Physical and Chemical Changes

Physical and Chemical Changes

You must have observed that when ice melts, it changes to water. Similarly, when we burn paper, it changes to ash. Thus, in both cases, a change is taking place. There are many changes taking place all around us. **Can we classify these changes?** All the changes can be broadly classified into two types:

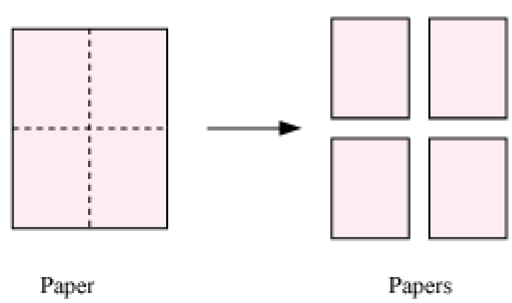
1. Physical changes

2. Chemical changes

Let us differentiate between physical changes and chemical changes.

We know that the shape, size, colour, and state of a substance are its physical properties. Physical changes usually involve changes in these properties of a substance. A change in any one of these physical properties is called a **physical change**.

For example, if you cut a piece of paper into 4 equal squares, then the shape of the paper changes, but there is no change in the properties of the paper. Also, no new substance gets formed in the process. Hence, the cutting of paper is a physical change.



In this case, we cannot join back the pieces to form the original paper. Hence, the cutting of paper is irreversible in nature. Let us now discuss a physical change, which is reversible in nature that is evaporation. **Evaporation is the process in which a liquid gets converted into its vapours**. This process depends on various factors such as,

- **Nature of liquids:** There are some liquids which evaporate quickly, such as petrol and kerosene, while there are other liquids which take some time to evaporate, such as water.
- **Surface area:** Evaporation of a liquid depends on the surface of the liquid. If surface area of the liquid increases, then evaporation increases.

- **Humidity:** It is the amount of water vapour present in the environment. If humidity or water vapour in air is high, evaporation will be slow and if air is free of water vapour then evaporation takes place rapidly.
- **Temperature:** As the temperature rises, the evaporation takes place more quickly.

Let's understand the evaporation process with the help of an illustration:

If we add a spoon of common salt in some water and stir the mixture for sometime, then the salt disappears. Now, if we place the salt solution in a china dish over a hot plate, then it will be observed that the water evaporates after sometime, leaving behind a white solid (as shown in the figure).

Evaporating salt solution

The white solid that is left after all the water is boiled is nothing but salt. This proves that when salt dissolves in water, no new substance is formed. However, this process is reversible. Thus, dissolution of salt in water is a physical change.

Hence, it can be concluded that in a physical change,

- a change in the physical properties of a substance such as state, shape, size, and colour takes place
- no new substances are formed

For example,

Ice → Water → Steam (They are all still water!)

- The original substance can generally be recovered again

Now, you know what physical changes are. **Do you know the characteristics of chemical changes?**

A chemical change is the one in which the formation of one or more new substances takes place. The new substance formed has different chemical properties from that of the substance that formed it.

Now, watch the following animation to see an example of chemical change.

Let us add more to our knowledge by performing the next activity.

When lime is added to water, the temperature of water increases and water almost starts boiling. A substance called slaked lime is produced during this change. Hence, it is a chemical change. The following chemical equation can be used to represent the chemical change.

