Chemical Reactions

Chemical Reactions

A chemical reaction is a process in which a chemical substance is completely transformed into another substance having different properties. This process is different from the process of physical changes because in physical changes no new substance is formed or the properties of the substance after the change remain the same.

Types of Chemical Reaction

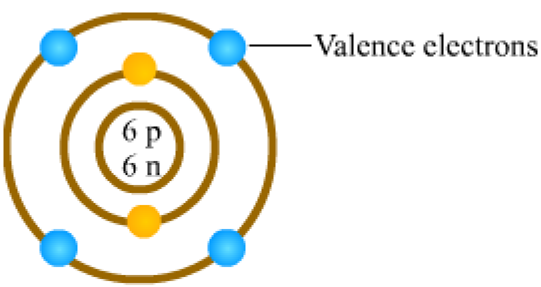
• **Simple chemical reaction:** Those chemical reactions which occur usually in one step and take lesser time to complete are known as simple chemical reactions. For example, burning of fuel and rusting of iron.

• **Complex chemical reaction:** Those chemical reactions which take longer time for their completion and occur in more than one step are known as complex chemical reactions. For example, photosynthesis reaction, digestion of food, and all biochemical reactions.

Chemical changes are very important to us as they lead to the formation of substances which help us to grow our food, make our lives more productive and comfortable, cure diseases etc.

Valence Electrons

All chemical reactions occur because the unstable atoms tend to achieve stable electronic configuration. In order to attain stable electronic configuration, an electron may get transferred from one atom to another or may be shared between two atoms. Those electrons which take part in chemical reactions are known as **valence electrons.** The valence electrons may be in the outer most orbit or in a penultimate orbit of an atom (Penultimate orbit means before the last orbit).



Note: Atoms combine with each other in order to complete their octet. All atoms try to achieve the nearest noble gas configuration, and in doing so they may lose electrons, gain electrons or share electrons. For example, sodium has a total of 11 electrons and its electronic configuration is 2, 8, 1. In order to attain the nearest stable noble gas configuration (Ne = 2,8), sodium losses 1 electron to form Na^+ ion with electronic configuration 2,8.

How do you know whether a chemical reaction has taken place or not?

Some changes that are observed during a chemical reaction are:

Evolution of gas: In many chemical reactions, gas is evolved sometimes with an effervescence. For, example when zinc metal reacts with hydrogen chloride, hydrogen gas is evolved.

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

Change of colour: Sometimes, a colour change occurs during a chemical reaction due to the formation of a new product. For example, when silver chloride is exposed to sunlight it produces black metallic coloured silver with the evolution of chlorine gas.

$2AgCl \rightarrow 2Ag + Cl_2$

Formation of precipitate: In certain chemical reactions, an insoluble solid substance called **precipitate** is formed. For example, silver nitrate reacts with sodium bromide to form a yellow precipitate of silver bromide.

$AgNO_3 + NaBr \rightarrow NaNO_3 + AgBr \downarrow$

Change of state: In some reactions, the change of state is observed in the formation of products from reactants. For example, sodium chloride (aqueous) when reacts with silver nitrate (aqueous) forms sodium nitrate (aqueous) and silver chloride (solid white precipitate).

$NaCl\ (aq) + AgNO_3\ (aq) \rightarrow NaNO_3\ (aq) + AgCl$ (white ppt)

Conditions for Chemical Change

For a chemical reaction to proceed, certain physical conditions are required. These conditions are:

- **Close contact**

Some chemical reactions proceed only when the reactant molecule are brought together in close contact with each other.

The intimate contact can be brought by

- grinding the reactants together
- dissolving the reactants in water

Example: Potassium iodide reacts with mercury chloride when they are thoroughly grinded together.

$$2KI + HgCl_2 \rightarrow HgI_2 + 2KCl$$

- **Heat**

  Certain chemical reactions proceed only when reactants are heated together.

  Example: The given reaction occurs only when reactant is heated.

$$KClO_3 \xrightarrow{\text{Heat}} 2KCl + 3O_2$$

- **Light**

  Certain chemical reactions proceed when reactants are exposed to sunlight or diffused sunlight.

  Example: The given reaction occurs only when reactants are exposed to sunlight.

$$H_2 + Cl_2 \xrightarrow{\text{sunlight}} 2HCl$$

- **Pressure**

  Certain chemical reactions proceed when reactants are subjected to a pressure higher than atmospheric pressure.

  Example: Nitrogen and hydrogen react in the presence of catalyst when subjected to a pressure between 200- 900 atms.

$$N_2 + 3H_2 \xrightarrow{200-900\text{ atms}} 2NH_3$$

- **Catalytic agent**

  Certain chemical reactions proceed in forward direction when brought in contact with a catalyst.

  Example: Sulphur dioxide and oxygen react in the presence of asbestos, which acts as a catalyst.

$$SO_2 + O_2 \xleftrightarrow[450°C]{\text{Pt-Asbestos}} 2\,SO_3$$

- **Electric Current**

  Certain chemical reactions proceed only when an electric current is passed through reactants in fused state or in aqueous solution.

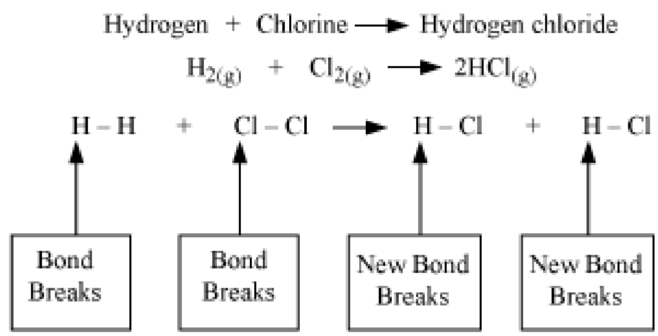
  Example: Acidulated water decomposes into hydrogen and oxygen only when electric current is passed.

$$2\,H_2O \xleftrightarrow[\text{Current}]{\text{Electric}} 2\,H_2 + O_2$$

Energy change in chemical reactions

During a chemical reaction, old bonds are broken and new bonds are formed. Energy is required to break the old bonds between the molecules of reactants; this energy is commonly called **activation energy.**

Also, energy is liberated when new bonds are formed. Thus, a chemical change is associated with absorption or release of energy. This energy can be in the form of **heat, sound, or electricity.**



Every substance is associated with a certain amount of energy stored in it in the form of latent energy. This stored energy is called **chemical energy or internal energy.** It is denoted by E.

