "heat changes" taking place in the reaction.**

3. By indicating the "conditions" under which the reaction takes place.

❖ We will discuss these three points in detail one by one

To Indicate the Physical States of Reactants and Products in an Equation.

❖ There can be four physical states for the reactants and products of a chemical reaction:
  - ❖ Solid
  - ❖ Liquid
  - ❖ Aqueous solution
  - ❖ Gas

➢ Solid state is indicated by the symbol (s)
➢ Liquid state is indicated by the symbol (L)
➢ Aqueous solution (solution made in water) is indicated by the symbol (aq)
➢ Gaseous state is indicated by the symbol (g)

❖ Zinc metal reacts with dilute sulphuric acid to form zinc sulphate solution and hydrogen gas. This can be written as:

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

Here,

$$Zn\,(s) + H_2SO_4\,(aq) \rightarrow ZnSO_4\,(aq) + H_2\,(g)$$

❖ In some cases an insoluble product (called precipitate)

❖ when calcium hydroxide solution (lime water) reacts with carbon dioxide gas, a white precipitate of calcium carbonate is formed along with water.

Examples

$$Ca(OH)_2\ (aq) + CO_2\ (g) \rightarrow CaCO_3\ (s) + H_2O\ (L)$$

Calcium hydroxide (Lime water) | Carbondioxide | Calciumcarbonate (White ppt.) | water

- calcium carbonate is formed as a solid product (precipitate)
- "precipitate' is written in short form as 'ppt

To indicate the heat change in an equation

There are two types of reactions on the basis of hear change involveds .

1. Exothermic reactions and
2. Endothermic reactions.

Exothermic reactions

- those reactions in which heat is evolved are known as exothermic reactions.

  when carbon burns in oxygen to form carbon dioxide, a lot of heat is produced in this reaction:

Example

$$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g) + Heat$$

carbon | oxygen | carbon dioxide

- The burning of carbon in oxygen is an exothermic reaction because heat is evolved in this reaction.
- An exothermic reaction is indicated by writing "+ Heat" or "+ Heat energy" or just "+ Energy" on the products' side of an equation (as shown in the above equation).
- (Natural gas is mainly methane (CH4). When natural gas burns in the oxygen of air, it forms carbon dioxide and

water vapour. A large amount of heat energy is also produced.

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O + \text{Heat energy}$$

Methane (Natural gas) Oxygen (From air) Carbon dioxide Water

- The burning of natural gas is an exothermic reaction become best in produced in this reaction
- Note : all the combustion reactions are exothermic reactions.
- combustion of fuels such as wood, coal, kerosene, petrol and diesel, are all exothermic reactions (because all these reactions produce heat energy) .
- the combustion of food (like glucose) in our body during respiration is an exothermic reaction.
- This glucose then undergoes slow combustion by combining with oxygen in the cells of our body to produce energy in a process called respiration, this energy maintains our body heat.
- During respiration, glucose combines with oxygen in the cells of our body to form carbon dioxide and water along with the production of energy:

$$C_6H_{12}O_6\,(aq) + 6O_2\,(g) \rightarrow 6CO_2\,(g) + 6H_2O\,(L) + \text{Energy}$$

Glucose Oxygen Carbondioxide Water

- Respiration is an exothermic process because energy is produced during this process .
- The burning of a magnesium wire in air to form magnesium oxide is an exothermic reaction

- **The decomposition of vegetable matter into compost is also an example of exothermic process (because heat energy is evolved during this process).**

Endothermic reactions

- **Those reactions in which heat is absorbed are known as endothermic reactions.**

when nitrogen and oxygen are heated to a very high temperature (of about 3000°C) they combine to form nitrogen monoxide, and a lot of heat is absorbed in this reaction:

$$N_2\ (g) + O_2\ (g) + \text{Heat} \rightarrow 2NO\ (g)$$

Nitrogen Oxygen Nitrogen monoxide

The reaction between nitrogen and oxygen to form nitrogen monoxide is an endothermic reaction

An endothermic reaction is usually indicated by written

"+ Heat" or "+ Heat energy" or just "+ Energy" on the reactants' side of an equation .

- **All the decomposition reactions require energy (in the form of heat, light or electricity) to take place. So, all the decomposition reactions are endothermic reactions.**
- **when calcium carbonate is heated, it decomposes to form calcium oxide and carbon dioxide:**

$$CaCO_3\ (s) + \text{Heat} \rightarrow CaO\ (s) + CO2\ (g)$$

Calcium carbonate Calciumoxide Carbon dioxide

- **The decomposition of calcium carbonate is an endothermic reaction because heat energy is absorbed in this reaction.**

- Photosynthesis is an endothermic reaction.

- This is because sunlight energy is absorbed during the process of photosynthesis by green plants. this is because sunlight is absorbed during the photosynthesis reaction.
- The electrolysis of water to form hydrogen and oxygen is also an endothermic reaction. This is because electric energy is absorbed during this reaction.

- absorbed in chemical reactions in the form of heat, light or electricity.

To Indicate the Conditions Under Which the Reaction Takes Place.

If heat is required for a reaction to take place, then the heat sign delta (Δ) is put over the arrow of the equation.

- If the reaction takes place in the presence of a catalyst
- then the symbol or formula of the catalyst is also written above or below the arrow sign in the equation. This will become more clear from the following example.
- When potassium chlorate ($KClO_3$) is heated in the presence of manganese dioxide catalyst, it decomposes to form potassium chloride and oxygen gas.

$$2KClO_3\,(s) \rightarrow KCl\,(s) + 3O_2\,(g)$$

Potassium chlorate Potassium chloride Oxygen

- Here delta (Δ) stands for heat and MnO, is the catalyst
- The conditions of temperature and pressure at which the reaction takes place can also be indicated in an equation.

- Methanol (or Methyl alcohol) is manufactured from carbon monoxide and hydrogen.
- The mixture of carbon monoxide and hydrogen gases is compressed to 300 atmospheric pressure and then passed

over a catalyst consisting of a mixture of zinc oxide and chromium oxide heated to a temperature of 300°C.

