1 Chemical Reactions and Equations

"When a chemical change occurs, chemical reaction is said to have taken place."

1. Introduction

Change is the law of nature.

There are so many situations of daily life, where we can observe various changes.

Like,

(i) Conversion of water into vapours from a cup of hot tea.

(ii) Corrosion of iron articles (rusting) if exposed to humid atmosphere.

(iii) Cooking of food.

(iv) Digestion of food in our body.

(v) Breaking of any article like glass.

(vi) Combustion of fuel in our vehicle.

Eggs, flour, sugar and baking powder are mixed and baked to yield a cake. This is a chemical change.

Scientist classify these changes as

(1) Physical changes (2) Chemical changes

(1) Physical changes

A change in which physical properties of a substance changes but the chemical composition do not change.

For example, Freezing, melting, boiling, condensation, etc.

Characteristic features of physical changes

(1) The identity of the substance is maintained.

(2) The change is generally temporary.

(3) Heat change may or may not take place.

(4) Only the physical state or some of the physical properties of the substances are changed.

(2) Chemical changes

A change in which one or more substances change into new substances with a different chemical composition.

For example, burning of a candle, rusting of iron, combustion of fuel, etc.

Characteristic features of chemical changes

(1) The identity of original substance is completely lost.

(2) The change is generally permanent.

(3) The change is generally accompanied by energy change.

(4) The change cannot be reversed generally.

1. From the given examples, identify the chemical changes.

(a) Fermentation of grapes.

(b) Burning of a candle.

(c) Evaporation of alcohol.

(d) Freezing of water.

(e) Turning brown of a freshly cut apple in the air.

(f) Growth of a plant.

(g) Dissolution of sugar in water.

(h) Fading of coloured clothes in the sun.

★ In earlier standards, we have seen how compounds are formed by chemical combination of element.

★ We have also learnt that the driving force behind formation of a chemical bond is to attain an electronic configuration with a complete octet.

★ The atoms attain a complete octet by giving, taking or sharing of electrons with each other.

2. Chemical reaction and its characteristics

The process in which a substance or substances undergo a chemical change to produce new substances, with entirely new properties are known as chemical reaction.

Active Chemistry 1

Magnesium burns in air to give a bright white flame

Aim

To study the reaction between magnesium and oxygen to form magnesium oxide.

Caution

Perform this activity in the presence of a teacher. It would be better to wear eye protection (as used by welders).

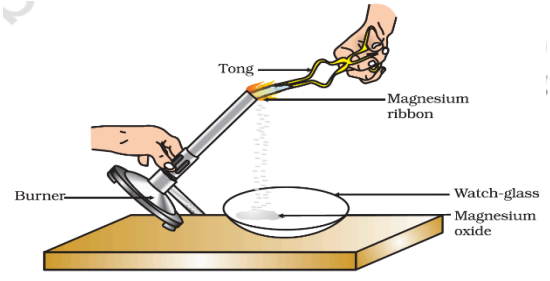
Materials Required

Burner, tong, magnesium ribbon, sand paper, watch glass.

Method

(i) Clean a magnesium ribbon about 2 cm long, by rubbing it with sand paper.

(ii) Hold it with a pair of tongs. Burn it using a spirit lamp or burner and collect the ash so formed in a watch glass. Burn the magnesium ribbon keeping it as far as possible from your eyes.



Magnesi
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Magnesi
um oxide

Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch glass.

Now answer

What do you observe?

Observation

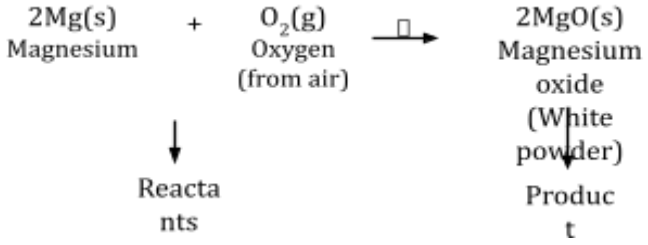
It is observed that magnesium ribbon burns with a dazzling white light and changes into a white powder. This powder is magnesium oxide.

Conclusion

Magnesium burns in air to combine with oxygen to form magnesium oxide.

Magnesium metal has a shining surface but due to attack of moist air, it is coated with a white layer of magnesium oxide. To use it for any chemical reaction, it is first rubbed with a sand paper.

SPOT LIGHT



Active Chemistry 2

Aim

To study the reaction between lead nitrate solution and potassium iodide solution.

Materials required

Test tube, lead nitrate solution, potassium iodide solution.

Method

(i) Take lead nitrate solution in a test tube.

(ii) Add potassium iodide solution to this.

Observation

It is observed that a yellow solid (precipitate) is formed.

Conclusion

Lead nitrate solution reacts with potassium iodide solution to form a yellow precipitate of lead iodide.

SPOT LIGHT

A precipitate is an insoluble solid that forms out of a liquid solution. i.e. does not dissolve in the solution.

$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI_2(s)(\downarrow) + 2KNO_3(aq)$

Lead nitrate (Colourless)	Potassium iodide (Colourless)	Lead (II) iodide (Yellow ppt.)	Potassium nitrate (Colourless)

Active Chemistry 3

Aim

To study the reaction between zinc and sulphuric acid or hydrochloric acid.

Materials required

Conical flask or test tube, zinc granules, dilute hydrochloric acid or sulphuric acid.

Method

(i) Take a few zinc granules in a conical flask or a test tube.

(ii) Add dilute hydrochloric acid or sulphuric acid to this.

Be Alert !

★ Handle the acid with care because acid is corrosive in nature.

Now answer

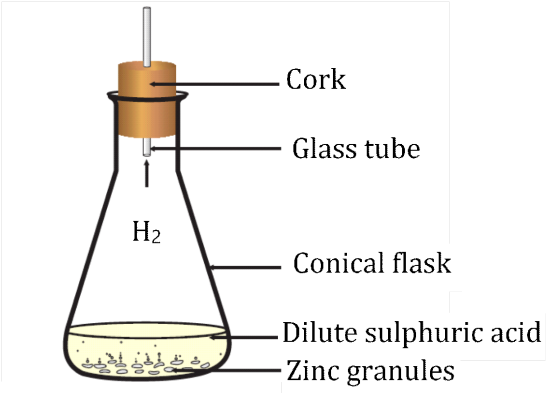
Touch the conical flask or test tube. Is there any change in the temperature?

Observation and discussion

Bubbles of hydrogen gas are found to rise briskly from the surface of zinc pieces. The gas evolved can be tested by bringing a lighted candle. It is found to burn with a popping sound. Further, the flask is found to be hot.

Conclusion

Zinc reacts with dilute sulphuric acid or hydrochloric acid with the evolution of hydrogen gas and heat (i.e., reaction is exothermic).



Formation of hydrogen gas by the action of dilute sulphuric acid on zinc.

$Zn(s) + dil.H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)(\uparrow) + Heat$

Zinc Sulphuric acid Zinc sulphate Hydrogen (gas)

With the help of above activities, we have seen that whenever a chemical reaction has taken place, some changes are observed. These observable changes are called the characteristics of a chemical reaction.

Characteristics of chemical reactions

(i) Change in state: The physical state of the substances normally changes.

For example,

Formation of solid MgO from solid Mg and gaseous O_2.

$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$

Magnesium Oxygen (from air) Magnesium oxide (White powder)

(ii) Change in colour: In some of the chemical reactions, change in colour can be observed.

For example,

(a) Formation of reddish-brown rust on grey iron nails.

(b) Formation of yellow ppt. of lead iodide from colourless solution of $Pb(NO_3)_2$ and KI.

(iii) Evolution of a gas: In some cases, a gas may be evolved.