$CaO + H_2O \rightarrow Ca\ (OH)_2 + \text{Energy}$

Lime   Water   Slaked lime

Thus, it can be concluded that in a chemical change,

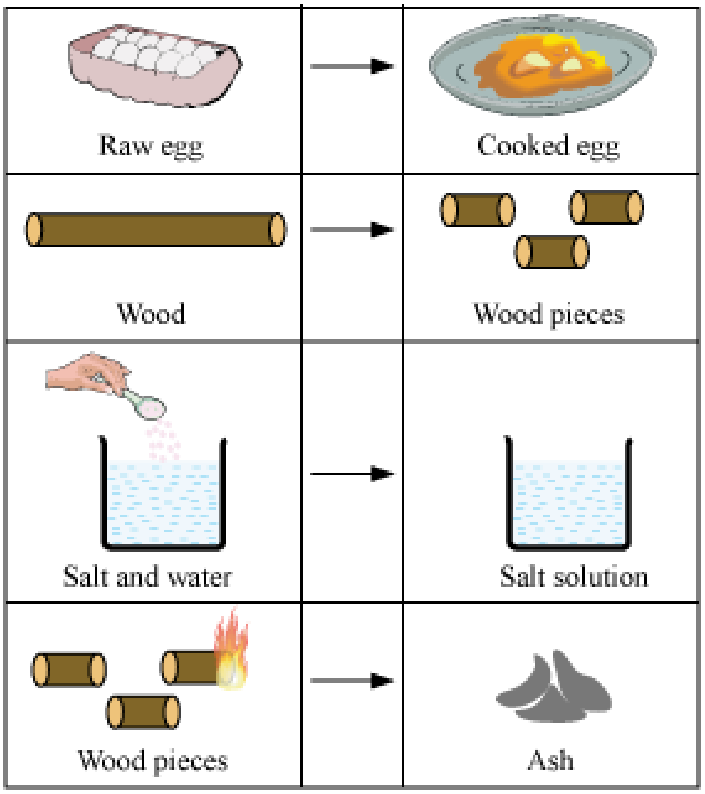
- one or more new substance(s) are formed
- the chemical properties of the new substance(s) are different from those of the starting material
- the original material cannot be recovered easily

For example, magnesium oxide and calcium hydroxide (formed in the above activities) cannot be converted back into their original substances.

Hence, we can summarize the differences between physical and chemical changes as given in the table below.

Physical Change	**Chemical Change**
1. The chemical composition of a substance does not change.	**1.** The chemical composition of a substance changes.
2. Most changes are reversible.	**2.** Most changes are irreversible.
3. No new substances are formed. For example, Ice → Water → Steam	**3.** New substances are formed. For example, Paper → Ashes

Can you specify the type of changes given in the table?



Some other examples of physical changes

- Melting of butter
- Boiling of water
- Condensation of water vapours
- Making of fruit salad with raw fruits
- Expansion or contraction of metals on heating or cooling
- Freezing of water
- Beating of metals into sheets
- Mixing of sugar and sand
- Crystallisation of salts from their solutions

Some other examples of chemical changes

- Digestion of food
- Cooking of food
- Rusting of iron
- Decaying of wood
- Burning of paper
- Souring of milk
- Ripening of mangoes
- Burning of candle

Burning a candle: We now understand the difference between physical change and chemical change. But there are a few changes in which simultaneous physical and chemical changes occur. Let us study about those changes.

When we burn a candle, heat and light is produced, which melts the candle. This process is a physical change. But at the same time, two new products, which are carbon dioxide and water vapours, are formed, making it a chemical change.

Thus, burning a candle is a combination of physical and chemical change.

- **Sublimation of ammonium chloride**

Sublimation of an element is a change from the solid directly to gas phase with no intermediate liquid stage. For example, when ammonium chloride is heated, it goes directly into the vapour state. When these vapours are cooled, ammonium chloride is obtained back.

Since we obtain the original substance back at the end, it is a physical change. **But do you know why does it sublime?**

Ammonium chloride sublimes because of dissociation of ammonium chloride into ammonia and hydrogen chloride in the vapour state. On cooling, ammonia and hydrogen chloride recombine to form ammonium chloride again.

Thus, the physical change taking place is the result of chemical dissociation and combination.

- **Cooking rice**

Do you know what happens when we cook rice?

While cooking rice, water molecules pierce the walls of the cells of the starch present in rice. Thus, some of the starch is decomposed. Therefore, this change is physical to a major extent since the composition of rice remains the same.

Allotropic Changes

Evaporating salt solution

Allotropy is exhibited by certain chemical elements, which can exist in two or more different forms, known as allotropes of that element. For example, carbon has graphite and diamond as its allotropes. Oxygen has ozone as its allotrope.

In each allotrope, the element's atoms are bonded together in a different manner. Also, they may differ in number of atoms forming the unit.

In oxygen gas, there are two atoms in a molecule. On the other hand, in ozone, three atoms are present in a molecule. Also, oxygen and ozone are different in some of their chemical properties.