The internal energy of a substance is the sum total of kinetic energy and potential energy.

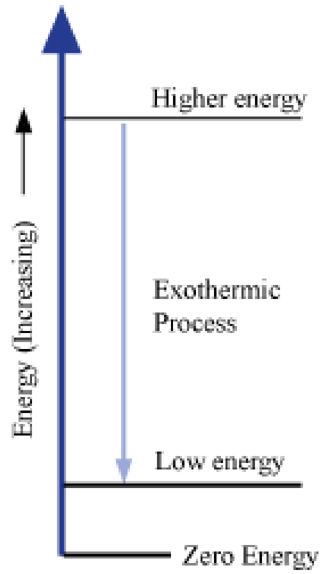
E = K.E. + P.E.

The internal energy is different for different substances. Hence, internal energy for reactants is different for internal energy of products.

The difference between chemical energy of reactants and chemical energy of products is called **energy change of a chemical reaction.**

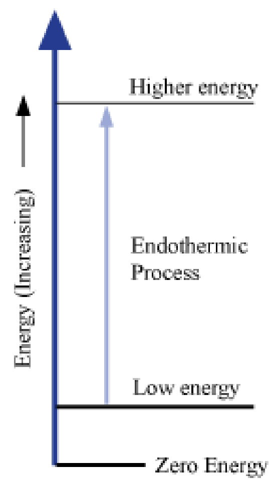
In an **exothermic reaction**, internal energy of reactants is greater than internal energy of the products and therefore, change in the energy will be negative. It is represented as:

Reactants + Energy → Products



In **an endothermic reaction**, the internal energy of reactants is less than the internal energy of the products. It is represented as:

Reactants → Products + Energy

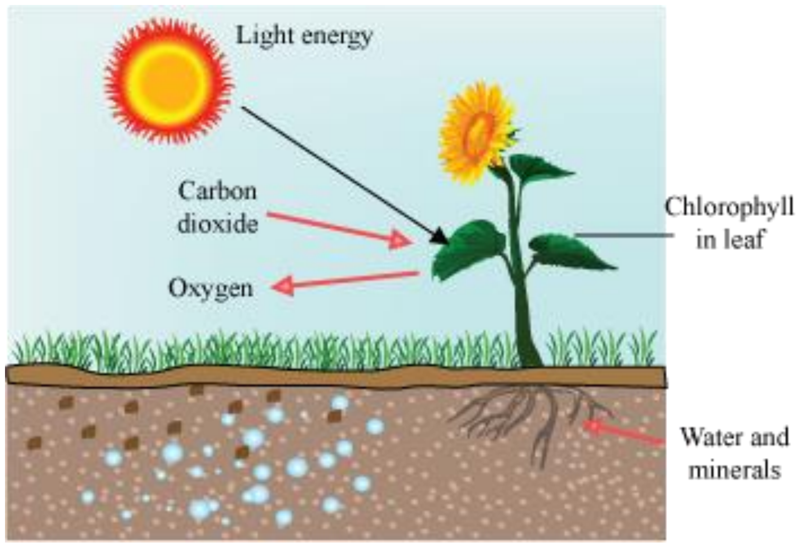


Energy released or absorbed is measured in kilocalories or kilojoules.

Endothermic reactions are the reactions, which absorb energy in the form of heat. The opposite of an endothermic process is an **exothermic process**, one that releases energy in the form of heat.

Photosynthesis is an example of an endothermic chemical reaction. In this process, plants use the energy from the sun to convert carbon dioxide and water into glucose and oxygen.

$$6\,CO_2 + 6\,H_2 \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} 6\,O_2 + C_6H_{12}O_6$$

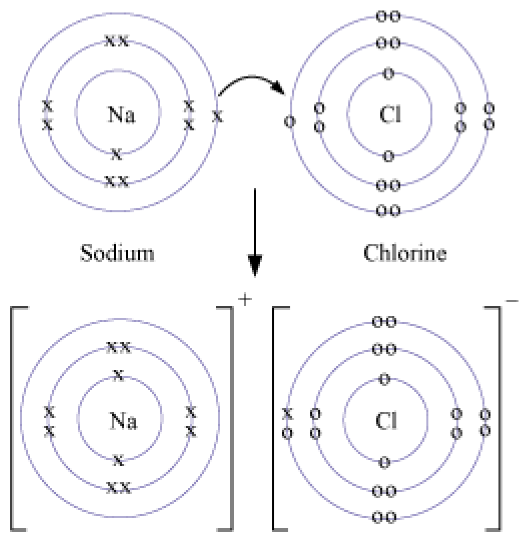


Some other examples of endothermic processes are:

- Depressurising a pressure can
- A chemical cold pack consisting primarily of ammonium nitrate and water
- Melting of solids
- Vaporisation, evaporation, fusion

An example of an exothermic reaction is the mixture of sodium and chlorine to yield table salt. This reaction produces 411 kJ of energy for each mole of salt that is produced.

$$Na_{(s)} + 0.5\ Cl_{2(s)} \longrightarrow NaCl_{(s)}$$



Sodium chloride

Some examples of exothermic processes are:

- Condensation of rain from water vapour
- Combustion of fuels such as petrol, wood, coal, and oil
- Hydration processes

Catalytic Reactions

Catalytic reactions

Catalysis is the process in which the rate of a chemical reaction is either increased or decreased by a chemical substance known as a **catalyst**.

Catalysts are substances which facilitate chemical reactions without themselves undergoing a chemical change.

Negative catalyst or inhibitor is a substance that slows down the rate of a reaction. It retards the efficiency of a catalyst.

Promoter is a substance which improves the rate of reaction by improving the efficiency of a catalyst.

Catalytic reactions can be represented as:

$$\text{Reactants} \xrightarrow{\text{catalyst}} \text{Products}$$

Examples

- **For positive catalysis**

  $2KClO_3 \rightarrow 2KCl + 3O_2$

- **For negative catalysis**

  $2H_2O_2 \rightarrow O_2 + H_2O$

- **For promoter**

  In the manufacture of ammonia from nitrogen and hydrogen, iron acts as the catalyst and molybdenum act as a promoter.

- **For inhibitor**

  In the manufacture of sulphuric acid, arsenic oxide acts as an inhibitor by inhibiting the action of the catalyst asbestos.

Enzymes as Catalysts

Enzymes catalyze biochemical reactions. Almost all enzymes are proteins, except ribozymes (nucleic acids acting like enzymes). They accelerate the reaction in a cell or outside it without getting consumed.