For example,

Evolution of H_2 gas, in the reaction between Zn and dil HCl.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2(g)\uparrow$$

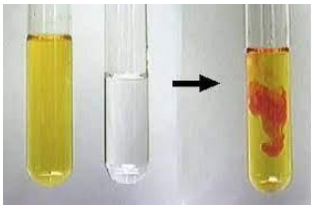
(iv) Change in temperature: Most of the reactions are accompanied by temperature change i.e. increase or decrease in temperature.

For example,

(a) In the reaction between Zn and H_2SO_4, flask was found to be warm. Thus, rise in temperature has taken place (Exothermic).

(b) If a reaction between barium hydroxide ($Ba(OH)_2$) and ammonium chloride, (NH_4Cl) is carried out in a test tube, it is observed that the bottom of the test tube becomes cooler (Endothermic).

Some clues that a chemical reaction has occurred



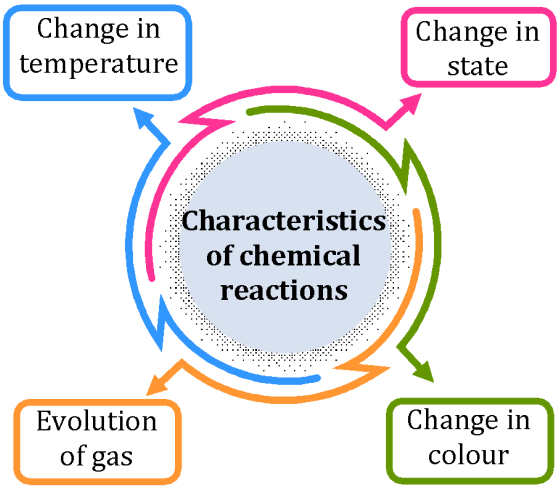
Colour change

Solid forms

Bubble forms

Heat or flame occurs

Quick Tips



Check your Answers 1

1. Chemical changes are a, b, e, f and h.

Do You Remember ?

★ **Here is some important chemical formulae which you have learnt in your previous class that can help you in writing a chemical equation.**

	Name of the Compound	Formula
1.	Calcium hydroxide (Lime water) (Slaked lime)	$Ca(OH)_2$
2.	Calcium carbonate (Marble, lime stone)	$CaCO_3$
3.	Calcium oxide (Quick lime)	CaO
4.	Iron (II) oxide	FeO
5.	Iron (III) oxide	Fe_2O_3
6.	Magnesium oxide	MgO
7.	Lead nitrate	$Pb(NO_3)_2$
8.	Lead iodide	PbI_2
9.	Potassium iodide	KI
10.	Ammonium chloride	NH_4Cl
11.	Ammoniun hydroxide	NH_4OH
12.	Potassium chlorate	$KClO_3$
13.	Lead oxide	PbO
14.	Aluminium oxide	Al_2O_3
15.	Silver chloride	$AgCl$

16.	Cuprous oxide	Cu_2O
17.	Cupric oxide	CuO
18.	Silver nitrate	$AgNO_3$
19.	Barium chloride	$BaCl_2$
20.	Barium sulphate	$BaSO_4$

3. Writing a chemical equation

Whenever a chemical reaction takes place, a number of steps are involved. Thus, description of chemical reaction in sentence form becomes long. So, some short-hand representation of chemical reaction is followed.

It can be done in two ways:

(a) Word equation

Chemical equation which represents a chemical reaction briefly in words is called word equation.

For example, the word equation can be written as:

Magnesium + Oxygen $\longrightarrow$ Magnesium oxide

(b) Chemical equation

A chemical equation is a short-hand method that describes a chemical reaction in terms of symbols and formulae of different reactants and products.

For example, the chemical equation can be written as

$Mg + O_2 \rightarrow MgO$

Rules for writing a chemical equation

(i) The symbols and formulae of the reactants are always written on the L.H.S. (left hand side) of arrow and a plus **(+)** sign is put between them.

(ii) The symbols and formulae of the products are always written on the R.H.S. (right hand side) of arrow and a plus (+) sign is put between them.

(iii) An arrow ($\longrightarrow$) sign is put between the reactants and the products, pointing from reactants towards products.

For example,

$Na + H_2O \rightarrow NaOH + H_2$

A chemical equation expressed in symbols and formulae, such that the number of atoms of different elements towards the side of the reactants is not equal to the number of atoms of different elements towards the side of the products, is called skeletal equation or unbalanced equation.

To make this equation meaningful, this equation is balanced according to law of conservation of mass, then it is called balanced chemical equation.

Check your Concepts 2

1. Write chemical equations for the following word equations.

(a) Potassium chlorate ⟶ Potassium chloride + Oxygen

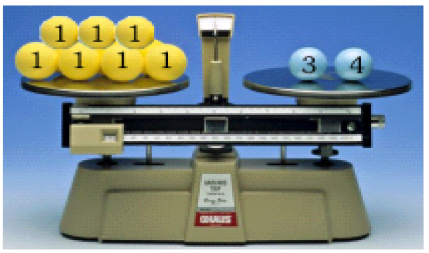
(b) Methane + Oxygen ⟶ Carbon dioxide + Water

(c) Nitrogen + Hydrogen ⟶ Ammonia

(d) Calcium carbonate ⟶ Calcium oxide + Carbon dioxide

(e) Graphite + Oxygen ⟶ Carbon dioxide

(f) Quicklime + Water ⟶ Calcium hydroxide



Mass of reactants = Mass of the products
Balanced chemical equation makes calculations easy.

How to balance an unbalanced chemical equation

According to the law of conservation of mass, the total mass of products must be equal to the total mass of the reactants (as mass can neither be created nor destroyed). This is possible only if the number of atoms of each element is equal on the two sides of the equation.

Balancing of a chemical equation means making the number of atoms of each element equal on both sides of the equation.

(i) To understand this, let us consider the following word equation for

Zinc + Sulphuric acid ⟶ Zinc sulphate + Hydrogen

Chemical equation for the above word equation will be,

Zn + H_2SO_4 ⟶ $ZnSO_4$ + H_2

Let us examine the number of atoms of different elements on both the sides of the arrow.

Element	Number of atoms in L.H.S.	Number of atoms in R.H.S.
Zn	1	1
H	2	2
S	1	1
O	4	4

As the number of atoms of each element is same on both sides of the arrow, the equation can be said a balanced chemical equation.

(ii) Now consider another chemical equation.

$$Fe + O_2 + H_2O \longrightarrow Fe_2O_3.H_2O$$

On counting number of different atoms on both the sides of the arrow, we will find that, this equation is not balanced.

Element	Number of atoms in L.H.S.	Number of atoms in R.H.S.
Fe	1	2
O	3	4
H	2	2

Let us learn about balancing a chemical equation step by step.

Step-I : Write the word equation for the given chemical reaction.

Step-II : Convert the formed word equation in the chemical equation (Skeletal chemical equation).

Step-III : Formula of each reactant and product has to be enclosed in boxes, so that during balancing, formula of substances cannot be changed.

Step-IV : Listing of number of reactants and products is done.

Step-V : Balancing should be started with compound which have biggest formula.

Step-VI : Then, balancing of different elements is done one by one.

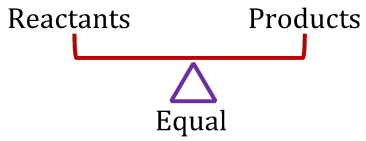
Step-VII : Finally, the equation is checked.

The method of balancing the equation by using smallest natural number coefficients is called as hit and trial method.

SPOT LIGHT

(iii) Let us balance some of the chemical reactions by following the above steps.

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



Law of conservation of mass

Step-I : Word equation

Magnesium + Hydrochloric acid $\longrightarrow$ Magnesium chloride + Hydrogen

Step-II : Chemical equation,

$Mg + HCl \longrightarrow MgCl_2 + H_2$

Step-III : Enclose all formulae into boxes.

$\boxed{Mg} + \boxed{HCl} \longrightarrow \boxed{MgCl_2} + \boxed{H_2}$

Step-IV : Count the number of atoms for all elements.