Hence, we conclude that allotropic changes are chemical changes.

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.

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

- **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.

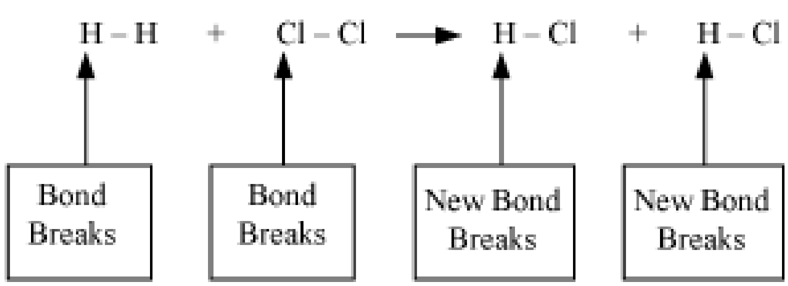
$$2\,H_2O \xleftrightarrow[\text{Current}]{\text{Electric}} 2\,H_2 + O_2$$

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.**

$$\text{Hydrogen} + \text{Chlorine} \longrightarrow \text{Hydrogen chloride}$$

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$$



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.

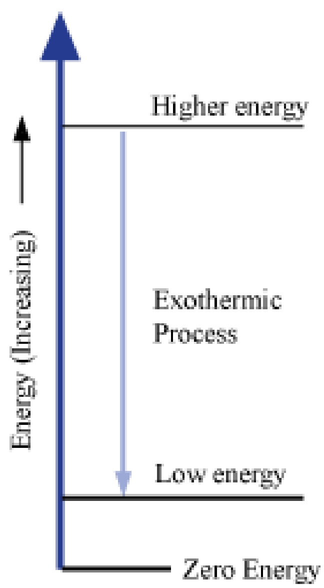
E = K.E. + P.E.

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.**

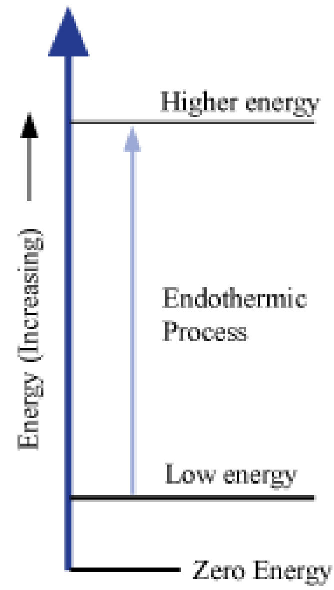
In an **exothermic reaction**, internal energy of reactants is greater than internal energy of the products and therefore, change in the energy will be negative. It is represented as:

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

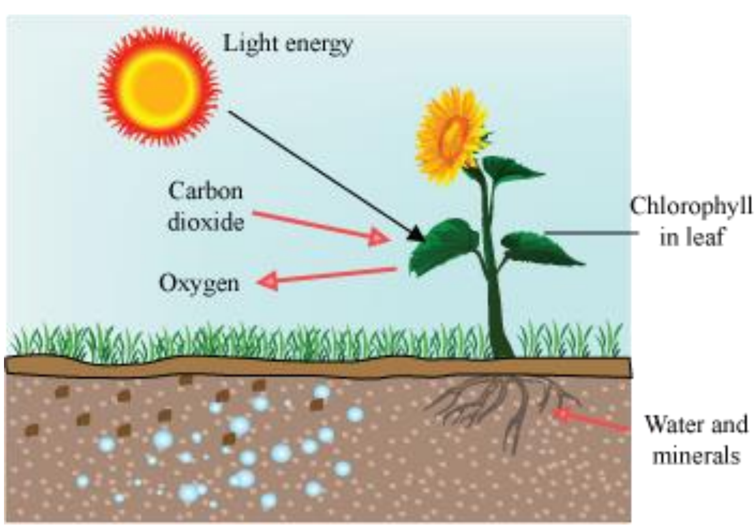


Energy released or absorbed is measured in kilocalories or kilojoules.

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.