- ❖ the conditions for this reaction to take place are: a pressure of 300 atmospheres (written as 300 atm), a temperature of 300°C .

a chemical equation for the reaction involved in the production of methanol along with conditions as follows:

Example

$$CO\ (g) + 2H_2\ (g) \rightarrow CH_3OH\ (l)$$

Carbon monoxide — hydrogen — methanol (methyl alcohol)

The green plants make food by photosynthesis.

During photosynthesis, carbon dioxide combines with water in the presence of 'sunlight' and the green pigment of leaves called 'chlorophyll' to make food like glucose and oxygen gas is given out.

$$6CO_2\ (g) + 6H_2O\ (l) \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} C_6H_{12}O_6\ (aq) + 602\ (g)$$

Carbon dioxide — Water — Glucose — Oxygen

Important Examples on Writing of Balanced Chemical Equations

We should remember the following four steps for writing equations for the chemical reactions:

- ➢ **First step:** Write down the chemical reaction in the form of a word equation, keeping the reactants on the left side and products on the right side.

- **Second step: Put the symbols and formulae of all the reactants and products in the word equation.**

- **Third step: Balance the equation by multiplying the symbols and formulae by the smallest possible figures (Do not change the formulae to balance the equation).**

- **Fourth step: If possible, make the equation more informative by indicating the physical states of reactants and products; by indicating the heat changes, if any, taking place in the reaction; and by indicating the conditions under which the reaction takes place. If, however, you do not have sufficient information regarding the physical states; heat changes and conditions of the reaction, this step may be avoided.**

TYPES OF CHEMICAL REACTIONS

- **Some of the important types of chemical reactions are:**

  - **Combination reactions,**
  - **Decomposition reactions,**
  - **Displacement reactions**
  - **Double displacement reactions, and**
  - **Reversible and Irreversible reactions**
  - **Oxidation and Reduction reactions.**

- **We will now discuss all these reactions in detail, one by one. Let us start with the combination reactions.**

Combination reactions

- **Those reactions in which two or more substances combine to form a single substance, are called combination reactions.**
- **An element and a compound can combine to form a new compound.**

Example 1. Magnesium and oxygen combine, when heated, to form magnesium oxide:

$$\underset{\text{Magnesium}}{2Mg\ (s)} + \underset{\text{Oxygen}}{O_2\ (g)} \xrightarrow{\text{Combination}} \underset{\text{Magnesium oxide}}{2MgO\ (s)}$$

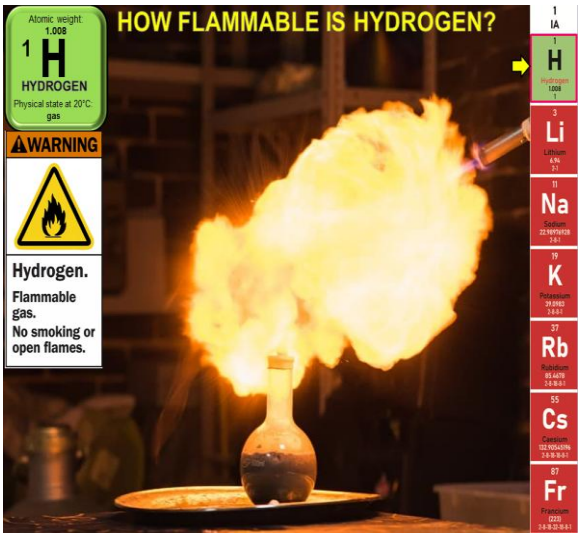
Example 2.Hydrogen burns in oxygen to form water:

$$\underset{\text{Hydrogen}}{2H_2\ (g)} + \underset{\text{Oxygen}}{O_2\ (g)} \xrightarrow{\text{Combination}} \underset{\text{water}}{2H_2O\ (l)}$$

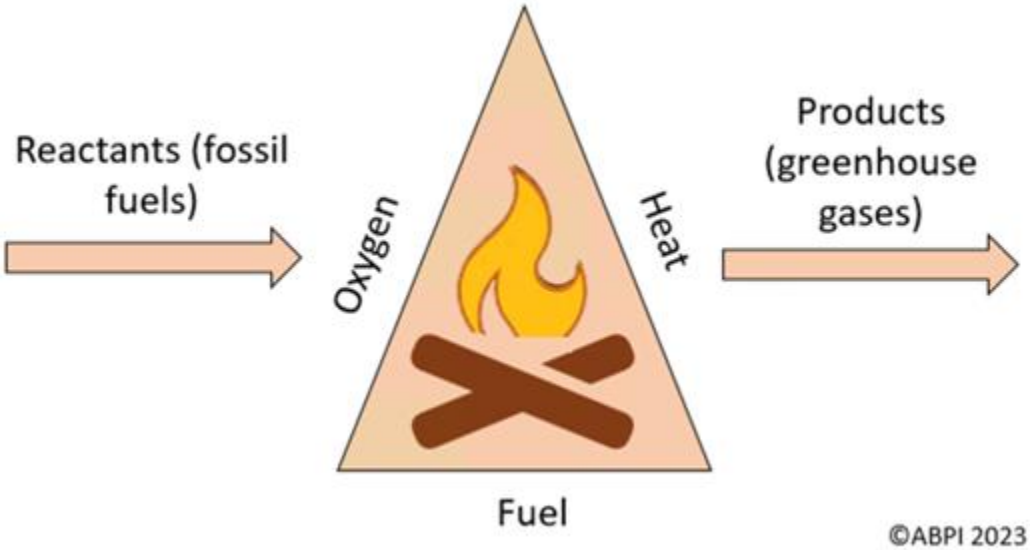
The formation of water from hydrogen and oxygen is a combination reaction.

Facts

1.Hydrogen is an element. It is a gas which burns explosively.



2.Oxygen is another element. It is a gas which does not burn itself but helps other things to burn.