Element	Number of atoms in L.H.S.	Number of atoms in R.H.S.
Mg	1	1
Cl	1	2
H	1	2

We can see that number of Mg is same on both sides but Cl and H atoms differs on both sides.

Step-V : As the number of atoms is deficient at reactant side, let's begin from here. At reactant side HCl is the bigger formula, so we will start with it.

Step-VI : Put coefficient 2 before HCl to make chlorine equal to reactant side.

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

Here, we can see that H, automatically gets balanced.

In a balanced chemical equation, an integer precedes the formula of each substance. This number is known as stoichiometric coefficient.

SPOT LIGHT

Step-VII : Now, check the number of atoms of different elements on both sides of the equation. These are equal. This means that the equation is balanced.

Balancing of equations

Steam is passed over red hot iron to form Iron (II, III) oxide and hydrogen in presence of air.

Step-I : Iron + Steam $\longrightarrow$ Iron (II, III) Oxide + Hydrogen

Step-II : $Fe + H_2O \longrightarrow Fe_3O_4 + H_2$

Step-III : $\boxed{Fe} + \boxed{H_2O} \longrightarrow \boxed{Fe_3O_4} + \boxed{H_2}$

Element	Number of atoms in L.H.S.	Number of atoms in R.H.S.
Fe	1	3
H	2	2
O	1	4

Step-IV : Formula selected is Fe_3O_4 to start balancing oxygen,

(i) To balance O-atoms, multiply H_2O in LHS by 4.

$$\boxed{Fe} + 4\boxed{H_2O} \longrightarrow \boxed{Fe_3O_4} + \boxed{H_2}$$

(ii) Now balance Fe atoms.

$$3\boxed{Fe} + 4\boxed{H_2O} \longrightarrow \boxed{Fe_3O_4} + \boxed{H_2}$$

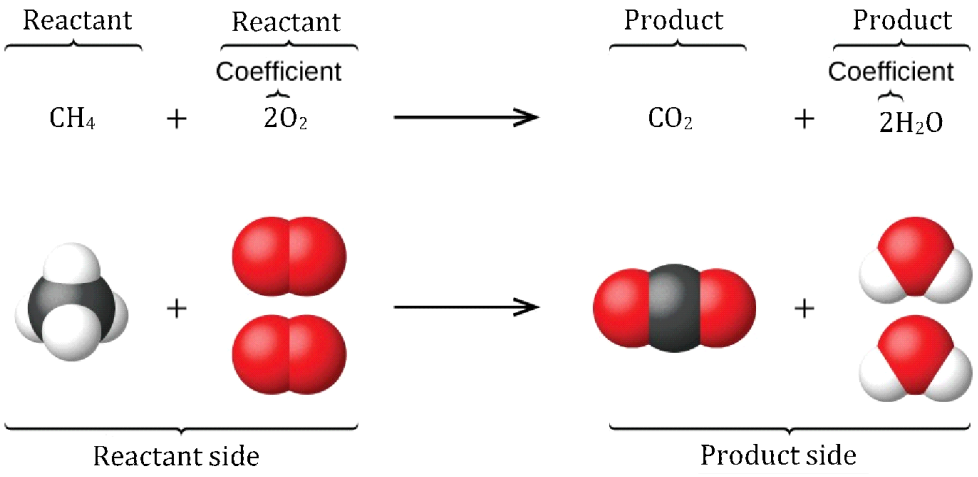
(iii) Balance H atoms.

$$3\boxed{Fe} + 4\boxed{H_2O} \longrightarrow \boxed{Fe_3O_4} + 4\boxed{H_2}$$

Note: Mixture of ferrous oxide (FeO) and ferric oxide (Fe_2O_3) is also known as magnetic oxide of iron.

Step-V : On checking the number of all elements, we found that equation is balanced now.

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$



The reaction between methane and oxygen.

Check your Answers 2

1. (a) $2KClO_3 \rightarrow 2KCl + 3O_2$

(b) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

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

(d) $CaCO_3 \rightarrow CaO + CO_2$

(e) $C + O_2 \rightarrow CO_2$

(f) $CaO + H_2O \rightarrow Ca(OH)_2$

Building Concepts 1

Write the balanced equation for the reaction involving the combustion of methane in presence of oxygen to form carbon dioxide and water.

Explanation

Step-I : Methane + Oxygen ⟶ Carbon dioxide + Water

Step-II : $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

Step-III : $\boxed{CH_4} + \boxed{O_2} \longrightarrow \boxed{CO_2} + \boxed{H_2O}$

Step-IV :

Element	Number of atoms in L.H.S.	Number of atoms in R.H.S.
C	1	1
H	4	2
O	2	3

Step-V : Balancing different elements

(i) C is already balanced.

(ii) For H, place 2 before H_2O in R.H.S.

$\boxed{CH_4} + \boxed{O_2} \longrightarrow \boxed{CO_2} + 2\boxed{H_2O}$

(iii) For O, as after (ii) number of oxygen atoms becomes 4 in R.H.S. therefore, place 2 before O_2 in L.H.S.

$\boxed{CH_4} + 2\boxed{O_2} \longrightarrow \boxed{CO_2} + 2\boxed{H_2O}$

Step-VI : Check the correctness of the balanced equation

Element	Number of atoms in L.H.S.	Number of atoms in R.H.S.
C	1	1
H	4	4
O	4	4

Hence, the equation is balanced.

Making chemical equations more informative

On examining a balanced chemical equation, we observe that it does not give any information about the physical state of the reactant and product.

Let's make balanced chemical equation more informative by following instructions.

(1) Writing symbols for the physical state of reactants and the products,

(s) for solid state

(ℓ) for liquid state

(g) for gaseous state

(aq) for aqueous solution i.e. solution prepared in water.

(2) Sometimes a gas evolved in a reaction is shown by the symbol (↑) i.e. by an arrow pointing upwards.

Similarly, the precipitate if formed during the reaction is indicated by the symbol (↓) i.e. by an arrow pointing downwards.

The abbreviation 'ppt' is also used to represent the precipitate if formed.

(3) The conditions of temperature, pressure and the presence of catalyst, if any, may be represented by writing these conditions above and / or below the arrow (⟶) between the reactants and the products.

Few examples,

$CO(g) + 2H_2(g) \xrightarrow{340\ atm} CH_3OH\ (g)$

$6CO_2(g) + 6H_2O(g) \xrightarrow[Chlorophyll]{Sunlight} C_6H_{12}O_6(aq) + 6O_2(g)$

$2Na(s) + 2H_2O(g) \longrightarrow 2NaOH(aq) + H_2(g)\ or\ H_2\uparrow$

$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(\downarrow)\ (or\ ppt) + H_2O(g)$

$C(s) + O_2(g) \longrightarrow CO_2(g) + heat$

$N_2(g) + O_2(g) + heat \longrightarrow 2NO(g)$

Check your Concepts 3

Write the balanced chemical equations for the following reactions

(a) Phosphorus burns in oxygen to form phosphorus pentaoxide.

(b) Marble (calcium carbonate) dissolves in hydrochloric acid to give calcium chloride, water and carbon dioxide.

(c) Ethyne (C_2H_2) gas burns in oxygen to form carbon dioxide and water along with evolution of heat.

(d) Liquid hydrogen peroxide decomposes to form water and oxygen gas.

Building Concepts 2

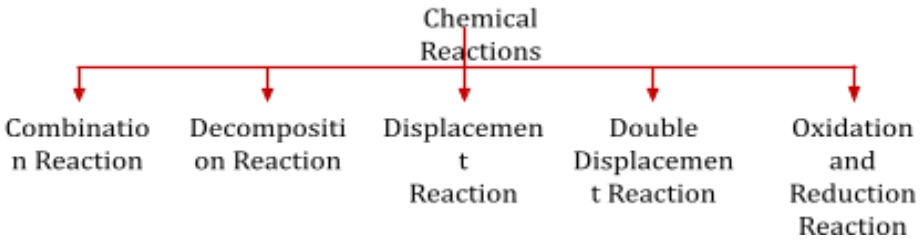
Why is it necessary to balance the chemical equation?