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

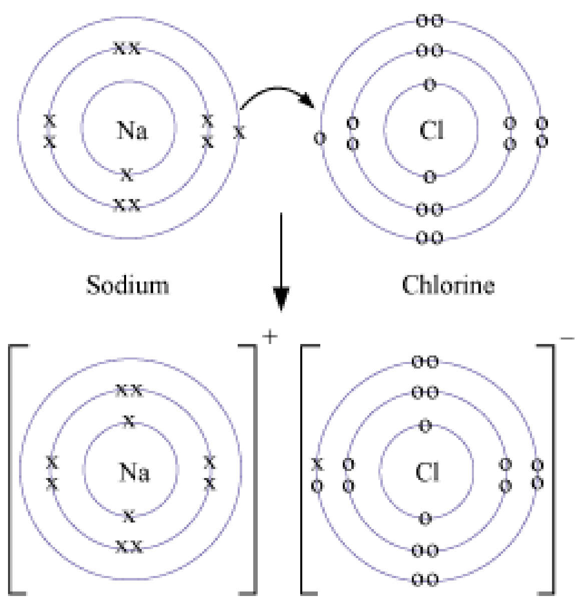


Some other examples of endothermic processes are:

- 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

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.

$$\underset{\substack{\text{Carbon} \\ \text{(From coal)}}}{C(s)} + \underset{\text{Oxygen}}{O_2(g)} \rightarrow \underset{\text{Carbon dioxide}}{CO_2(g)}$$

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.

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

Hydrogen and oxygen combine to form water.

$$\underset{\text{Hydrogen}}{2H_2\ (g)} + \underset{\text{Oxygen}}{O_2\ (g)} \rightarrow \underset{\text{Water}}{2H_2O\ (l)}$$

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:

$CaO(s)$	$+$	$H_2O(l)$	$\rightarrow$	$Ca(OH)_2(aq)$
Calcium oxide		Water		Calcium hydroxide
(Quick lime)				(Slaked lime)

Hence, in combination reactions, two or more compounds combine to produce only one product. Generally, combination reactions are exothermic in nature i.e. energy is released when two or more compounds combine. This can be explained with the help of the given activity.

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:

N_2 (g)	$+$	O_2 (g)	$\rightarrow$	$2NO$ (g)
Nitrogen		Oxygen		Nitric oxide

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.

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:

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

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. 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

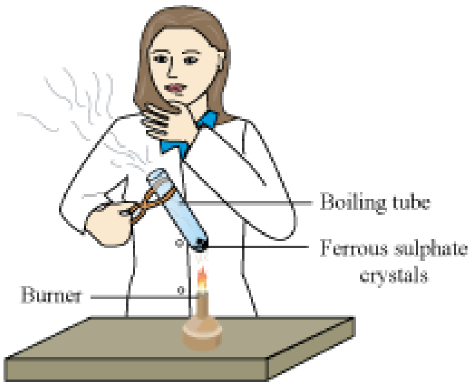
4. Double displacement reactions

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

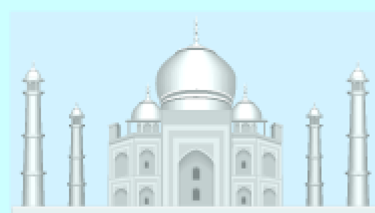
Take about 3 g of solid lead nitrate in a boiling tube. Note the colour of the compound. Heat it in the flame of the Bunsen burner. Observe the change taking place.

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.
  - $2\ NaHCO_3\,(s) \xrightarrow{\Delta} Na_2\,CO_3\,(s) + CO_2\,(g) + H_2O\,(g)$
- 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.

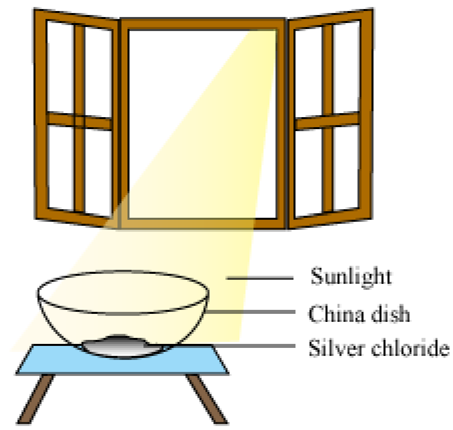


Figure 3: Photolysis of silver chloride

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

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)**

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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.