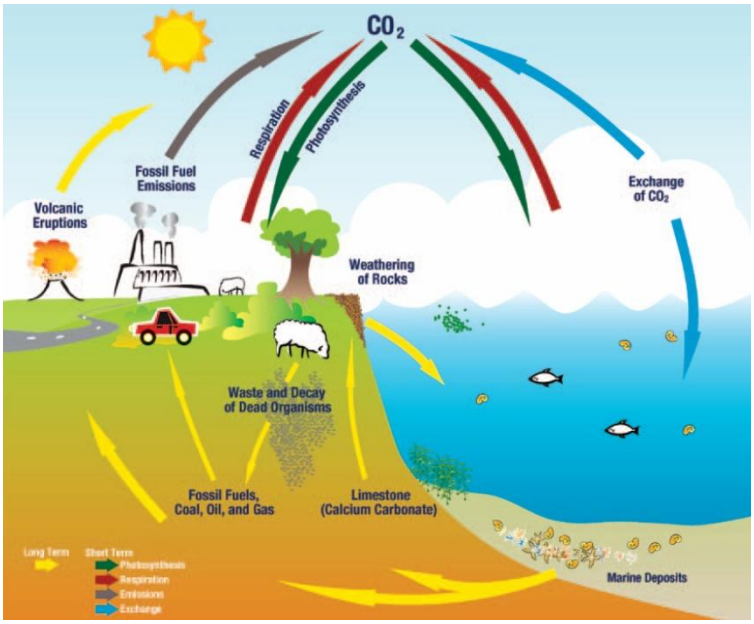


- **Hydrogen and oxygen combine to form a compound "hydrogen oxide' which is commonly known as water.it puts out fire.**

- 

Example 3. **Carbon (coal) burns in air to form carbon dioxide:**

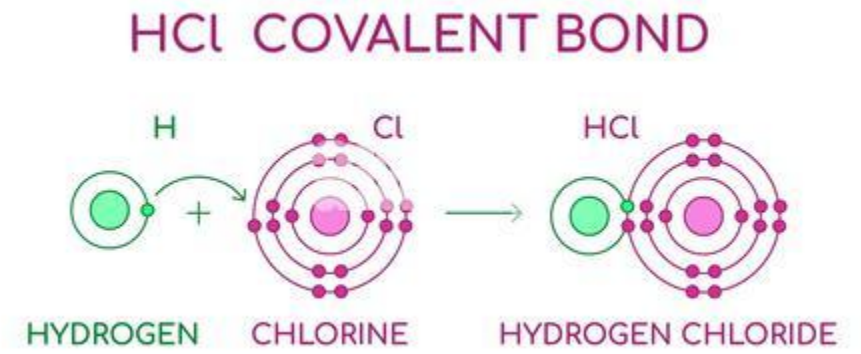
C (s)	**+**	**O_2 (g)**	Combination →	**CO2 (g)**
Carbon (Coal)		Oxygen (From air)		Carbon dioxide



Example 4. **Hydrogen combines with chlorine to form hydrogen chloride**

$H_2\,(g)$ + $Cl_2\,(g)$ $\xrightarrow{\text{Combination}}$ $2HCl\,(g)$

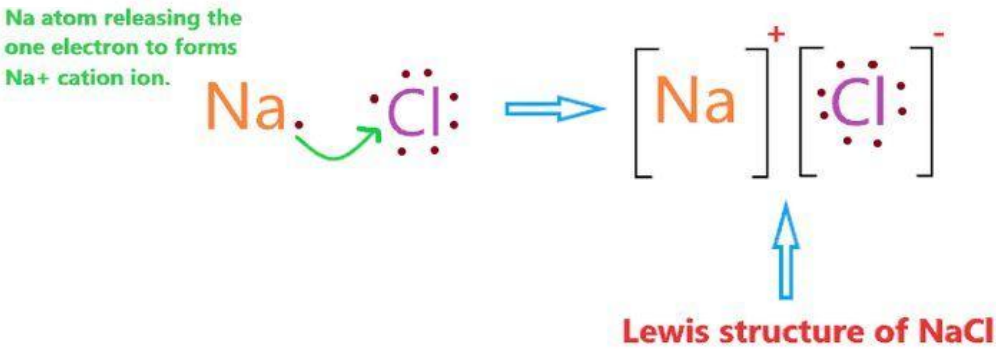
Hydrogen — Chlorine — Hydrogen chloride



Example 5. **Sodium metal burns in chlorine to form sodium chloride.**

$2Na\,(s)$ + $Cl_2\,(g)$ $\xrightarrow{\text{Combination}}$ $2NaCl\,(s)$

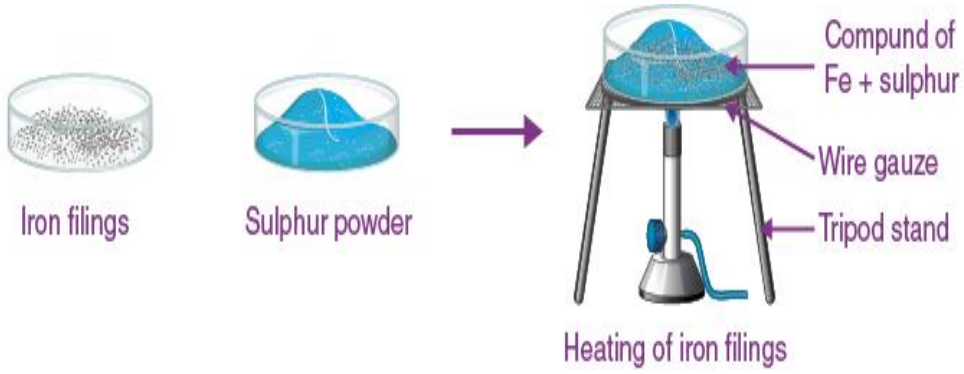
Sodium — Chlorine — Sodium chloride



Example 6. **When iron powder is heated with sulphur, iron sulphide is formed:**

$Fe\,(s)$ + $S\,(s)$ $\xrightarrow{\text{Combination}}$ $FeS\,(s)$

Iron — Sulphur — Iron sulphide



Heating of iron filings

- **two elements combine to form a single compound.**
- **two or more compounds combine together to form a new compound.**