Explanation

In order to make law of conservation of mass applicable on the given chemical equation, it is necessary to make number of atoms of all elements equal in L.H.S. and R.H.S. Thus, balancing of equation is required.

4. Types of chemical reactions

Accordingly, the reactions are classified in different types.



(i) Combination Reaction

The reactions in which two or more substances combine to form a single new substance are called combination reaction.

Active Chemistry 4

Aim

To study combination reaction between calcium oxide, i.e., quick lime and water.

Materials required

Beaker, water, calcium oxide or quick lime

Method

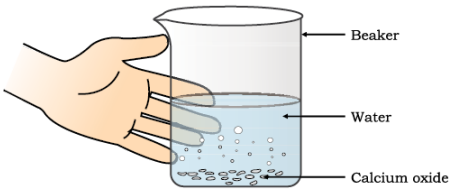
(i) Take a small amount of calcium oxide or quick lime in a beaker.

(ii) Slowly add water to this.

When quick lime (CaO) was added to water, formation of slaked lime [$Ca(OH)_2$] has taken place, with the evolution of a large amount of heat. i.e. its an exothermic reaction.

SPOT LIGHT

(iii) Touch the beaker.



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Calcium
oxide

Formation of slaked lime by the reaction of calcium oxide with water.

Now answer

Do you feel any change in temperature?

Observation and discussion

A vigorous reaction is found to occur and the beaker is found to become very hot.

Conclusion

Calcium oxide combines with water to form calcium hydroxide (slaked lime) and this reaction is highly exothermic.

$CaO(s) + H_2O(\ell) \longrightarrow Ca(OH)_2(aq) + Heat$

Calcium oxide (Quick lime) Water Calcium hydroxide (Slaked lime)

Which solution is used for white washing of walls and which compound gives a shiny finish to the walls?

Explanation

A solution of slaked lime produced by the reaction above is used for white washing of walls. Calcium hydroxide reacts slowly with the carbon dioxide in the air to form a thin layer of calcium carbonate on the walls. Calcium carbonate is formed after two to three days of white washing and gives a shiny finish to the walls.

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

Some more examples of combination reactions

(i) Burning of Coal

$C(s) + O_2(g) \longrightarrow CO_2(g)$

Carbon Oxygen Carbon dioxide

(ii) Formation of Water

$2H_2(g) + O_2(g) \longrightarrow 2H_2O(\ell)$

Hydrogen Oxygen Water

(iii) Burning of Magnesium in air

$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

Magnesium Oxygen Magnesium oxide

(iv) Formation of Iron sulphide

$Fe(s) + S(s) \longrightarrow FeS(s)$

Iron Sulphur Iron sulphide

(v) Formation of Ammonium chloride

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

Ammonia Hydrogen Chloride Ammonium Chloride

(vi) Formation of Calcium Carbonate

$CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$

Calcium oxide (Quick lime) Carbon dioxide Calcium carbonate

(vii) Reaction of carbon monoxide with oxygen

$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$

This is also an exothermic reaction.

(viii) $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

How the process of respiration is said exothermic?

Explanation

$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(\ell)$ + energy

Glucose Oxygen Carbon dioxide Water

Since energy is liberated, we can say that respiration is an exothermic reaction.

(ii) Decomposition Reaction

The reaction in which a single compound breaks up into two or more simpler substances is known as decomposition reaction. The decomposition reaction generally takes place when energy in some form such as heat, electricity or light is supplied to the reactants.

Types of decomposition reaction

(a) Thermal decomposition: The reaction in which a single compound breaks up into two or more simpler substances by the action of heat is called thermal decomposition reaction.

Following activity shows thermal decomposition reaction:

Active Chemistry 5

Aim

To study the decomposition of ferrous sulphate on heating.

Materials required

Burner, boiling tube, tong, ferrous sulphate crystals

Method

(i) Take about 2g ferrous sulphate crystals in a dry boiling tube.

(ii) Note the colour of ferrous sulphate crystals.

(iii) Heat the boiling tube over the flame of a burner or spirit lamp.

(iv) Observe the colour of the crystals after heating.

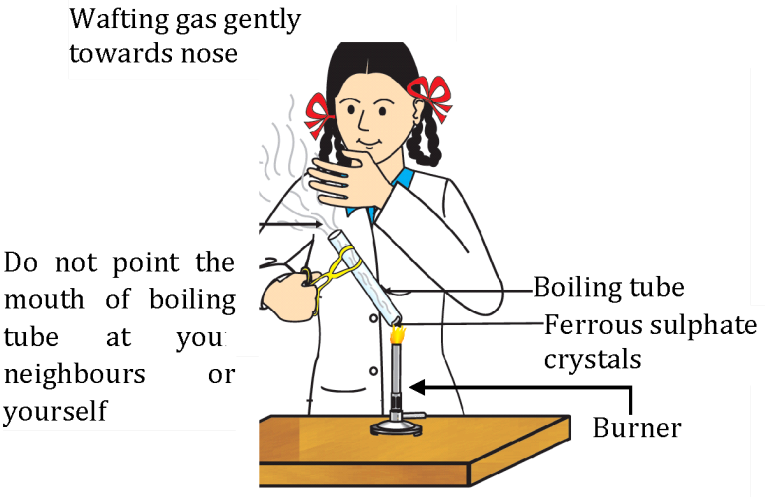
SPOT LIGHT

A biochemical reaction is the transformation of one molecule to a different molecule inside a cell. For example, Respiration, digestion, etc.

Be Alert !

★ Always remember not to point the mouth of boiling tube at your neighbours or yourself.

Correct way of holding the boiling tube is given below.



Correct way of heating the boiling tube containing crystals of ferrous sulphate and of smelling the odour.

Observation

It is observed that green coloured ferrous sulphate crystals ($FeSO_4.7H_2O$) on heating first change colour by losing water to form $FeSO_4$ which on further heating decomposes to leave behind a reddish-brown residue along with evolution of sulphur dioxide and sulphur trioxide gases.

Conclusion

The reddish-brown residue is of ferric oxide. Hence, the following decomposition reaction takes place:

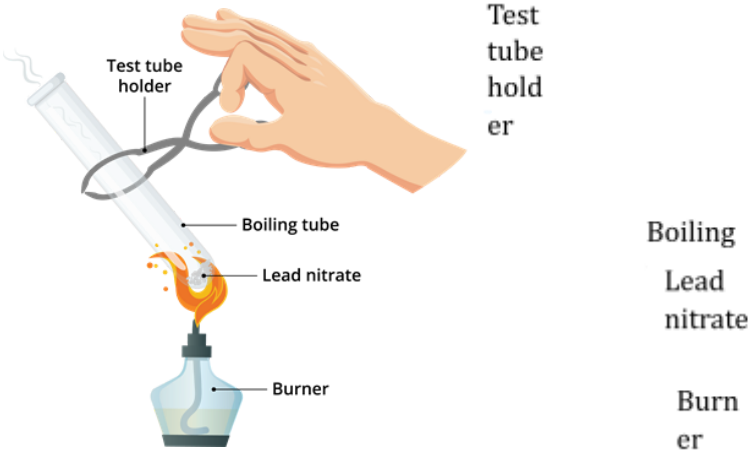
$$FeSO_4.7H_2O(s) \xrightarrow{Heat} FeSO_4(s) + 7H_2O(l)$$

Ferrous sulphate crystals (Green) — Anhydrous ferrous sulphate (White)

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

Anhydrous ferrous Sulphate (White) — Ferric oxide (Reddish Brown) — Sulphur dioxide — Sulphur trioxide

SO_2 is a colourless gas which smells like burnt matches and turns moist blue litmus paper red. It also turns acidified potassium dichromate solution from orange to green. It is collected by downward displacement of water and is a pungent smelling gas.