For example, maltase helps in the breakdown of maltose into glucose.

$$Maltose + Water \xrightarrow{Maltase(enzyme)} Glu\cos e$$

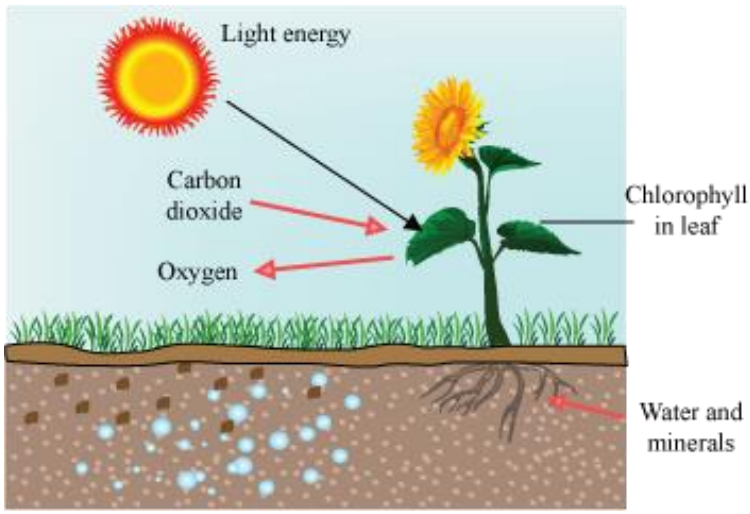
Photochemical reactions

Photochemical reactions are the reactions which proceed with the absorption of light energy. Let us look at some examples of photochemical reactions.

- **Photosynthesis**

  Plants produce glucose and oxygen from carbon dioxide and water in the presence of sunlight and chlorophyll.

$$6\,CO_2 + 6\,H_2 \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} 6\,O_2 + C_6H_{12}O_6$$



- When an equal volume of hydrogen gas and chlorine gas are mixed and exposed to direct sunlight, a violent explosion takes place with the formation of hydrochloric acid gas.

$$H_2 + Cl_2 \xrightarrow{\text{Sunlight}} 2\,HCl$$

Combination Reactions

You know that chemical changes involve chemical reactions. Chemical reactions are primarily of five types. They are listed as follows:

1. Combination reactions

2. Decomposition reactions

3. Displacement reactions

4. Double displacement reactions

5. Oxidation and reduction reactions

Here, we will discuss combination reactions in detail. **Do you know what actually happens in a combination reaction?**

Combination reactions

In these reactions, two or more substances combine to form a new compound. The reactants in such reactions can be elements as well as compounds. The general equation used to represent a combination reaction is:

$A + Z \longrightarrow AZ$

For example, coal is primarily carbon. When it burns, it combines with oxygen present in the air to form carbon dioxide.

$$\begin{array}{ccccc} C(s) & + & O_2(g) & \rightarrow & CO_2(g) \\ \text{Carbon} & & \text{Oxygen} & & \text{Carbon dioxide} \\ \text{(From coal)} & & & & \end{array}$$

Some other examples of combination reactions are discussed below.

1. Combination of two elements

On heating, magnesium combines with oxygen present in the air to form magnesium oxide.

$$\begin{array}{ccccc} 2\,Mg\ (s) & + & O_2\,(g) & \rightarrow & 2\,MgO\ (g) \\ \text{Magnesium} & & \text{Oxygen} & & \text{Magnesium oxide} \end{array}$$

Hydrogen and oxygen combine to form water.

$$\begin{array}{ccccc} 2H_2\ (g) & + & O_2\ (g) & \rightarrow & 2H_2O\ (l) \\ \text{Hydrogen} & & \text{Oxygen} & & \text{Water} \end{array}$$

2. Combination of two compounds

Calcium oxide, also known as quick lime, when mixed with water reacts with it to form calcium hydroxide, also known as slaked lime. The chemical equation for the same is given as:

$$\begin{array}{ccccc} CaO(s) & + & H_2O(l) & \rightarrow & Ca(OH)_2\,(aq) \\ \text{Calcium oxide} & & \text{Water} & & \text{Calcium hydroxide} \\ \text{(Quick lime)} & & & & \text{(Slaked lime)} \end{array}$$

Hence, it can be concluded that combination reactions are generally exothermic in nature. In the above activity, CaO combines with water to give only a single product, Ca $(OH)_2$.

However, there are very few combination reactions which are endothermic in nature. One of the examples of such a reaction is combination of nitrogen and oxygen gas to form nitrogen dioxide gas:

$$\begin{array}{ccccc} N_2\ (g) & + & O_2\ (g) & \rightarrow & 2\,NO\ (g) \\ \text{Nitrogen} & & \text{Oxygen} & & \text{Nitric oxide} \end{array}$$

In this reaction, reactants absorb energy from the surroundings in order to form product.

What happens when coal is burned? On burning, coal combines with oxygen to produce carbon dioxide. It also gives a lot of heat energy. Hence, burning of coal is an exothermic reaction.

DO YOU KNOW?

Lime water or slaked lime ($Ca(OH)_2$) is used in white washing of walls. It combines with carbon dioxide present in the air to form a thin layer of calcium carbonate. The chemical formula of calcium carbonate is $CaCO_3$. The chemical equation involved in the reaction can be represented as:

$$\underset{\text{Slaked lime}}{Ca(OH)_2\,(aq)} + \underset{\text{Carbon dioxide}}{CO_2\,(g)} \rightarrow \underset{\text{Calcium carbonate}}{CaCO_3\,(s)} + \underset{\text{Water}}{H_2O\,(l)}$$

Questions asked in previous years' board examinations

Ques. Define a combination reaction. Give one example of a combination reaction which is also exothermic.

(2 marks)

−2009 CBSE Delhi

Sol: In combination reactions, two or more substances combine to form a new compound. Only one product is obtained in such reactions. The reactants in such reactions can be elements as well as compounds. The general equation used to represent a combination reaction is:

$$A + Z \longrightarrow AZ$$

For example, calcium oxide reacts vigorously with water to produce calcium hydroxide.

$$\underset{\text{Calcium oxide}}{CaO\,(s)} + \underset{\text{Water}}{H_2O\,(l)} \rightarrow \underset{\text{Calcium hydroxide}}{Ca(OH)_2\,(aq)}$$

A large amount of heat is also evolved during this process, which increases the temperature of the system. Hence, the combination of calcium oxide and water is exothermic in nature.