Example 7. (Calcium oxide (lime or quicklime) reacts vigorously with water to form calcium hydroxide (slaked lime)

$$CaO\ (s) + \underset{\text{water}}{H_2O\ (l)} \xrightarrow{\text{Combination}} \underset{\text{Calcium hydroxide (Slaked lime)}}{Ca(OH)_2\ (s)}$$

- **A large amount of heat is released when calcium oxide reacts with water to form calcium hydroxide (or slaked lime.**
- **it is solid calcium hydroxide which is known as slaked lime.**
- **Slaked lime is a white powder.**
- **Discussion. The substance which we use for white-washing our house is lime (or quicklime) which calcium oxide (CaO).**

- **We put calcium oxide in a drum and add water to it slowly.**

- **Calcium oxide reacts with water vigorously to form a white solid called calcium hydroxide (or slaked lime) with the evolution of heat.**

- **More water is then added to get calcium hydroxide solution. This calcium hydroxide solution is then applied to the walls of the house with a brush.**

- **"The calcium hydroxide solution, when applied to the walls, reacts slowly with the carbon dioxide gas present in air to form a thin, shining layer of calcium carbonate on the walls of the house:**

$$\underset{\text{Calcium Hydroxide}}{Ca(OH)_2\,(aq)} + \underset{\substack{\text{Carbon dioxide}\\\text{(From air)}}}{CO_2\,(g)} \longrightarrow \underset{\text{Calcium carbonate}}{CacO_3\,(s)} + \underset{\text{Water}}{H_2O\,(l)}$$

- **Since this process gives a white, shiny appearance to the walls of a house, it is called white-washing. The calcium carbonate is actually formed after two to three days of white-washing and gives a shiny finish to the walls.**
- **Figure 28. Calcium oxide reacts vigorously with water to form calcium hydroxide.**

- **Figure 29. Calcium hydroxide solution reacts with carbon dioxide gas to form a white solid, calcium carbonate.**
- **Figure 30. Ammonia combines with hydrochloric acid to form dense white fumes of ammonium chloride.**

Example 8. Ammonia reacts with hydrogen chloride to form ammonium chloride. This can be written as:

$$\underset{\text{Ammonia}}{NH_3\,(g)} + \underset{\text{Hydrogen chloride}}{HCl\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Ammonium chloride}}{NH_4Cl\,(s)}$$

Example 9.Carbon monoxide reacts with oxygen to form carbon.

$$\underset{\text{Carbon monoxide}}{2CO\,(g)} + \underset{\text{Oxygen}}{O_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Carbon dioxide}}{2CO_2\,(g)}$$

Example 10. sulphur trioxide. This reaction can be written as:

$$2SO_2\,(g) + \underset{\text{Oxygen}}{O_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Sulphur trioxide}}{2SO_3\,(g)}$$

DECOMPOSITION REACTIONS

- **Those reactions in which a compound splits up into two or more simpler substances are known as decomposition reactions.**

- **The decomposition reactions are carried out by applying heat, light or electricity. Heat, light or electricity provide energy which breaks a compound into two or more simpler compounds.**

- **Please note that a decomposition reaction is just the opposite of a combination reaction.**

Example 1. When calcium carbonate is heated, it decomposes to give calcium oxide and carbon dioxide:

$$\underset{\text{Calcium carbonate (limestone)}}{CaCO_3\,(s)} \xrightarrow{\text{Heat}} \underset{\text{Calcium oxide (Lime)}}{CaO\,(s)} + \underset{\text{Carbon dioxide}}{CO_2\,(g)}$$

- **Please note that calcium carbonate is also called 'limestone.**
- **Calcium oxide formed from it is called 'lime (or quicklime).**
- **Calcium oxide (or lime) is used on a large scale in the manufacture of cement and glass.)**
- **When a decomposition reaction is carried out by heating, it is called 'thermal decomposition. ("Thermal' means 'relating to heat').**
- **The decomposition of calcium carbonate into calcium oxide and carbon dioxide is an example of thermal decomposition (because it is carried out by heating).**

Example 2.When potassium chlorate is heated in the presence of manganese dioxide catalyst, it decomposes to give potassium chloride and oxygen:

$$\underset{\text{Potassium chlorate}}{2KClO_3\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Heat}} \underset{\text{Potassium chloride}}{2KCl\ (s)} + \underset{\text{Oxygen}}{3O_2\,(g)}$$

Example 3. When ferrous sulphate is heated strongly, it decomposes to form ferric oxide, sulphur dioxide and sulphur trioxide:

$$\underset{\text{Ferrous sulphate (Green colour)}}{2FeSO_4\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Heat}} \underset{\text{Ferric oxide (Brown colour)}}{Fe_2O_3\,(s)} \quad \underset{\text{Sulphur dioxide}}{SO_2\,(g)} + \underset{\text{Sulphur trioxide}}{SO_3}$$

- ❖ **Please note that ferrous sulphate is also known as iron**
  - **sulphate (or just iron sulphate)**
  - **ferric oxide is also known as iron**
  - **oxide.**
- ✓ **ferrous sulphate crystals**
- ✓ **ferrous sulphate heptahydrate $FeSO_4.7H_2O$.**

Example 4. When lead nitrate is heated strongly, it breaks down to form simpler substances like lead monoxide, nitrogen dioxide and oxygen. This can be written as:

Heat

$$2Pb(NO_3)_2\ (s) \xrightarrow{\text{(Decomposition)}} 2PbO\ (s) + 4NO_2 + O_2\ (g)$$

Lead nitrate (Colourless) | Lead monoxide (Yellow) | Nitrogen dioxide (Brown fumes) | Oxygen

- ✓ **some examples of those decomposition reactions which are carried out by using electricity.**
- ✓

Example 5. When electric current is passed through acidified water, it decomposes to give hydrogen gas and oxygen gas. This reaction can be represented as:

$$2H_2O\ (l) \xrightarrow[\text{(Decomposition)}]{\text{Electricity}} 2H_2\ (g) + O_2$$

water | Hydrogen | Oxygen

(Decomposition)

Example 6. When electric current is passed through molten sodium chloride, it decomposes give sodium metal and chlorine gas:

$$2NaCl\ (l) \xrightarrow[\text{(Decomposition)}]{\text{Electricity}} 2Na\ (s) + Cl_2\ (g)$$