Heating of lead nitrate and emission of nitrogen dioxide

Active Chemistry 6

Aim

To study the thermal decomposition of lead nitrate.

Materials required

Burner, test tube holder/tong, boiling tube, lead nitrate powder

Method

(i) Take about 2 g lead nitrate powder in a boiling tube.

(ii) Hold the boiling tube with a pair of tongs and heat it over the flame.

Observation

Brown fumes of nitrogen dioxide (NO_2) are found to evolve and a yellow residue is left in the test tube.

Conclusion

Solid lead nitrate decomposes on heating to give out brown fumes of NO_2 and a yellow residue of lead (II) oxide.

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

Lead nitrate	Lead (II) oxide	Nitrogen dioxide	Oxygen
(white)	(Yellow)	(Brown fumes)	

In the above two activities, it is seen that chemical reaction has taken place on absorption of heat. These type of decomposition reactions are called thermal decomposition reactions.

Some more examples of thermal decomposition reactions.

(i) Decomposition of limestone on heating

$$CaCO_3 \xrightarrow{Heat} CaO + CO_2(g)$$

Limestone (Calcium carbonate)	Quick lime (Calcium oxide)	Carbon dioxide

(ii) Decomposition of zinc carbonate on heating

$$ZnCO_3 \xrightarrow{Heat} ZnO + CO_2(\uparrow)$$

Zinc carbonate	Zinc oxide	Carbon dioxide

(b) Electric Decomposition: The reaction in which a single compound breaks up into two or more simpler substances by the action of electricity is called electric decomposition reaction.

Following activity shows electric decomposition reaction:

Aim

To study electrolytic decomposition, i.e., electrolysis of water.

Materials required

Plastic mug, rubber stoppers, carbon electrodes, battery, water, dilute sulphuric acid, test tubes, candle.

Method

(i) Take a plastic mug. Drill two holes at its base and fit rubber stoppers in these holes. Insert carbon electrodes in these rubber stoppers.

(ii) Connect these electrodes to a 6-volt battery.

(iii) Fill the mug with water such that the electrodes are immersed. Add a few drops of dilute sulphuric acid to the water.

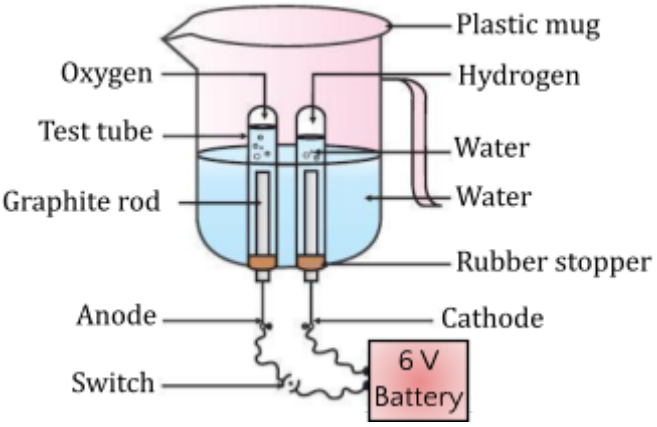
(iv) Take two test tubes filled with water and invert them over the two carbon electrodes.

(v) Switch on the current and leave the apparatus undisturbed for some time.

(vi) It will be observed that formation of bubbles takes place at both the electrodes. These bubbles displace water in the test tubes.

(vii) Once the test tubes are partially filled with the respective gases, remove them carefully.

(viii) Test these gases one by one by bringing burning candle close to the mouth of the test tubes.



Electrolysis of water

Now answer

(i) Is the volume of the gas collected the same in both the test tubes?

(ii) What happens in each test tube on bringing close to the burning candle?

(iii) Which gas is present in each test tube?

Observation and discussion

It is observed that the volume of gas collected over the cathode is double than that collected over the anode. The gas with double volume burns with a popping sound whereas the other gas supports burning (combustion). Thus, the gas with double the volume is hydrogen whereas the gas in the other tube is oxygen.

Conclusion

Acidified water undergoes electrolysis producing H_2 and O_2 gases in the ratio of 2 : 1 by volume

$$2H_2O(l) \quad Electrolysis \rightarrow \quad 2H_2(g) \quad + \quad O_2(g)$$

Water Hydrogen Oxygen

(Burns with a popping sound) (Supporter of combustion)

These kind of reactions in which electric current is passed through the compound in liquid (or molten) or aqueous solution are called electrolytic decomposition reactions or simply electrolysis.

Some more examples of electrolysis.

(i) Electrolysis of molten sodium chloride.

$$2NaCl \xrightarrow[\text{Current}]{\text{Electric}} 2Na + Cl_2(\uparrow)$$

Sodium chloride (molten) Sodium metal Chlorine

(ii) Electrolytic decomposition of molten alumina.

$$2Al_2O_3 \xrightarrow[\text{Current}]{\text{Electric}} 4Al + 3O_2(\uparrow)$$

Alumina (molten) Aluminium Oxygen

(c) Photo decomposition: The reaction in which a single compound breaks up into two or more simpler substances by the action of light is called photo decomposition reaction.

Note: Photolysis of AgBr is used in black and white photography.

Following activity shows photo decomposition reaction.

Active Chemistry 8

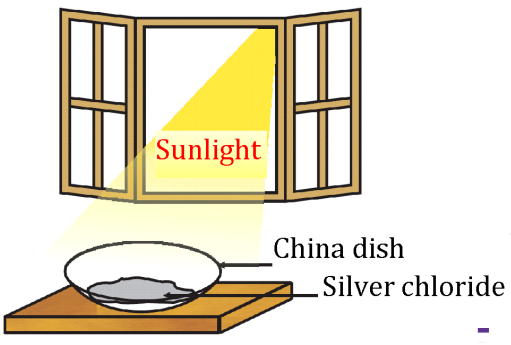
Aim

To study photo-decomposition of silver chloride.

Materials required

China dish, silver chloride

Method



Photochemical decomposition of silver chloride to grey silver metal

(i) Take about 2 g of silver chloride in a china dish. Note its colour.

(ii) Place this china dish in sunlight for some time.

(iii) Observe the colour of the silver chloride after some time.

Observation

It is observed that white crystals of silver chloride turn grey in the sunlight.

Conclusion

Silver chloride decomposes into grey silver and chlorine in the presence of light.

$$2AgCl(s) \quad Sunlight \rightarrow 2Ag(s) + Cl_2(g)$$

Silver chloride	Silver	Chlorine
(white)	(grey)	(Greenish yellow)

These types of decomposition reactions which take place on absorption of light are called photo decomposition reactions or simply photolysis.

Some more examples of photolysis

(i) Photolysis of silver bromide

$$2AgBr(s) \xrightarrow{\text{Sunlight}} 2Ag(s) + Br_2(g)$$

(Yellow) (Grey) (Reddish brown)

(ii) Photolytic decomposition of hydrogen peroxide.

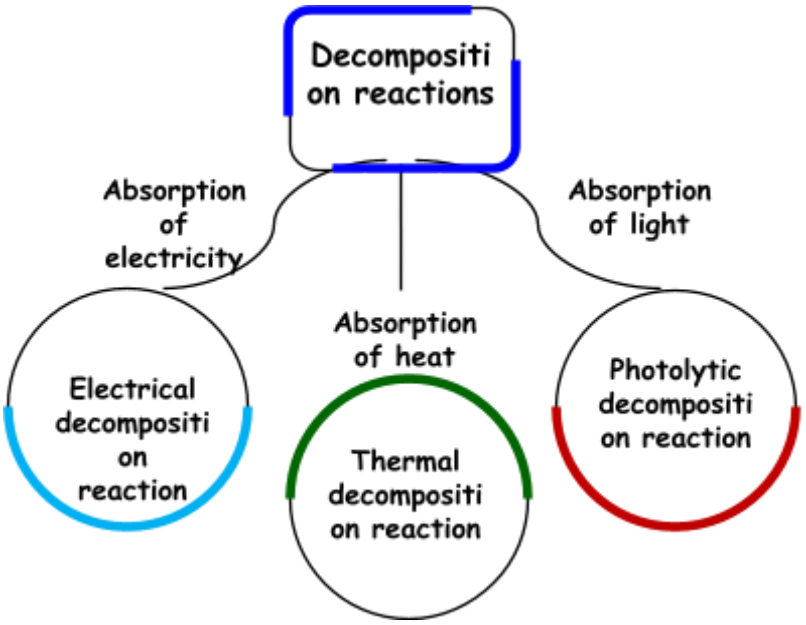
$$2H_2O_2 \xrightarrow{\text{Light}} 2H_2O + O_2(\uparrow)$$

Hydrogen peroxide Water Oxygen

Here, we can see that all decomposition reactions need energy in one or other form.