Decomposition Reactions

We know that chemical reactions are primarily of five types. They are listed as follows:

1. Combination reactions

2. Decomposition reactions

3. Displacement reactions

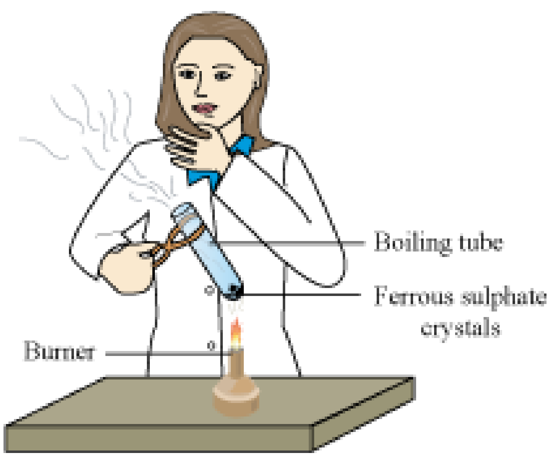
4. Double displacement reactions

5. Oxidation and reduction reactions

The following activity can be performed to understand decomposition reactions.

Activity:

Take 3 g of green ferrous sulphate crystals in a dry boiling tube. Heat the boiling tube over the flame of a burner. Observe the change in color of the crystals on heating.



It will be observed that the colour of the crystals undergoes a change. Also, the characteristic smell of burning sulphur is observed. **Do you know why this happens?**

Here, green crystals of ferrous sulphate lose water on heating. Hence, a change in colour is seen in the crystals. On further heating, it decomposes into ferric oxide, sulphur dioxide, and sulphur trioxide. The chemical equation involved in the reaction can be represented as:

$$\underset{\text{Ferrous sulphate}}{2FeSO_4\ (s)} \xrightarrow{\text{Heat}} \underset{\text{Ferric oxide}}{Fe_2O_3\ (s)} + \underset{\text{Sulphur dioxide}}{SO_2\ (g)} + \underset{\text{Sulphur trioxide}}{SO_3\ (g)}$$

Here, ferrous sulphate breaks down or decomposes to form three new substances. Hence, it is an example of decomposition reactions.

What are decomposition reactions?

In these reactions, a compound breaks down or decomposes to form two or more substances. These reactions are exactly opposite to combination reactions. We know that there is only one product in combination reactions. Similarly, there is only one reactant in decomposition reactions. The general equation used to represent a decomposition reaction is:

$$XY \longrightarrow X + Y$$

Decomposition reactions require a source of energy in the form of heat, light, or electricity to decompose the compound involved. Hence, these reactions can be classified into three types, depending on the source of energy for the reaction.

a) Decomposition by heat or thermal decomposition

b) Decomposition by electricity or electrolysis

c) Decomposition by light or photolysis

Let us now study three different types of decomposition reactions.

a) Decomposition by heat or thermal energy

One of the most common examples of thermal decomposition reactions is the decomposition of calcium carbonate. Calcium carbonate when heated decomposes to form calcium oxide and carbon dioxide.

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

In this reaction, one compound i.e. calcium carbonate breaks down to form two compounds, namely calcium oxide and carbon dioxide. Hence, it is an example of decomposition reactions. Commercially, this reaction is very important as calcium oxide (obtained as a product in this reaction) is used in cement and glass industries.

Hands-on Activity

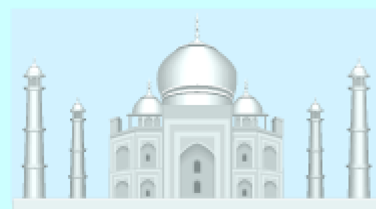
Take about 3 g of solid lead nitrate in a boiling tube. Note the colour of the compound. Heat it in the flame of the Bunsen burner. Observe the change taking place.

You will observe that emission of brown fumes occurs. These fumes are of nitrogen dioxide.

During this reaction, lead nitrate decomposes to form lead oxide, nitrogen dioxide, and oxygen gas. The following reaction takes place:

$2Pb(NO_3)_2\,(s) \rightarrow 2PbO\,(s) + 4NO_2\,(g) + O_2\,(g)$
Lead nitrate   Lead oxide   Nitrogen dioxide   Oxygen

The Taj Mahal is made up of marble. **Do you know that chemically, marble is nothing but calcium carbonate?**



Thermal decomposition of some compounds:

- Metal hydroxides: They decompose on heating to produce metal oxide and water or steam.

  - $Cu(OH)_2\,(s) \xrightarrow{\Delta} CuO\,(s) + H_2O\,(g)$
- Metal carbonates: They decompose on heating to produce metal oxide and carbon dioxide gas.
  - $CuCO_3\,(s) \xrightarrow{\Delta} CuO\,(s) + CO_2\,(g)$
- Metal bicarbonates: They decompose on heating to produce metal carbonate, carbon dioxide gas, and water.
  - $2\ NaHCO_3\,(s) \xrightarrow{\Delta} Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(g)$
- Metal nitrates: They decompose on heating to produce metal oxide, nitrogen dioxide, and oxygen gas.
  - $2\ Mg(NO_3)_2\,(s) \xrightarrow{\Delta} 2\ MgO\,(s) + 4\ NO_2\,(g) + O_2\,(g)$

b) Decomposition by electricity

When electricity is passed through water containing a few drops of sulphuric acid, it breaks down to give its constituent elements as products i.e. hydrogen and oxygen. This is known as electrolysis of water. Let us understand decomposition by electricity with one of its application in the real world.

c) Decomposition by light

When silver chloride is kept in the sun, it decomposes to form chlorine gas and metallic silver. As the reaction proceeds, the white coloured silver chloride turns grey because of the formation of silver. Chlorine produced in the reaction escapes into the environment as it is produced in the gaseous state.

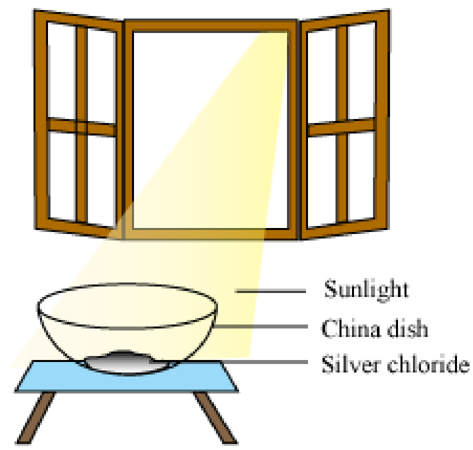


Figure 3: Photolysis of silver chloride

$$\underset{\text{Silver chloride}}{2AgCl(s)} \xrightarrow{\text{Light}} \underset{\text{Silver}}{2Ag(s)} + \underset{\text{Chlorine}}{Cl_2(g)}$$

Silver bromide also undergoes decomposition in a similar manner when exposed to sunlight.

$$\underset{\text{Silver bromide}}{2AgBr(s)} \xrightarrow{\text{Light}} \underset{\text{Silver}}{2Ag(s)} + \underset{\text{Bromine}}{Br_2(g)}$$

As the above reactions are sensitive to light, they are used in black and white photography.