Sodium oxide (Molten) | Sodium metal | Chlorine gas

Example 7. When electric current is passed through molten aluminium oxide, it decomposes 10 give aluminium metal and oxygen gas:

$$2Al_2O_3(l) \xrightarrow[\text{(Decomposition)}]{\text{(Decomposition) Electricity}} 4Al\ (l) + 3O2\ (g)$$

Aluminium oxide (Molten) | Aluminium metal | Oxygen gas

- **some decomposition reactions which are brought about by light energy**

Example 8. **(When silver chloride is exposed to light, it decomposes to form silver metal and chlorine gas:**

$$\underset{\text{Silver chloride (white)}}{2AgCl\ (s)} \xrightarrow[\text{(Decomposition)}]{\text{Light}} \underset{\text{Silver (Greyish white)}}{2Ag(s)} + \underset{\text{Chlorine (Yellowish green)}}{Cl_2}$$

- **The decomposition of silver chloride is caused by light (It may be sunlight or bulb light).**
- **This reaction is used in black and white photography.**

❖ **Silver bromide also behaves in the same way as silver chloride with light energy.**

❖ **Thus, when silver bromide is exposed to light, it decomposes to form silver metal and bromine vapour.**

$$\underset{\text{Silver bromide (Pale yellow)}}{2AgBr\ (s)} \xrightarrow[\text{(Decomposition)}]{\text{Light}} \underset{\text{Silver (Greyish white)}}{2Ag\ (s)} + \underset{\text{Bromine (Red-brown)}}{Br_2\ (g)}$$

❖ **The light may be sunlight or bulb light.**

❖ This reaction of decomposition of silver bromide is also used in black and white photography.

- **Uses of Decomposition Reactions.** The decomposition reactions carried out by electricity are used to extract several metals from their naturally occurring compounds like chlorides or oxides.
- sodium chloride whereas aluminium metal is extracted by the electrolysis of molten aluminium oxide (see examples 6 and 7 given on page .

- **Decomposition Reactions in Our Body.** The digestion of food in the body is an example of decomposition reaction.

- When we eat foods like wheat, rice or potatoes, then the starch present in them decomposes to give simple sugars like glucose in the body; and the proteins decompose to form amino acids

DISPLACEMENT REACTIONS

- Those reactions in which one element takes the place of another element in a compound, are known as displacement reactions.
- In general, a more reactive element displaces a less reactive element from its compound.

Example 1. When a strip of zinc metal is placed in copper sulphate solution, then zinc sulphate solution and copper are obtained:

$$\underset{\text{Copper sulphate (Blue solution)}}{CuSO_4\ (aq)} + \underset{\text{Zinc (Silvery white)}}{Zn\ (s)} \longrightarrow \underset{\text{Zinc sulphate (Colourless solution)}}{ZnSO_4\ (aq)} + \underset{\text{Copper (Red brown)}}{Cu\ (s)}$$

- **this displacement reaction takes place because zinc is more reactive than copper.**

Example 2. When a piece of magnesium metal is placed in copper sulphate solution, then magnesium sulphate solution and copper metal are formed:

$$\underset{\text{Copper sulphate (Blue solution)}}{CuSO_4\,(aq)} + \underset{\text{Magnesium (silvery white)}}{Mg(s)} \longrightarrow \underset{\text{Magnesium sulphate (Colourless solution)}}{MgSO_4\ (aq)} + \underset{\text{Copper (Red – brown)}}{Cu\ (s)}$$

- **Here, magnesium is able to displace copper from copper sulphate solution because magnesium is more reactive than copper.**

Example 3. (When a piece of iron metal (say, an iron nail) is placed in copper sulphate solution then iron sulphate solution and copper metal are formed:

$$\underset{\text{Copper sulphate (Blue solution)}}{CuSO_4\,(aq)} + \underset{\text{Iron (Grey)}}{Fe\ (s)} \longrightarrow \underset{\text{Iron sulphat (Greenish solution)}}{FeSO_4(aq)} + \underset{\text{Copper (Red - brown)}}{Cu\ (s)}$$

- **this displacement reaction takes place because zinc is more reactive than copper.**

Example 4. **When a strip of lead metal is placed in a solution of copper chloride, then lead chloride solution and copper metal are formed:**

$CuCl_2\,(aq)$ (Copper chloride, Blue solution) + $Pb\,(s)$ (Lead, Bluish grey) ⟶ $PbCl_2\,(aq)$ (Lead chloride, Colourless solution) + $Cu\,(s)$ (Copper, Red - brown)

- **lead is able to displace copper from copper chloride solution because lead is more reactive than copper.)**
- **copper chloride ($CuCl_2$) used in this reaction is actually copper (II) chloride.**

Example 5. **When a copper strip is placed in a solution of silver nitrate, then copper nitrate solution and silver metal are formed:**

$2AgNO_3\,(aq)$ (Silver nitrate, Colourless solution) + $Cu\,(s)$ (Copper, Red - brown) ⟶ $Cu(NO_3)_2\,(aq)$ (Copper nitrate, Blue solution) + $2Ag\,(s)$ (Silver, Greyishwhite

- **This displacement reaction occurs because copper is more reactive than silver.**

Example 6. **Iron metal reacts with dilute hydrochloric acid to form iron (II) chloride and hydrogen.**

$Fe\,(s)$ (Iron, As iron filings) + $2HCl\,(aq)$ (Hydrochloric acid) ⟶ $FeCl_2\,(aq)$ (Iron (II) chloride, Ferrous chloride) + $H_2\,(g)$ (Hydrogen)

- **iron displaces hydrogen from hydrochloric acid solution to form hydrogen gas.**
- **his displacement reaction takes place because iron is more reactive than hydrogen.**
- **that iron metal is usually taken in the form of iron filings in this reaction.**

Example 7. Magnesium metal reacts with hydrochloric acid to form magnesium chloride and hydrogen.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Hydrochloric acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2\,(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

- **This displacement reaction occurs because magnesium is more reactive than hydrogen.**

Example 8. Sodium metal reacts with water to form sodium hydroxide solution and hydrogen gas.