Reactions which need energy (or energy is absorbed) are called endothermic reactions.

Note: Hydrogen peroxide is kept in amber coloured bottles so as to cut off light.



Building Concepts 5

If 2 g of barium hydroxide is added to 1 g of ammonium chloride, in a test tube, it was observed that bottom of test tube becomes cold. Give an explanation.

Explanation

When barium hydroxide is added to ammonium chloride, an endothermic reaction takes place. Since heat is absorbed, so the bottom of the test tube becomes cold.

$$Ba(OH)_2 + 2NH_4Cl \xrightarrow{\Delta} BaCl_2 + 2NH_4OH$$

(iii) Displacement reactions

A reaction in which a more reactive element displaces a less reactive element from its compounds is called displacement reaction. The elements involved may be metals or non-metals.

Relative reactivity of metals

Different metals possess different reactivities. The arrangement of metals in a vertical column in order of their decreasing reactivity from top to bottom is called reactivity series or activity series of metals.

Quick Tips

★ Here is a tip to memorise the activity series of some metals which can help you to perform displacement reaction.

Element	Symbol	Trick
Potassium	K	Please
Sodium	Na	Stop
Calcium	Ca	Calling
Magnesium	Mg	Me
Aluminium	Al	A
Zinc	Zn	Zebra
Iron	Fe	Instead
Tin	Sn	Try
Lead	Pb	Learning
Hydrogen	H	How
Copper	Cu	Copper
Mercury	Hg	Merc
Silver	Ag	Saves
Gold	Au	Gold

Most reactive

DECREASING REACTIVITY

Least reactive

Activity series of some metals

Aim

To study displacement of copper from copper sulphate solution by iron.

Materials required

Test tubes, iron nails, sand paper, copper sulphate solution, test tube stand.

Method

(i) Take three iron nails and clean them by rubbing with sandpaper.

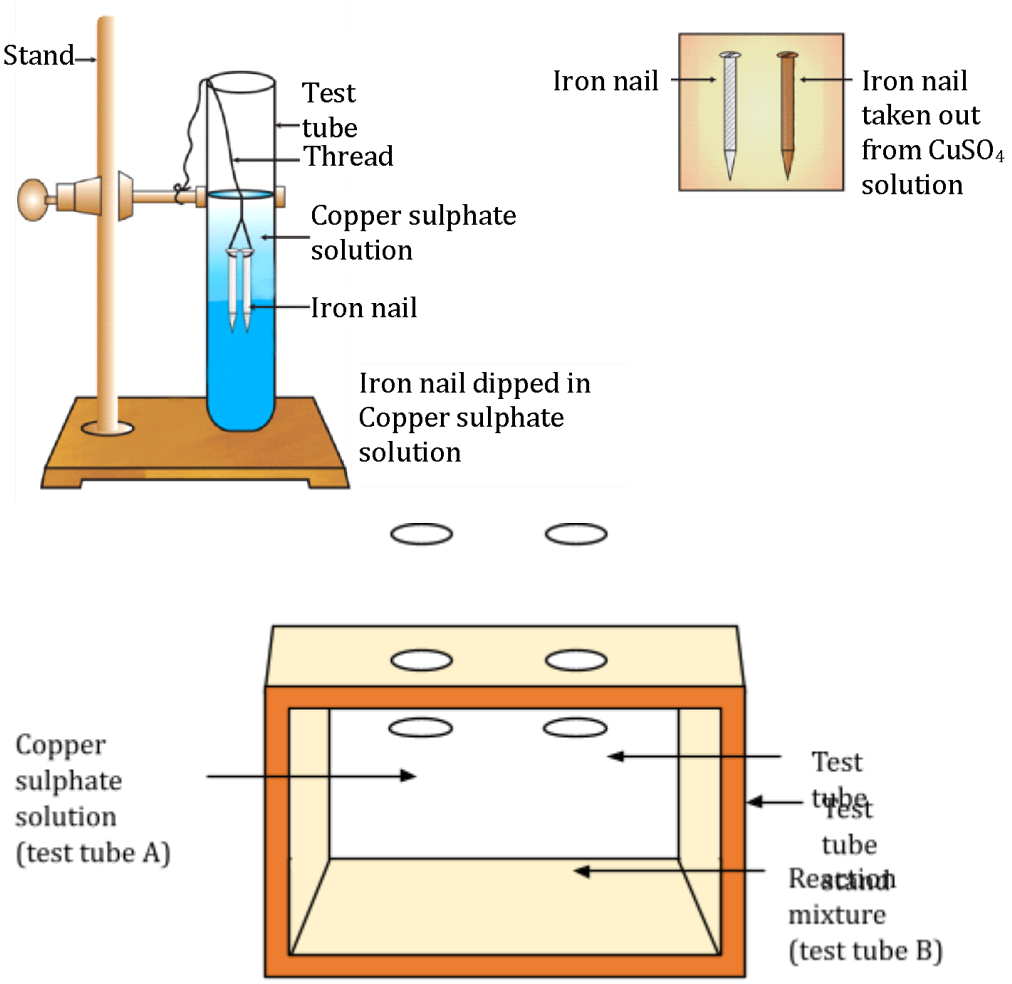
(ii) Take two test tube marked as (A) and (B). In each test tube, take about 10 mL copper sulphate solution.

(iii) Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes. Keep one iron nail aside for comparison.

(iv) After 20 minutes, take out the iron nails from copper sulphate solution.

(v) Compare the intensity of the blue colour of copper sulphate solutions in test tubes (A) and (B).

(vi) Also compare the colour of the iron nails dipped in copper sulphate solution with the one kept aside.



Iron nails and copper sulphate solutions compared before and after the experiment.

Observation

It is observed that iron nail becomes brownish in colour and the blue colour of copper sulphate solution fades and changes to light green colour.

Conclusion

Iron displaces copper from copper sulphate solution forming Iron (II) Sulphate in the solution, which has a light green colour. Hence, blue colour of copper sulphate solution fades. The displaced copper is deposited on the iron nail, giving it a brownish colour.

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$

Iron	Copper sulphate (Blue solution)	Ferrous sulphate (Light green solution)	Copper (Brown)

Relative reactivity of some non-metals

Among halogens, fluorine is most reactive, and iodine is least reactive.

$F_2 > Cl_2 > Br_2 > I_2$

Some more examples of displacement reactions.

(i) $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$

Zinc (Bluish-Silver) | Copper sulphate (Blue) | Zinc sulphate (Colourless) | Copper (Reddish Brown)

(ii) $Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Copper | Silver nitrate | Copper nitrate (Dark blue) | Silver

(iv) Double displacement reactions

Those reactions in which two different atoms or groups of atoms are exchanged are called double displacement reactions or double decomposition reactions or metathesis reactions.

Aim

To study double decomposition reaction between barium chloride solution and sodium sulphate solution.