It is seen that decomposition reactions require a source of energy in the form of heat, light, or electricity to decompose the compound involved. Hence, it can be concluded that decomposition reactions are **endothermic** in nature.

Questions asked in previous years' board examinations
Ques. (a) What is the colour of ferrous sulphate crystals? How does this colour change after heating?

(b) Name the products formed on strongly heating ferrous sulphate crystals.

(2 marks)

−2009 CBSE Delhi

Sol: a) The colour of ferrous sulphate crystals is green.

On heating, ferrous sulphate crystals ($FeSO_4.7H_2O$) lose their water of crystallisation and due to this, the colour of the compound changes to white/colourless.

(b) On strong heating, ferrous sulphate crystals give ferric oxide (Fe_2O_3), sulphur dioxide (SO_2) and sulphur trioxide (SO_3) as products.

$$2\,FeSO_4(s) \xrightarrow{\text{Heat}} Fe_2O_3(s) + SO_2(g) + SO_3(g)$$

Decomposition reaction occurs in this change.

Ques. Give an example of a decomposition reaction. Describe an activity to illustrate such a reaction by heating.

(2 marks)

−2008 CBSE Delhi

Sol: Example of decomposition reaction: $$\underset{\text{Calcium carbonate}}{CaCO_3(s)} \xrightarrow{\Delta} \underset{\text{Calcium oxide}}{CaO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

2 g of lead nitrate is taken in a boiling tube and heated. On heating, lead nitrate decomposes to produce lead oxide, nitrogen dioxide, and oxygen. The chemical equation involved in the reaction is:

$$\underset{\text{Lead nitrate}}{2Pb(NO_3)_2(s)} \xrightarrow{\text{Heat}} \underset{\text{Lead oxide}}{2PbO(s)} + \underset{\text{Nitrogen dioxide}}{4NO_2(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

Displacement Reactions

We know that chemical reactions are primarily of five types. They are listed below.

1. Combination reactions

2. Decomposition reactions

3. Displacement reactions

4. Double displacement reactions

5. Oxidation and reduction reactions

In this part, we will discuss displacement and double displacement reactions in detail.

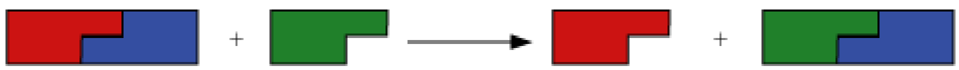
In displacement reactions, a more reactive metal replaces a less reactive metal from the latter's salt.

Reactions in which a more reactive element replaces a less reactive element from the salt solution of the less reactive element are called displacement reactions.

Do you know that displacement reactions are of two types? They are:

1. Single Displacement Reactions
2. Double Displacement Reactions

Single Displacement Reactions can be better understood with the help of the following figure.



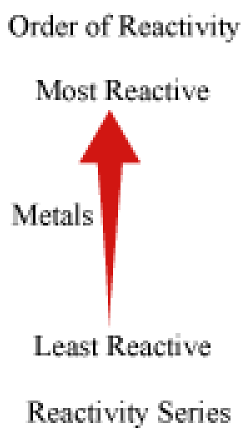
In the above figure, you have three blocks. It will be observed that while red and blue blocks are fixed in, green block is aloof. Now, if a blue block is detached from the red and fixed with the green, it will mean that the green block displaces the red block.

Thus, in a single displacement reaction, an uncombined single element replaces the other element present in a compound.

Another example of single displacement reaction is:

$$\underset{\text{Zinc}}{Zn\ (s)} + \underset{\text{Copper sulphate}}{CuSO_4\ (aq)} \rightarrow \underset{\text{Zinc sulphate}}{ZnSO_4\ (aq)} + \underset{\text{Copper}}{Cu\ (s)}$$

The reactivity of metals can be known from the reactivity series, which lists metals in their respective order of reactivity (most reactive at the top, least reactive at the bottom).



Now, consider the following figure.

Do you observe any difference from the first block sequence? In the above figure, there are four different blocks with different colours in two pairs. These blocks are detached. Then, the blue block is exchanged with the yellow block. This represents a double displacement reaction.

A Double Displacement Reaction is a bimolecular process in which parts of two compounds are exchanged to give two new compounds. The general equation used to represent double displacement reactions can be written as:

$AB + CD \rightarrow AD + BC$

Double Displacement Reactions have two common features:

- Firstly, two compounds exchange their ions resulting in the formation of new compounds.
- Secondly, one of the new products formed would be separated from the mixture in some way (commonly as a solid or gas).

Hands on activity
Activity - I Take 2 mL each of lead nitrate and potassium iodide solution in two separate test tubes. Gently pour the potassium iodide solution into the lead nitrate solution.

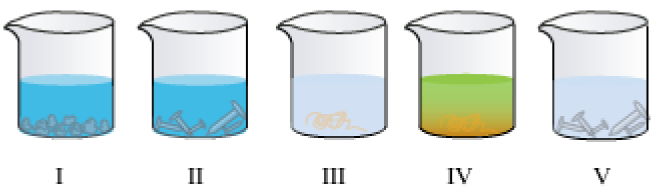
As soon as you do this, you will observe the formation of a yellow precipitate. This yellow precipitate is of lead iodide. In this reaction, the two compounds lead nitrate and potassium iodide react by exchanging their ions to form new compounds, lead iodide and potassium nitrate.

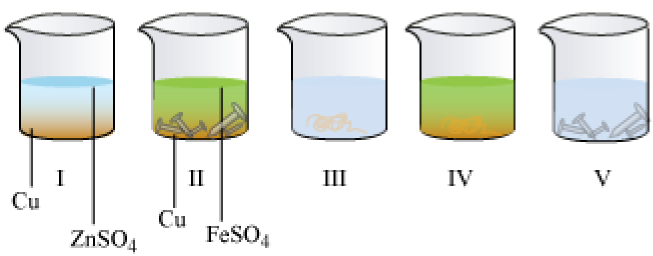
The equation involved in this reaction is:

$$Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3$$

Activity - II

Take five 100 mL beakers and add 20 mL water in them. Label the beakers as **I, II, III, IV,** and **V**. Add 5 g copper sulphate to beakers **I** and **II**, 5 g zinc sulphate to **III** and **V,** and 5 g iron sulphate to beaker **IV**. Now, add some iron nails to beakers **II** and **V,** copper turnings to beakers **III** and **IV**, and zinc granules to beaker **I.** Then, keep the beakers undisturbed for some time and observe carefully.