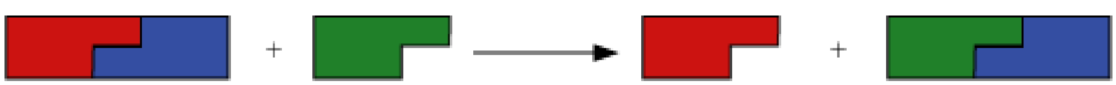
In displacement reactions, a more reactive metal replaces a less reactive metal from the latter's salt.

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.



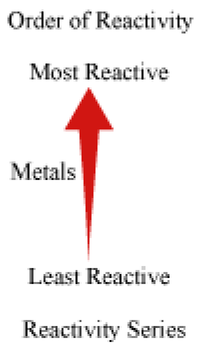
In the above figure, you have three blocks. It will be observed that while red and blue blocks are fixed in, green block is aloof. Now, if a blue block is detached from the red and fixed with the green, it will mean that the green block displaces the red block.

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:

$$\begin{array}{lllll} Zn\ (s) & + \ CuSO_4\ (aq) & \rightarrow \ ZnSO_4\ (aq) & + \ Cu\ (s) \\ \text{Zinc} & \text{Copper sulphate} & \text{Zinc sulphate} & \text{Copper} \end{array}$$

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 → 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.

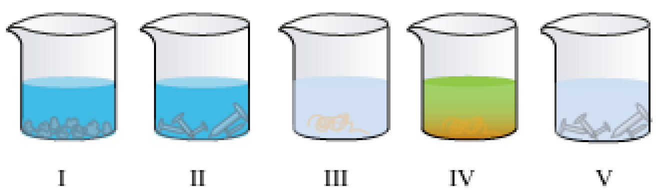
As soon as you do this, you will observe the formation of a yellow precipitate. This yellow precipitate is of lead iodide. In this reaction, the two compounds lead nitrate and potassium iodide react by exchanging their ions to form new compounds, lead iodide and potassium nitrate.

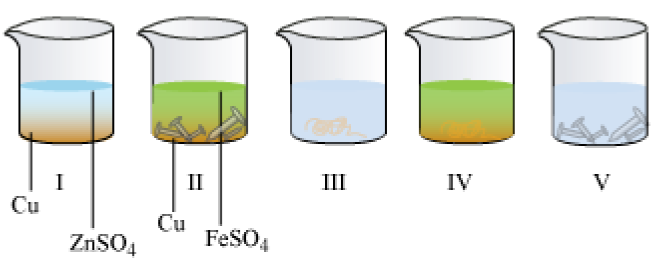
The equation involved in this reaction is:

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

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.

$$\underset{\text{Iron}}{Fe\ (s)} + \underset{\text{Copper sulphate}}{CuSO_4\ (aq)} \rightarrow \underset{\text{Iron(II) sulphate}}{FeSO_4\ (aq)} + \underset{\text{Copper}}{Cu\ (s)}$$

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

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

- **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:

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

Essential Conditions Required for Combustion

When a burning matchstick is brought near kerosene gas stove (with its knob turned on), it will be observed that the gas coming out of the stove (i.e., LPG) starts burning instantly to produce heat and light. Hence, LPG is a fuel.

Similarly, there are fuels such as wood, coal, charcoal, petrol, diesel, etc. that are used for various purposes at home, in industry, and for running vehicles. These fuels burn to produce large amounts of heat. **Thus, burning or combustion is a chemical process in which substances react with oxygen to produce heat (and sometimes light).**

Do you know that there are some substances which do not burn in air?

Glass, stone, and iron nail are examples of non-combustible substances. The substances which do not burn in air are called **non-combustible substances.**

On the other hand, the substances that undergo combustion are said to be **combustible substances** or **fuels**. Fuels may be solid, liquid, or gaseous. The light that is produced during their combustion may either be in the form of a flame or glow.