Materials required

Test tube, sodium sulphate solution, barium chloride solution

Method

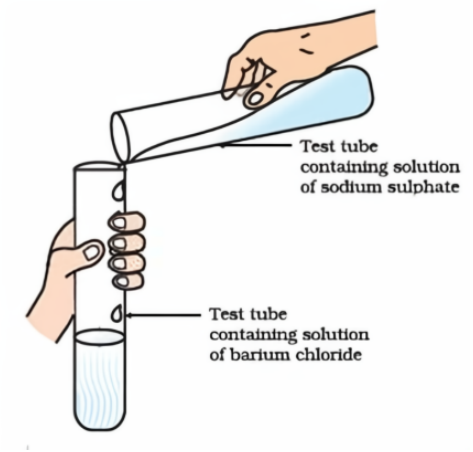
(i) Take about 3 mL of sodium sulphate solution in a test tube.

(ii) In another test tube, take about 3 mL of Barium Chloride solution.

(iii) Mix the two solutions.

Observation

It is observed that a white solid insoluble in water is formed. This white insoluble solid is called precipitate.



Formation of barium sulphate and sodium chloride.

Conclusion

Barium chloride solution reacts with sodium sulphate solution to form a white precipitate of barium sulphate along with sodium chloride in the solution.

$BaCl_2(aq)$ + $Na_2SO_4(aq) \longrightarrow$ $BaSO_4(S)$ + 2NaCl(aq)

Barium chloride solution	Sodium sulphate solution	Barium sulphate (White ppt.)	Sodium chloride solution

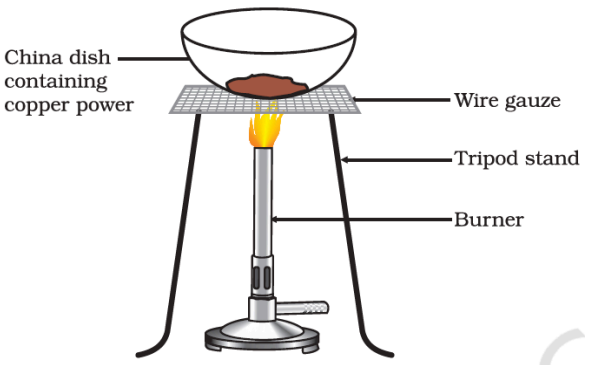
The above reaction is a double displacement reaction as well as precipitation reaction.

In the above reaction, an insoluble compound i.e. $BaSO_4$ is formed because of double displacement reaction. This type of double displacement reaction where precipitate formation takes places, is called **precipitation reaction**.

(v) Oxidation and reduction reactions

The reaction which involves addition of oxygen or removal of hydrogen from a substance is called oxidation reaction.

The reaction which involves addition of hydrogen or removal of oxygen from a substance is called reduction reaction.



Active Chemistry 11

Aim

To study oxidation of copper to copper (II) oxide.

Materials required

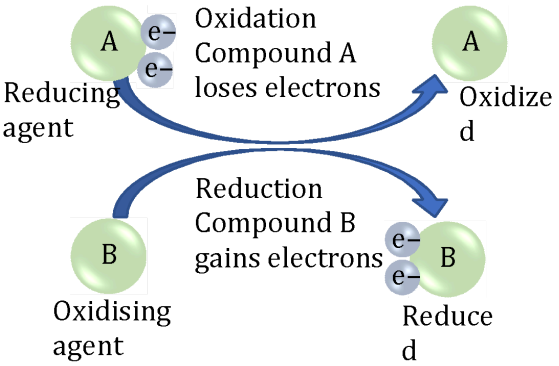
China dish, copper powder, burner, tripod stand, wire gauze

Method

Heat a China dish containing about 1g copper powder.

Observation

It is observed that the brown copper powder gets coated with black copper (II) oxide.



Oxidation and reduction in term of electronic concept.

Conclusion

Copper (brown in colour) on heating combines with oxygen to form black copper (II) oxide.

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

Copper (Brown)	Oxygen	Copper (II) oxide (Black)

Here, we can say that copper is being oxidised, as it is gaining oxygen.

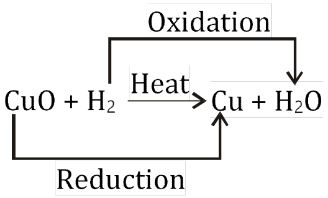
In the above activity, if hydrogen gas is passed over the product (CuO) then black coating on the surface becomes brown because reverse reaction takes place.

$CuO(s) + H_2(g) \xrightarrow{\Delta} Cu(s) + H_2O(\text{✈})$

Here, we can say that copper oxide is being reduced, as it is loosing oxygen and hydrogen is being oxidised, as it is gaining oxygen.

OR

One reactant (H_2) gets oxidised and other (CuO) gets reduced during this reaction.



Substance which gives oxygen or gains hydrogen is called an oxidising agent whereas substance which gives hydrogen or gains oxygen is called a reducing agent.

SPOT LIGHT

Quick Tips

★ Here is a tip to learn electronic concept of oxidation and reduction.

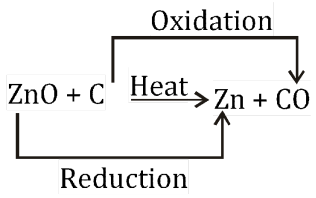
Oxidation **I**s **L**oss of e^- **R**eduction **I**s **G**ain of e^-

O I L **R I G**

Such reactions where both oxidation and reduction reactions take place are called oxidation- reduction reactions or **redox reactions**.

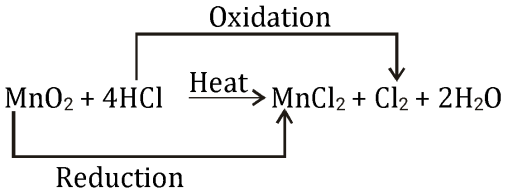
Some other examples of redox reactions are:

(i)



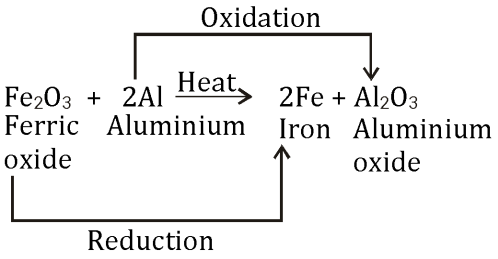
ZnO : Oxidizing agent
C : Reducing agent

(ii)



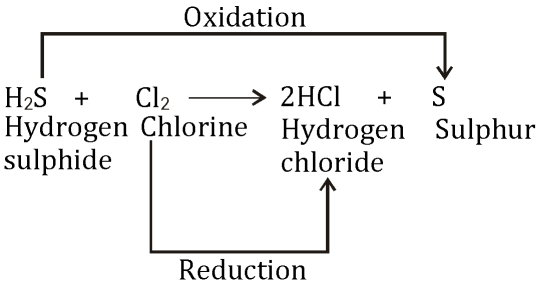
MnO_2 : Oxidizing agent
HCl : Reducing agent

(iii)



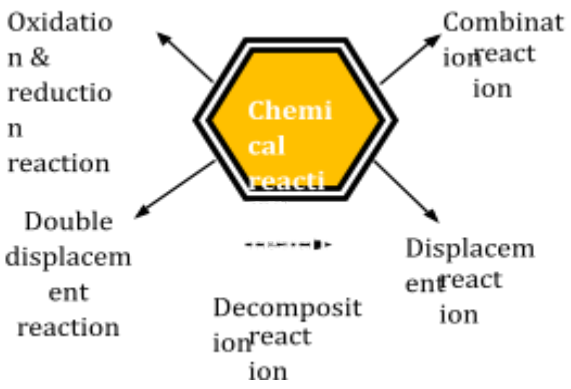
Fe_2O_3 : Oxidizing agent
Al : Reducing agent

(iv)



H_2S : Reducing agent
Cl_2 : Oxidizing agent

Memory map



5. Corrosion

The chemical or electrochemical reaction between a material, usually a metal, and its environment that produces a deterioration of the material and its properties is called corrosion.