You will observe that the colour of copper sulphate solution changes in beakers **I** and **II.** On the other hand, no change is observed in beakers **III, IV,** and **V**.

Can you explain these observations using the concept of displacement reactions?

In beaker **I,** zinc (Zn) replaces copper (Cu) from copper sulphate ($CuSO_4$) solution to form zinc sulphate ($ZnSO_4$) and copper. Because of this, the blue colour of copper sulphate disappears and a reddish brown substance i.e. copper gets deposited at the bottom of the beaker. The chemical equation for the reaction can be represented as:

$$\underset{\text{Zinc}}{Zn\ (s)} + \underset{\text{Copper sulphate}}{CuSO_4\ (aq)} \rightarrow \underset{\text{Zinc sulphate}}{ZnSO_4\ (aq)} + \underset{\text{Copper}}{Cu\ (s)}$$

Similarly, in beaker **II,** iron replaces copper from copper sulphate solution. Hence, the colour of the solution changes from blue to green and a reddish brown substance gets deposited on the iron nail.

$$\begin{array}{llll} \mathrm{Fe\ (s)} + & \mathrm{CuSO_4\ (aq)} \rightarrow & \mathrm{FeSO_4\ (aq)} + & \mathrm{Cu\ (s)} \\ \text{Iron} & \text{Copper sulphate} & \text{Iron(II) sulphate} & \text{Copper} \end{array}$$

Do you know why there are no changes in beakers III, IV, and V?

Since no change is observed in beakers **III, IV,** and **V**, it can be concluded that copper is less reactive than zinc and iron. Hence, copper can not replace zinc from zinc sulphate solution and iron from iron sulphate solution. Therefore, we can also say that iron is less reactive than zinc. Hence, iron cannot replace zinc from zinc sulphate solution.

Hence, it can be concluded that in displacement reactions, a more reactive metal replaces a less reactive metal from its salt solution, whereas a less reactive metal cannot replace a more reactive metal.

Types of double displacement reaction: A Double Displacement Reaction is of three types.

- **Precipitation reaction**

  In precipitation reaction, soluble ions in separate solutions are mixed together to form an insoluble compound that settles out of the solution as a solid. This insoluble compound is called a precipitate.

  Example:

  If an aqueous solution of sodium sulphate is mixed with barium chloride, it will be observed that a white insoluble substance is formed. The white insoluble substance is called a **precipitate**. Here, barium chloride reacts with sodium sulphate to produce barium sulphate (white insoluble precipitate) and sodium chloride. Thus, this is an example of a double displacement reaction. The chemical equation involved in the reaction is

$$\begin{array}{llll} \mathrm{BaCl_2\ (aq)} + & \mathrm{Na_2\,SO_4\ (aq)} \rightarrow & \mathrm{BaSO_4\ (s)} + & 2\,\mathrm{NaCl}\ (s) \\ \text{Barium chloride} & \text{Sodium sulphate} & \text{Barium sulphate} & \text{Sodium chloride} \end{array}$$

- **Neutralisation reaction**

  Neutralisation reaction is a chemical reaction in which an acid and a base react to produce salt and water (H_2O).

  Example:

$$\underset{\text{Sodium hydroxide}}{2NaOH\ (aq)} + \underset{\text{Sulphuric acid}}{H_2SO_4\ (aq)} \rightarrow \underset{\text{Sodium sulphate}}{Na_2SO_4\ (s)} + \underset{\text{Water}}{2H_2O\ (l)}$$

- **Gas forming reaction**

Gas forming reactions are those reactions in which either, one of the product is formed in gaseous state or a product decomposes instantly to form a gaseous compound.

Example:

$$\underset{\text{Nitric acid}}{2HNO_3\ (aq)} + \underset{\text{Sodium sulphite}}{Na_2SO_3\ (aq)} \rightarrow \underset{\text{Sodium nitrate}}{2NaNO_3\ (aq)} + \underset{\text{Water}}{H_2O\ (l)} + \underset{\text{Sulphur dioxide}}{SO_2\ (g)}$$

Natural Indicators

In our daily life, we use many substances such as curd, lemon juice, vinegar etc. that are either acidic or basic in nature. However, to determine the acidic or basic nature of an unknown substance, it should not be tasted, as it may be harmful for us. **Then, how can we determine whether a substance is acidic or basic?**

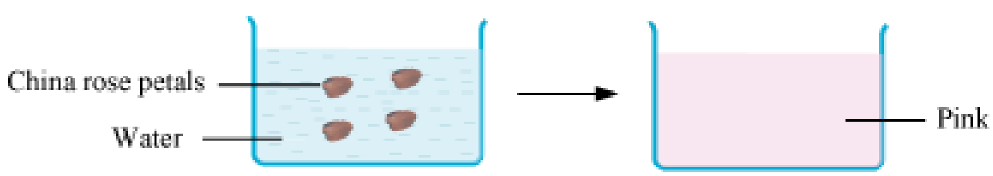
To distinguish an acid from a base, special compounds are used called **indicators**.

Hence, substance can be identified as acids, bases, or neutrals on the basis of the change in colour produced by red and blue litmus paper.

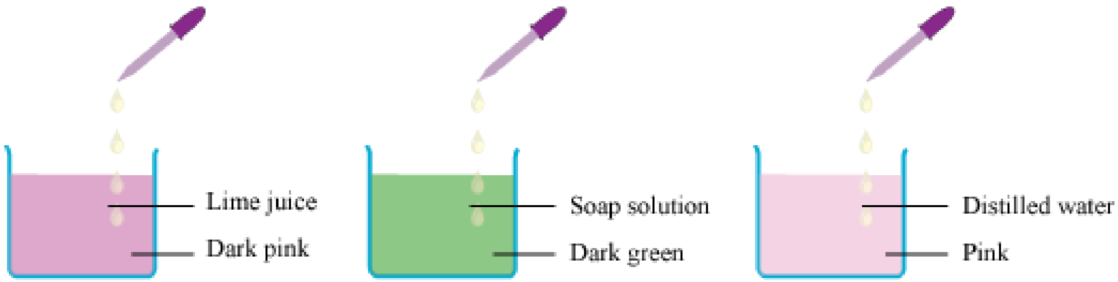
Solution	Red litmus paper	Blue litmus paper
Acidic	Red	Red
Basic	Blue	Blue
Neutral	No change in colour	No change in colour

Do you know that turmeric and China rose can also act as acid base indicators?