Now, we know that substances can be classified as combustible or non-combustible. **However, if a combustible substance is exposed to air, then will it start burning on its own or does a fuel require special conditions for combustion to take place?**

In our houses, there are many substances that are made of wood. Even though they are exposed to air all the time, they do not burn on their own. However, when brought in contact with a burning matchstick or candle, they start burning. This shows that some conditions are required to be fulfilled before a substance starts burning.

Let us now discuss the essential requirements for combustion. The essential conditions for combustion to take place include the presence of a fuel, air, and heat.

Now, let us watch an animation before proceeding further.

We have learnt that heat is one of the necessary requirements for combustion to take place. However, does a fuel start burning as soon as it is exposed to a source of heat?

To observe this, try burning a candle by using a gas lighter. A gas lighter provides enough heat to initiate the combustion of LPG, but not enough heat to burn a candle, paper, or wood.

You must have heard that during extreme heat in summers, dry forests catch fire and very soon the fire spreads throughout the whole area. However, forests do not burn in winters when the temperatures are low. All these examples show that different substances catch fire at different temperatures.

Therefore, we can conclude that some minimum amount of heat is required for a substance to start burning. **The lowest temperature at which a substance burns is called its ignition temperature.**

Phosphorus is a very reactive substance having an ignition temperature of 35°C. Therefore, when it is exposed to air, it starts burning as this temperature is easily attained at room temperature. Therefore, it is stored in water to cut off its contact with air.

Matchsticks contain a mixture of antimony trisulphide and potassium chloride. The rubbing surface consists of powdered glass and red phosphorus. When a matchstick is struck against a rough surface, the red phosphorus gets converted into white phosphorus, and reacts with potassium chlorate. The heat produced in this reaction ignites anitmony trisulphide, thus triggering combustion.

Thus, a combustible substance cannot catch fire as long as its temperature is lower than its ignition temperature. This can be demonstrated by performing a simple activity as shown in the following animation.

Do you know that there are some substances which have very low ignition temperatures?

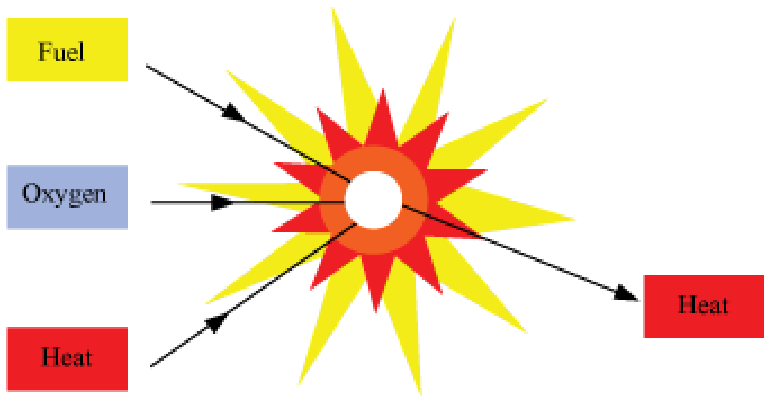
The substances which have very low ignition temperature are called **inflammable substances**. These substances catch fire very easily. LPG, petrol, and alcohol are some examples of inflammable substances. Hence, one should be very careful when using these substances.

Supporter of combustion

The gaseous environment that supports combustion of a combustible substance is called supporter of combustion. The oxygen present in air is the main supporter of combustion.

Following is the summary of the factors on which rate of combustion depends:

- Size of combustible particles: Smaller the size of combustible particles, faster is the rate of combustion.
- Nature of combustible substance: Inflammable substances burn faster as compared to substances such as wood.
- Nature of gaseous environment
- Ignition temperature of combustible substances



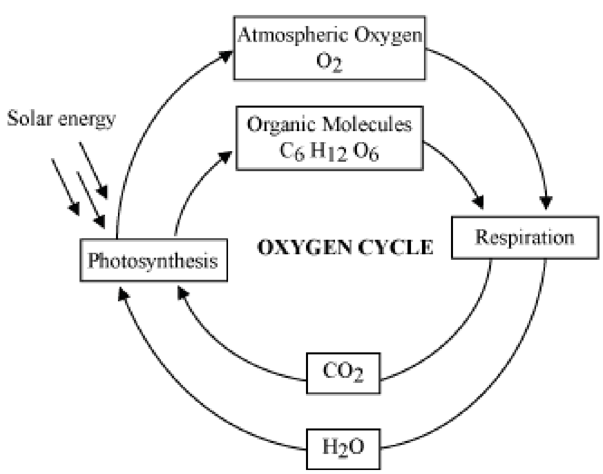
Do you know what happens if you pour water over fire?