$$\underset{\text{Sodium}}{2Na\,(s)} + \underset{\text{Water}}{2H_2O\,(l)} \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

- **This displacement reaction takes place because sodium is more reactive than hydrogen.**

Example 9. Chlorine gas reacts with potassium iodide solution to form potassium chloride.

$$\underset{\text{Chlorine}}{Cl2(g)} + \underset{\text{Potassium iodide}}{2kI\,(aq)} \longrightarrow \underset{\text{Potassium chloride}}{2KCl\,(aq)} + \underset{\text{Iodine}}{I2\,(s)}$$

- **This displacement reaction occurs because chlorine is more reactive than iodine.**

Example 10. When copper oxide is heated with magnesium powder, then magnesium oxide.

$$\underset{\text{Copper oxide}}{CuO\ (s)} + \underset{\text{Magnesium}}{Mg\ (s)} \longrightarrow \underset{\text{Magnesium oxide}}{MgO\ (s)} + \underset{\text{Copper}}{Cu\ (s)}$$

- **In this displacement reaction, a more reactive metal, magnesium is displacing a less reactive metal, copper, from its oxide, copper oxide.**

Example 11. When iron (III) oxide is heated with aluminium powder, then aluminium oxide and iron metal are formed:

$$\underset{\text{Iron (III) oxide (Ferric oxide)}}{Fe_2O_3\ (s)} + \underset{\text{Aluminium}}{2Al\ (s)} \longrightarrow \underset{\text{Aluminium oxide}}{Al_2O_3\ (s)} + \underset{\text{Iron (Molten)}}{2Fe\ (l)}$$

- **heat is produced in this reaction that iron is obtained in the molten state (liquid state).**
- **(single displacement reactions.**

❖ DOUBLE DISPLACEMENT REACTIONS

- **Those reactions in which two compounds react by an exchange of ions to form two ne compounds are called double displacement reactions.**
- **A double displacement reaction usually occurs in solution and one of the products, being insoluble, precipitates out (separates as a solid)**

Example 1. When silver nitrate solution is added to sodium chloride solution, then a white precipitate of silver chloride is formed along with sodium nitrate solution:

$$\underset{\text{Silver nitrate}}{AgNO_3\,(aq)} + \underset{\text{Sodium cloride}}{NaCl\,(aq)} \longrightarrow \underset{\text{Sodium cloride (white ppt.)}}{AgCl\,(s)} + \underset{\text{Sodium nitrate}}{NaNO_3\,(aq)}$$

Example 2. When barium chloride solution is added to sodium sulphate solution, then a white precipitate of barium sulphate is formed along with sodium chloride solution:

$$\underset{\text{Barium chloride}}{BaCl_2\,(aq)} + \underset{\text{Sodium sulphate}}{Na_2SO_4\,(aq)} \longrightarrow \underset{\text{Barium sulphate (white ppt.)}}{BaSO_4\,(s)} + \underset{\text{Sodium chloride}}{2NaCl\,(aq)}$$

- **Any reaction in which an insoluble solid (called precipitate) is formed that separates from t solution is called a precipitation reaction.**

Example 3. If barium chloride solution is added to copper sulphate solution, then a white precipitate of barium sulphate is produced along with copper chloride solution:

$$\underset{\text{Barium chloride}}{BaCl_2\,(aq)} + \underset{\text{Copper sulphate}}{CuSO_4\,(g)} \longrightarrow \underset{\text{Copper sulphide (Black ppt.)}}{CuS(s)} + \underset{\text{Copper chloride}}{H_2SO_4\,(aq)}$$

Example 4. **When hydrogen sulphide gas is passed through copper sulphate solution, then a black precipitate of copper sulphide is formed along with sulphuric acid solution:**

$$\underset{\text{Copper sulphate}}{CuSO_4\,(aq)} + \underset{\text{Hydrogen sulphide}}{H_2S\,(g)} \longrightarrow \underset{\text{Copper sulphide (Black ppt.)}}{CuS(s)} + \underset{\text{Sulphuric acid}}{H_2SO_4\,(aq)}$$

Example 5. **When ammonium hydroxide solution is added to aluminium chloride solution, then white precipitate of aluminium hydroxide is formed along with ammonium chloride solution:**

$$\underset{\text{Aluminium chloride}}{AlCl_3\,(aq)} + \underset{\text{Ammonium hydroxide}}{3NH_4OH\,(aq)} \longrightarrow \underset{\text{Aluminium hydroxide (white ppt.)}}{Al(OH)_3\,(s)} + \underset{\text{Ammonium chloride}}{3NH_4Cl\,(aq)}$$

Example 6. **When potassium iodide solution is added to lead nitrate solution, then a yellow precipitate of lead iodide is produced alongwith potassium nitrate solution:**

$$\underset{\text{Lead nitrate}}{Pb(NO_3)_2\,(aq)} + \underset{\text{Potassium iodide}}{2KI\,(aq)} \longrightarrow \underset{\text{Lead iodide (Yellow ppt.)}}{PbI_2\,(s)} + \underset{\text{Potassium nitrate}}{2KNO_3\,(aq)}$$

Example 7. **The reactions between acids and bases to form salts and water are also double displacement reactions. For example, sodium hydroxide and hydrochloric acid react to form sodium chloride and water:**

$$\underset{\text{Sodium hydroxide}}{NaOH\,(aq)} + \underset{\text{Hydrochloric acid}}{HCl\,(aq)} \longrightarrow \underset{\text{Sodium chloride}}{NaCl\,(aq)} + \underset{\text{Water}}{H_2O\,(l)}$$

❖ Reversible and Irreversible reactions

Reversible Reactions

- **Those reactions in which the products can re-combine to give back reactants are called reversible reactions.**

Example 1. When steam is passed over red-hot iron, then iron (II, III) oxide and hydrogen gas are obtained:

$$\underset{\text{Iron (Red-hot)}}{3Fe\,(s)} + \underset{\text{Stea}}{4H_2O\,(g)} \longrightarrow \underset{\text{Iron (II,III) oxide}}{Fe_3O_4\,(s)} + \underset{\text{Hydrogen}}{4H_2\,(g)}$$