China rose indicator is prepared by adding petals of China rose in warm water. Then, the mixture is kept for some time until the water becomes pink in colour (as shown in the figure).



Now, if we add a few drops of this pink solution to lime juice, it turns dark pink. If we add a few drops of this pink solution to soap solution, it turns green in colour (as shown in the figure below). However, if the pink solution is added to water, no change in colour will be observed.

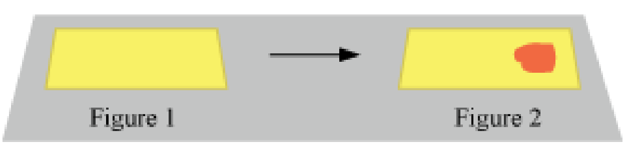


Thus, it can be concluded that China rose indicator gives dark pink colour with acids and green colour with bases. With neutral solutions, there is no change in colour of the China rose indictor.

Turmeric paper is another natural indicator. Turmeric paper is prepared by drying a blotting or filter paper, after applying turmeric paste on it. Turmeric paste is prepared by mixing turmeric powder and water. This strip of yellow paper is then used as an indicator strip, in the same way as litmus paper is used.

Figure - I

The colour of turmeric paper remains yellow in acidic solutions and changes to red in basic solutions. For example, if a drop of soap solution is put on a strip of turmeric paper with the help of a dropper, then that portion of the turmeric paper will turn red (as shown in figure 2). Hence, it shows that soap solution is basic in nature.



Activity:

Take hydrochloric acid, sulphuric acid, acetic acid, sodium hydroxide, lime water, ammonium hydroxide, sugar solution, common salt solution, and magnesium hydroxide. Test the nature of these substances with the help of three indicators: litmus paper, turmeric paper, and China rose solution.

The presence of acids and bases can also be tested by using phenolphthalein indicator. A drop of phenolphthalein when added to a basic solution, changes the colour of the solution to pink. On the other hand, an acidic solution has no effect of phenolphthalein and hence remains colourless.

Common Examples of Neutralization Reactions

Do you know that our stomach contains hydrochloric acid? Hydrochloric acid is essential for the digestion of food. Sometimes, our stomach produces excess acid, which causes pain and irritation. This condition is known as **acidity** or **indigestion**.

To get relief from this condition, milk of magnesia is used as an antacid. Milk of magnesia contains magnesium hydroxide, which is a mild base. It reacts with excess acid present in the stomach, and neutralizes it. This is an example of neutralization reaction, i.e., reaction taking place between an acid and a base.

When an acid is mixed with a base, both neutralize the effect of each other. **Thus, the reaction between an acid and a base is known as neutralization reaction**. Salt and water are produced in this process with the evolution of heat.

Acid + Base → Salt + Water (Heat is evolved)

In our everyday life, we observe many examples of neutralization reactions. For example, a honeybee’s sting causes pain and irritation as it contains formic acid. Similarly, when an ant bites, it injects formic acid into the skin, which causes pain and irritation.

To neutralize the effect of formic acid, baking soda (sodium hydrogen carbonate) or zinc carbonate can be applied on the skin for relief.

Do you know that soil becomes acidic when an excess of chemical fertilizers are used? Plants do not grow well in acidic or basic soil. Hence, to neutralize the acidity of soil, quick lime (calcium oxide) or slaked lime (calcium hydroxide) is added to soil. To neutralize excess basicity, soils are treated with organic matter, containing organic acids.

Do You Know:

Tooth enamel, which is made of calcium phosphate, is the hardest substance in the human body. It does not dissolve in water. However, it breaks down, or disintegrates, or decays on reacting with acids. Acids are produced in the mouth due to degradation of sugar and food particles by certain bacteria.

Toothpaste, which we use daily for cleaning our teeth, is generally basic. Hence, it can neutralize excess acid present in the mouth and prevent tooth decay.

The wastes of many industries contain acids. This waste, when thrown directly into the water bodies, harms the aquatic life. Hence, this waste is first treated with basic chemicals to neutralize the effect of acids present in it.

Reactions of Metals and Non-Metals with Oxygen and Water

Metals such as aluminium, copper, and iron are widely used around us. Metals are used for the construction of bridges, automobiles, airplanes, ships, trains etc.

We have earlier studied about the physical properties of metals. Now, let us try to learn about their chemical properties. Here, we will study about the reaction of metal with oxygen, water, and acids.

You must have observed that when a piece of iron is kept in the open for some time, it gets covered with a brownish substance. This brownish substance is called **rust** and the process is called **rusting**.

Rust is formed when iron reacts with oxygen (present in air) to form iron oxide. Also, a ribbon of magnesium burns in air to form magnesium oxide. These reactions represent reactions of metals with oxygen. Hence, metals react with oxygen to produce metals oxides.

Metals react with oxygen to produce metal oxides which are basic in nature. These oxides thus turn red litmus paper blue, but have no effect on blue litmus paper.

Sulphur (S) is a non-metal. It reacts with oxygen to produce sulphur dioxide (SO_2), which is an acidic oxide. Sulphur dioxide then reacts with water to produce sulphurous acid (H_2SO_3), which changes blue litmus to red. The chemical equations involved in the reaction can be represented as:

$$\underset{\text{Sulphur}}{S} + \underset{\text{Oxygen}}{O_2} \rightarrow \underset{\text{Sulphur dioxide}}{SO_2}$$

Non-metals react with oxygen to produce their oxides, which are generally acidic in nature.

We will now study the reaction of metals and non-metals with water.

While some metals react very vigorously with water, others react very slowly. However, there are some metals which do not react with water at all. For example, sodium metal reacts vigorously with water and iron reacts slowly with water.

Metals react with water to produce hydrogen gas and metal hydroxides. These metal hydroxides are basic in nature. However, non-metals usually do not react with water.

Do You Know:

- Sodium and potassium are very reactive metals. They react vigorously with oxygen and water to produce a lot of heat. Hence, to prevent their reaction with air and water, they are stored under kerosene.