The corrosion causes damage to building, ships and many other articles especially made of iron. During corrosion, metal gets changed into its oxide, sulphide, carbonate etc.

Most common example of corrosion is rusting of iron.

Rust

Iron corrodes readily when exposed to moisture and gets covered with a brown flaky substance called rust. It is called as rusting of iron.

Rust is hydrated Iron (III) oxide [$Fe_2O_3.2H_2O$]

Some more examples of corrosion

(i) Copper reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and acquires a green coating of basic copper carbonate.

(ii) Silver articles become black after sometime when exposed to air.

Rancidity

The oxidation of oils or fats in food, resulting into a bad taste and bad smell is called rancidity. It is caused due to prolonged exposure of food in air. Oxygen present in air oxidise fats / oils present in food and form volatile substances, which have bad odour.

Prevention of rancidity

(i) Rancidity can be prevented by adding antioxidants to foods containing fats and oils. Antioxidants are reducing agents which when added to food, the food do not get oxidised easily and hence do not turn rancid easily.

Common antioxidants are:

(a) BHA (Butylated Hydroxy Anisole)

(b) BHT (Butylated Hydroxy Toluene)

Vitamin-E (tocopherol) and vitamin-C (ascorbic acid) are the two naturally occuring antioxidants.

(ii) Rancidity can be prevented by packaging fat and oil containing foods in the presence of nitrogen gas.

(iii) It can be retarded by keeping food in refrigerator.

(iv) It can also be retarded by storing food in airtight containers.

(v) It can be retarded by storing foods away from light.

(a) $P_4 + 5O_2 \rightarrow 2P_2O_5$

(b) $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

(c) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$ + heat

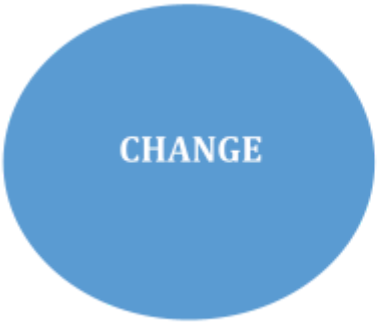
(d) $2H_2O_2 \rightarrow 2H_2O + O_2$

Basic terminology

1. **Tongs -** Clips to hold test-tube or any other substance.
2. **Decomposition -** To break down.
3. **Slaked lime -** Lime water $Ca(OH)_2$
4. **Antioxidant -** Prevents oxidation.
5. **Precipitate -** An insoluble substance which settles at the bottom of test tube in solid form.
6. **Briskly -** In an active, quick, or energetic way
7. **Evolution -** The giving off of a gaseous product
8. **Coefficients -** Constant quantity placed before and multiplying the variable in a reaction
9. **Vigorous -** Strong and full of energy
10. **Liberated -** Release (gas, energy, etc.) as a result of chemical reaction
11. **Wafting -** To pass gently through the air
12. **Residue -** A substance that remains after a process or reaction
13. **Pungent -** Having a sharply strong taste or smell
14. **Fumes -** An amount of gas or vapour that smells strongly or is dangerous to inhale

15. **Electrodes -** A conductor through which electricity enters or leaves an object
16. **Invert -** Put upside down or in the opposite position
17. **Extraction -** A way to separate a desired substance
18. **Molten -** Liquefied by heat
19. **Acquires -** To obtain
20. **Retarded -** Delay or hold back in terms of progress

Memory Map



Chemical change
[New substance is formed]
Eg. Baking, Burning

Chemical reaction
(Process of
chemical change)

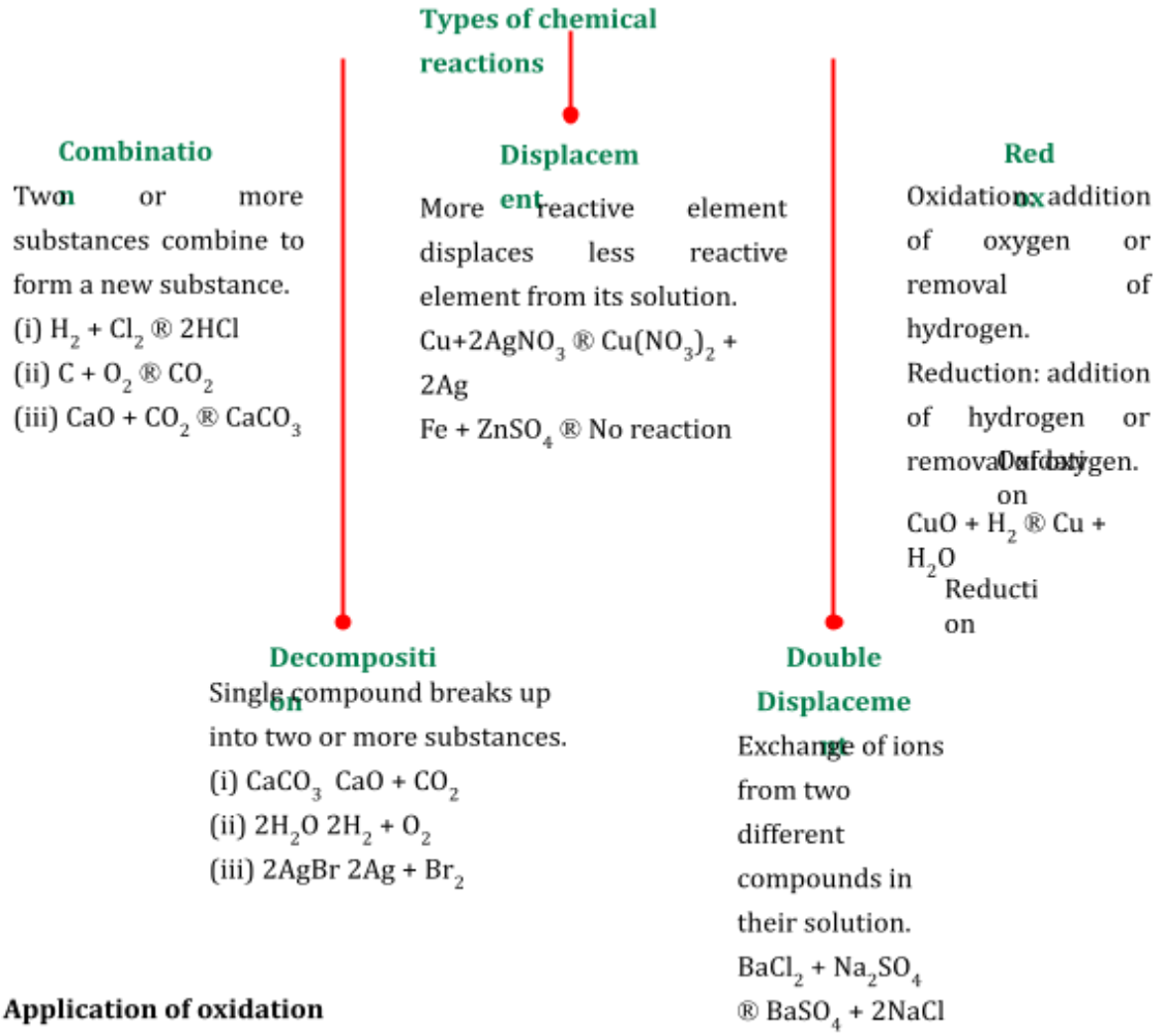
Physical change
[No new substance is
formed]
Eg. Melting, Boiling

Characteristics
1. Change in state
2. Change in temperature
3. Evolution of gas
4. Change in colour

Chemical equation
(Shorthand represent
- tation of chemical
reaction)

Skeletal equation
Number of different elements are
not equal in both product &
reactant side

Balanced Equation
Number of different
elements are equal in
both product & reactant
side



Application of oxidation

1. Corrosion ® Deterioration of metals due to presence of moisture, air and salts.
2. Rancidity ® Deterioration of oils in food. Occurs smell in food items.