- **Now, if hydrogen gas is passed over hot iron (II, III) oxide, then iron and steam are formed. This can be written as:**

$$\underset{\text{Iron (II,III) oxide (Hot)}}{Fe_3O_4\,(s)} + \underset{\text{Hydrogen}}{4H_2\,(g)} \longrightarrow \underset{\text{Iron}}{3Fe(s)} + \underset{\text{Steam}}{4H_2O(g)}$$

- **The above two reactions can be presented by a single chemical equation by using a double arrow as shown below:**

$$3Fe\,(s) + 4H_2O\,(g) \rightleftharpoons Fe_3O_4\,(s) + 4H_2\,(g)$$

- **The double arrow indicates that this reaction can go in both directions (forward and reverse),**
- **The double arrow indicates that this reaction can go in both directions (forward and reverse).this is a reversible reaction.**

- **The arrow pointing towards right (→) indicates the forward reaction whereas the arrow pointing towards left (←) indicates the reverse reaction.** ⇋
- **the reaction between red-hot iron and steam to form iron (II, III) oxide and hydrogen is a reversible reaction.**
- **The reversible reactions are a kind of "two-way reactions")**

Example 2. When mercury (II) oxide is heated strongly, it decomposes to form mercury metal and oxygen gas. This can be written as:

$$2HgO\ (s) \longrightarrow 2Hg\ (l) + O2\ (g)$$

Mercury (II) oxide — Mercury — oxygen

And when mercury and oxygen are heated gently, they combine to form mercury (II) oxide:

$$2Hg\ (l) + O_2\ (g) \longrightarrow 2HgO\ (s)$$

Mercury — oxygen — Mercury (II) oxide

- **The second reaction is the reverse of the first reaction. We can write this reversible reaction in a single equation as follows:**

$$2HgO\ (s) \rightleftharpoons 2Hg\ (l) + O_2(g)$$

Example 3. When nitrogen and hydrogen are heated together under pressure in the presence of finely divided iron as catalyst, then ammonia gas is formed (in the forward reaction):

$$N_2(g) + 3H_2\ (g) \longrightarrow 2NH_3\ (g)$$

Nitrogen — Hydrogen — Ammonia

- **Some of the ammonia formed begins to break down into nitrogen and hydrogen by the back reaction (or reverse reaction)**

$$2NH_3\,(g) \longrightarrow N_2\,(g) + 2H_2\,(g)$$

Ammonia — Nitrogen — Hydrogen

- **So, the best way to write the above two reactions together is like this:**

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

- **Thus, the reaction to produce ammonia from nitrogen and hydrogen is an example of reversible chemical reaction.**

Example 4. When hydrogen and iodine are heated together at high temperature in a sealed container, they combine to make hydrogen iodide (in the forward reaction).

$$H_2(g) + I_2\,(g) \longrightarrow 2HI$$

Hydrogen — Iodine — Hydrogen Iodine

- **Almost as soon as hydrogen iodide is formed, some hydrogen iodide molecules decompose in the back reaction (reverse reaction) to form hydrogen and iodine. That is:**

$$\underset{\text{Hydrogen iodide}}{2HI\ (g)} \longrightarrow \underset{\text{Hydrogen}}{H_2\ (g)} + \underset{\text{Iodide}}{I_2\ (g)}$$

- **We can write the above forward reaction and reverse reaction in one equation as follows.**

$$H_2\ (g) + I2\ (g) \rightleftharpoons 2SO_3\ (g)$$

- **Thus, the reaction between hydrogen and iodine to form hydrogen iodide is an example of a reversible reaction.)**

Example 5. **Sulphur dioxide combines with oxygen (of air) in the presence of vanadium (V) oxide as catalyst at 450°C to form sulphur trioxide:**

$$\underset{\text{Sulphur dioxide}}{2SO_2\ (g)} + \underset{\text{Oxygen (From air)}}{O_2\ (g)} \rightleftharpoons \underset{\text{Sulphur trioxide}}{2SO_3\ (g)}$$

Irreversible Reactions

- **Those reactions in which the products cannot re-combine to give back reactants are called irreversible reactions.**
- **Irreversible reactions are just the opposite of reversible reactions,**

Example 1. **Sodium metal reacts with water to form sodium hydroxide gas.**

$$\underset{\text{Sodium}}{2Na\,(s)} + \underset{\text{Water}}{2H_2O\,(l)} \longrightarrow \underset{\text{Sodium Hydrogen}}{2NaOH\,(aq)} + \underset{\text{Hydrogen}}{H2\,(g)}$$

- **An irreversible chemical reaction is represented by putting a single arrow(); in this equation (pointing from reactants to products side) as shown above.**
- **Actually, the irreversible reactions go completely in the forward direction,**

Example 2. **When methane gas burns in air, it combines with the oxygen (of air) to form carbon dioxide and water vapour. This can be written as:**

$$\underset{\text{Calcium carbonate (Marble chips)}}{CH4\,(g)} + \underset{\text{Oxygen (in air)}}{2O_2\,(g)} \longrightarrow \underset{\text{Carbon dioxide}}{CO2} + \underset{\text{Water vapour}}{2H_2O(g)}$$

- **The reactions like this which cannot be reversed are called irreversible reactions,**

Example 3. **Calcium carbonate (marble chips) react with dilute hydrochloric acid to form calcium chloride, water and carbon dioxide gas:**

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2\,(g)$$

Calcium carbonate (Marble chips) | Hydrochloric acid | Calcium chloride | Water | Carbon dioxide

- **We cannot re-combine calcium chloride, water and carbon dioxide gas to give back calcium carbonate and hydrochloric acid**

Example 4, Zinc metal reacts with dilute sulphuric acid to form zinc sulphate solution and hydrogen.

$$Zn\,(s) + H_2SO_4\,(aq) \longrightarrow ZnSO_4\,(aq) + H2\,(g)$$

Zinc (Metal) | Sulphuric acid (Dilute) | Zinc Sulphate solution | Hydrogen gas

- **The products of this reaction (zinc sulphate and hydrogen) cannot re-combine to give back reactants (zinc and sulphuric acid), so this is an irreversible reaction.**

Example 5. When manganese dioxide reacts with hydrochloric acid, then manganese dichloride.

$$MnO_2\,(s) + 4HCl\,(aq) \longrightarrow MnCl_2 + Cl_2\,(g) + 2H_2O\,(l)$$

Manganese dioxide | Hydrochloric acid | Manganese dichloride | chlorine | Water

❖ OXIDATION AND REDUCTION REACTIONS

- ❖ The earlier concept of oxidation and reduction is based on the addition or removal of oxygen or hydrogen elements.
- ❖ So, in terms of oxygen or hydrogen, oxidation and reduction reactions can be defined as follows:

Oxidation:

(i) The addition of oxygen to a substance is called oxidation.