The fire dies. This is because of the fact that air is necessary for combustion and when water is poured over the flame, air supply is cut off by the water vapours. Water also lowers the ignition temperature of the burning material below its ignition temperature.

Do you know?

Incomplete combustion occurs when there is not enough oxygen to allow the fuel to burn completely. A lot of smoke and other harmful substances are formed during incomplete combustion.

Oxygen is an important component of everyday life. We cannot survive without oxygen. It comprises about 21% of atmospheric air. It is a component of several biological molecules such as carbohydrates, proteins, nucleic acids and fats. Like carbon dioxide, **oxygen too is cycled through the process of photosynthesis and respiration**. It is also utilised during combustion or burning.



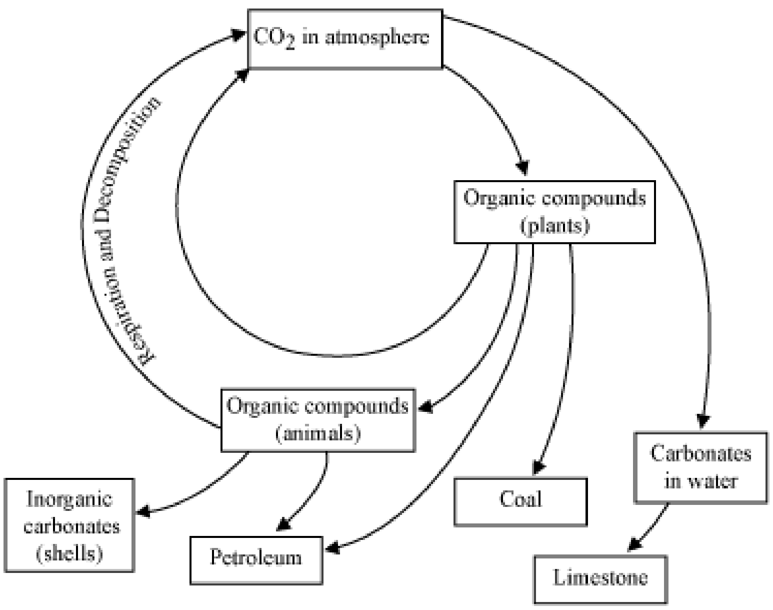
During photosynthesis, carbon dioxide and water combine to produce glucose and oxygen. This changes the atmospheric carbon into glucose molecules.

Glucose, which is a source of food, is utilised by organisms to produce energy during **respiration**. **During this process, glucose is broken down in the presence of oxygen to produce carbon dioxide and energy.**

Thus, through the processes of photosynthesis and respiration, carbon is utilised and then returned to the environment.

All organisms do not need oxygen to break down glucose and produce energy.

Combustion is another process that releases carbon dioxide. Many substances release carbon dioxide on burning. Carbon dioxide is a part of vehicular emissions and industrial fumes.



Comparison between Respiration and Burning

The following are the similarities between the two processes.

1. Both the processes utilise oxygen and liberate carbon dioxide.
2. Both the processes are exothermic in nature i.e., both the processes release energy.

The difference in the two processes is the amount of energy released during the reactions. During burning, energy is released at a much faster rate than during respiration. Also, the rate of liberation of energy can be easily controlled for respiration. Respiration occurs inside living cells, while burning occurs outside living cells.

During the process of respiration, enzymes are required and energy is trapped in ATP molecules. On the other hand, enzymes are not required and ATP molecules are not formed during the process of burning.

Do you know what happens to the mass of a substance on burning?

When magnesium is burnt in air, a white-coloured metallic oxide is formed. Now, if 3 g of magnesium is taken, then 5 g of magnesium oxide is produced. Thus, there is an increase in mass by 2 g. The increase in mass is because of the mass of oxygen that combines with magnesium to form magnesium oxide.

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

Thus, substances gain mass on burning.