- Non-metals react very vigorously with air, but generally do not react with water. Phosphorus is a very reactive non-metal, which catches fire when exposed to air. Hence, phosphorus is stored under water to prevent contact between phosphorus and air.

1. **Reaction of metals with oxygen**

On heating, magnesium burns with a dazzling white flame to form magnesium oxide. Similarly, when aluminium is heated, it reacts with oxygen present in the air to form aluminium oxide.

$$\underset{\text{Aluminium}}{4Al(s)} + \underset{\text{Oxygen}}{3O_2(g)} \rightarrow \underset{\text{Aluminium oxide}}{2Al_2O_3(s)}$$

Almost all metals combine with oxygen to form metal oxides. The general reaction for the process is:

$$\text{Metal} + \text{Oxygen} \rightarrow \text{Metal oxide}$$

All metals are not equally reactive. Therefore, the reactivity of metals with oxygen also varies. Some metals such as sodium react with oxygen at room temperature. Metals such as magnesium do not react with oxygen at room temperature and require heating.

On the other hand, metals such as zinc do not react with oxygen easily and require very strong heating. Silver and gold do not react with oxygen even at high temperatures.

All metal oxides are basic in nature and turn red litmus paper blue. These basic oxides react with acids to form salt and water. However, the oxides of aluminium and zinc show the properties of both acids and bases.

Chemicals that show both acidic and basic properties are said to be amphoteric in nature. Hence, aluminium oxide and zinc oxide are amphoteric oxides. They react with both acids and bases to give their respective salts and water.

Almost all metal oxides are insoluble in water. However, the oxides of sodium and potassium dissolve in water to form hydroxides.

$$\underset{\text{Sodium oxide}}{Na_2O(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Sodium hydroxide}}{2\ NaOH(aq)}$$

$$\underset{\text{Potassium oxide}}{K_2O(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Potassium hydroxide}}{2\ KOH(aq)}$$

2. **Reaction of metals with water**

Do you know that sodium reacts explosively with cold water? The reaction results in the formation of their respective hydroxides and hydrogen gas. The reaction is so violent and exothermic that the evolved hydrogen catches fire. These metals give hydroxides with water as their oxides are soluble in water.

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

On the other hand, metals such as iron do not react with cold water or hot water. However, they react with steam to give their respective oxides and hydrogen gas.

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

Thus, metals react with water to form metal oxides and hydrogen gas. Some metal oxides are soluble in water. These metal oxides form hydroxides by reacting with one or more water molecules. The general reaction for the process is given as:

$$\text{Metal} + \text{Water} \rightarrow \text{Metal oxide} + \text{Hydrogen}$$
$$\text{Metal oxide} + \text{Water} \rightarrow \text{Metal hydroxide} \quad (\text{if metal oxide is soluble in water})$$

The vigour with which a metal reacts with water differs from metal to metal. Some metals react with cold water, others with hot water, while some react only with steam. There are also metals that do not even react with steam. For example, silver and gold do not react with water at all.

Reaction of non-metals with hydrogen.

Non-metals react with hydrogen under specific conditions to form their corresponding compounds containing hydrogen. Few examples are given below:

$O_2 + 2H_2 \rightarrow 2H_2O$ (Water)

$S + H_2 \rightarrow H_2S$ (Hydrogen sulphide)

$N_2 + 3H_2 \rightarrow 2NH_3$ (Ammonia)

$Cl_2 + H_2 \rightarrow 2HCl$ (Hydrogen chloride)

Unlike metals, non-metals do not react with water or dilute acids.

Reaction of Metals and Non-Metals with Acids and Bases

You know that the substances which turn blue litmus paper to red are called **acids,** and the substances which turn red litmus paper to blue are called **bases**. **Do you how these substances react with metals and non-metals?**

Let us study how metals and non-metals react with acids. The reaction of metals and non-metals with acids can be observed by performing the following experiment.

Therefore, it can be concluded that metals react with acids to release hydrogen gas, which burns with a 'pop' sound. On the other hand, non-metals do not react with acids.

Some Interesting Facts:

- Hydrogen gas is colourless and odourless. It has no effect on moist litmus paper. It burns with a characteristic 'pop' sound when a flame is introduced.

- Copper is a less reactive metal. It does not react with dilute hydrochloric acid, even on heating.

Let us now study how metals and non-metals react with bases by performing the following experiment.

Thus, metals react with bases to produce hydrogen gas. However, not all the metals react with bases to produce hydrogen gas. The reactions of non-metals with bases are complex.

- **Reaction of metals with acids**

Metals react with hydrochloric acid in the similar fashion as they do with sulphuric acid. Sodium reacts very vigorously with hydrochloric acid to form a salt, and hydrogen gas is evolved in the reaction.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{Sodium chloride}}{2NaCl(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Magnesium reacts vigorously with hydrochloric acid, but not as vigorously as sodium and potassium.

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

Zinc and iron also react with dilute hydrochloric acid to give zinc chloride and iron (II) chloride respectively. These reactions are comparatively less vigorous than the reaction of hydrochloric acid with aluminium metal.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{Zinc chloride}}{ZnCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

$$\underset{\text{Iron}}{Fe(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{Iron (II) Chloride}}{FeCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Thus, it can be concluded that metals react with acids to give a salt and hydrogen gas. The general equation for the process can be represented as:

$$\text{Metal} + \text{acid} \rightarrow \text{Salt} + \text{Hydrogen}$$

However, all metals do not react with dilute hydrochloric and sulphuric acids. Also, hydrogen gas is not evolved when a metal reacts with nitric acid.

This is because nitric acid acts as an oxidizing agent and oxidizes hydrogen gas produced in the reaction to form water.

At the same time, nitric acid itself gets reduced to form nitrogen oxides such as nitrous oxide (N_2O), nitric oxide (NO), and nitrogen dioxide (NO_2).

However, there are some metals such as magnesium, which react with very dilute nitric acid to evolve hydrogen gas.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Nitric acid}}{2HNO_3(aq)} \rightarrow \underset{\text{Magnesium nitrate}}{Mg(NO_3)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Metals such as gold and silver, which are very less reactive, do not react with acids.
The only acid that dissolves gold is *aqua regia*. *Aqua regia* is the Latin name for ‘holy water’ or ‘royal water’. It is called so because it is the only liquid that dissolves gold. It is prepared by mixing three parts of concentrated hydrochloric acid and one part of concentrated nitric acid. It is a highly corrosive and fuming solution having yellow or red colour. It can also dissolve platinum metal.

Reaction of metals with bases

When metals react with base they forms hydrogen.