(ii) The removal of hydrogen from a substance is also called oxidation.

Reduction

(i) The addition of hydrogen to a substance is called reduction.

(ii) The removal of oxygen from a substance is also called reduction.

- ➢ the process of reduction is just the opposite of oxidation. Moreover, oxidation and reduction occur together.

Oxidising agent

(i) The substance which gives oxygen for oxidation is called an oxidising agent.

(ii) The substance which removes hydrogen is also called an oxidising agent.

Reducing agent

(i) The substance which gives hydrogen for reduction is called a reducing agent.

(ii) The substance which removes oxygen is also called a reducing agent. The oxidation and reduction reactions are also called redox reactions (In the name 'redox, the

- **The oxidation and reduction reactions are also called redox reactions (In the name 'redox', the term 'red' stands for 'reduction' and 'ox' stands for oxidation).**

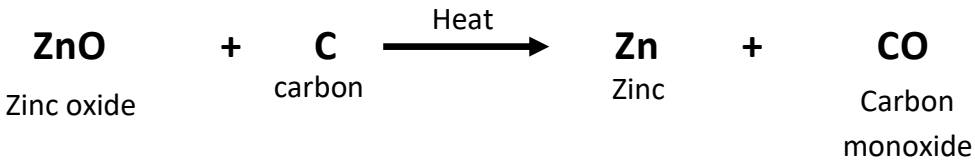
Example 1. When copper oxide is heated with hydrogen, then copper metal and water are formed:

$$\underset{\text{Copper oxide}}{CuO} + \underset{\text{Hydrogen}}{H_2} \xrightarrow{\text{Heat}} \underset{\text{Copper}}{Cu} + \underset{\text{Water}}{H_2O}$$

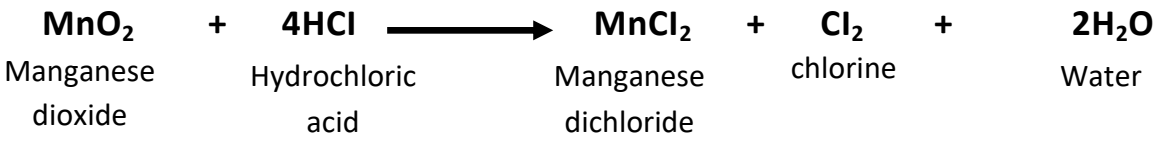
Example 2. When hydrogen sulphide reacts with chlorine, then sulphur and hydrogen chloride are formed.

$$\underset{\text{Hydrogen Sulphide}}{H_2S} + \underset{\text{chlorine}}{Cl_2} \longrightarrow \underset{\text{Sulphur}}{S} + \underset{\text{Hydrogen chloride}}{2HCl}$$

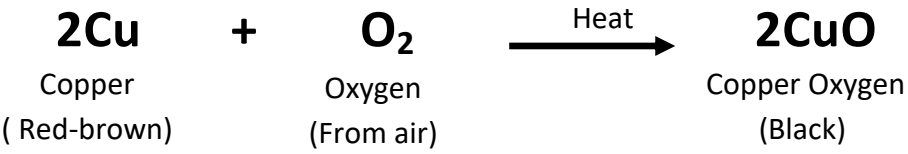
Example 3. **When zinc oxide is heated with carbon, then zinc metal and carbon monoxide are formed:**



Example 4. **When manganese dioxide reacts with hydrochloric acid, then manganese dichloride, chlorine and water are formed:**



. Example 5. **When copper is heated in air, it reacts with the oxygen of air to form a black compound copper oxide:**



EFFECTS OF OXIDATION REACTIONS IN EVERYDAY LIFE

- **Oxidation has damaging effect on metals as well as on food.**
- **there are two common effects of oxidation reactions which we observe in daily life. These are:**

  - **Corrosion of metals, and**

- **Rancidity of food.**

Corrosion

- **Corrosion is the process in which metals are eaten up gradually by the action of air, moisture or a chemical (such as an acid) on their surface.**
- **Corrosion is caused mainly by the oxidation of metals by the oxygen of air.**
- **Rusting of iron metal is the most common form of corrosion.**
- **During the corrosion of iron (or rusting of iron), iron metal is oxidised by the oxygen of air in the presence of water (moisture) to form hydrated iron (III) oxide called rust:**

$$\underset{\text{Iron}}{4Fe} + \underset{\text{Oxygen}}{3O_2} + \underset{\text{Water}}{2xH_2O} \longrightarrow \underset{\text{Hydrated iron (III) oxide (Rust)}}{2Fe_2O_3xH_2O}$$

- **he rusting of iron is a redox reaction.**
- **Corrosion weakens the iron and steel objects and structures such as railings, car bodies, bridges and ships, etc., and cuts short their life.**

Rancidity

- **The condition produced by aerial oxidation of fats and oils in foods marked by unpleasant smell and taste is called rancidity.**
- **Rancidity spoils the food materials prepared in fats and oils which have been kept for a considerable time and makes them unfit for eating.**

1. **Rancidity can be prevented by adding anti-oxidants to foods containing fats and oils.**
2. **Rancidity can be prevented by packaging fat and oil containing foods in nitrogen gas.**
3. **Rancidity can be retarded by keeping food in a refrigerator.**
4. **Rancidity can be retarded by storing food in air-tight containers.**
5. **Rancidity can be retarded by storing foods away from light.**