# CHEMICAL EQUATION

**A chemical reaction** is a process where one or more substances are changed into new substances with different chemical properties.

The process in which one or more substances are changed into one or more new substances.

Any chemical reaction consists of ***reactants*** and ***products.***

**Reactants** are substances that start the chemical reaction.

**Products** are the new substances formed as a result of the chemical reaction.

**Important observable characteristics of chemical reactions are:**

(i) Evolution of gas (bubbles)
(ii) Formation of precipitate
(iii) Change in colour
(iv) Change in temperature
(v) Change in state

All chemical reactions are represented by chemical equations.

**Chemical equation** is a symbolic or words representation of a chemical reactions.

In a chemical equation:

- **Reactants:** They are written on the **left side** of the equation.

  If there's more than one reactant, they are separated by a **"+"** sign, meaning "reacts with."

- **Products:** They are written on the **right side** of the equation.

  If there's more than one product, they are also separated by a **"+"** sign, meaning "and."

- **Arrow (→):** This separates the reactants from the products. It indicates the direction in which the reaction proceeds.

  It stands for "yields" "produces" "gives" or "forms."

  In some chemical equation, double arrows (⇋) are used.

  **A + B ⇋ C + D**

  This means that the reaction can proceed in both forward and backward direction. These are known as **reversible reactions.**

- Sometimes the conditions required for a chemical reaction to take place may be placed above or below the arrow.
- The physical states of reactants and products (s, l, g, aq.) are written next to each substance in the equation.

There are three main types of chemical equations:

A + B → **C + D**
*A reacts with B to produce C and D*
*Where :*
*A and B are the reactants*
*C and D are the products*

- Word Equations
- Formula or Molecular Equation
- Ionic Equations

## 1. Word Equations

**Word Equations** are the equations which use the names of the reactants and products to describe the chemical reaction.

**a.** Hydrogen peroxide → water + oxygen
b. Copper + oxygen → copper (II) oxide

**Advantages**

It does not require high knowledge to write them. It involves the memory of the names of the reactants and products.

**Disadvantages**

- It uses a lot of space to write the names of reactants and products.
- It does not show the quantities of the reactants and products involved in the chemical reaction.
- It does not show the number of atoms in the reactants and products.
- It does not show the physical states of the products and reactants.

# 2. Formula / Molecular Equations

**Formula / Molecular Equations** are the equations which use chemical symbols and formulae of the reactants and products to describe the chemical reaction.

**a.** $H_2O_2 \rightarrow \mathbf{H_2O + O_2}$
**b.** $\mathbf{Cu + O_2 \rightarrow CuO}$

**Advantages of formula/molecular equation**

- This provides a more precise representation than a word equation.
- It indicates the quantities of the reactants and products involved in the chemical reaction.
- It shows the composition of reactants and products involved in the chemical reaction.
- It indicates the states of the reactants and products involved in the chemical reaction.
-

## ❖ Balancing Chemical Equations

The chemical reaction obeys the **law of conservation of matter.**

**Law of Conservation of Matter**

States "***In a chemical reaction the total mass of the products equals to the total mass of the reactants".***

Therefore, the number of atoms of each element must remain constant before and after the reaction

**Balancing chemical equation** is the process of making the number of different types of atoms equal on both sides of an equation.

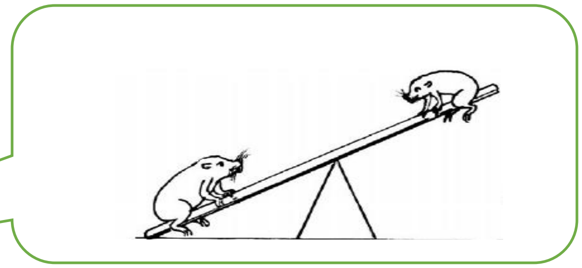
**Balanced chemical equation** is a chemical equation in which the number of each type of atom is equal on the two sides of the equation.

**Steps of balancing chemical equation:**

- ✓ Write down the word equation for the reaction.
- ✓ Write the formula/molecular equation for the reaction.
- ✓ Count the numbers of atoms in each reactant and product.
- ✓ Balance the number of atoms by adjusting the coefficients until the number of each atom on reactant side is equal to those of product side.
- ✓ Use appropriate state symbols in the equation.

### Things to consider when balancing chemical equations:

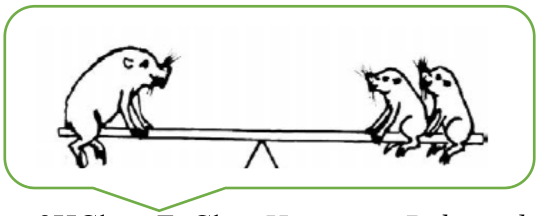
- ✓ Only change the coefficients (these are the numbers in front of the substance).
- ✓ Never change the subscripts (the small numbers after elements).



**Examples**

Balance the following equation

$Zn + HCl \rightarrow ZnCl_2 + H_2$



| Atom | Reactant | product |
| --- | --- | --- |
| Zn | 1 | 1 |
| H | 1 | 2 |
| Cl | 1 | 2 |

From this table hydrogen and chlorine are not balanced.
But both appear together on the reactant side.
So putting 2 in front of HCl, the equation will be balanced.

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$ ………*Balanced*

### Importance of Balancing Chemical Equations

1. **Mass Conservation**: Balancing ensures that the mass of reactants equals the mass of products, adhering to the conservation law.
2. **Correct Ratios**: It reveals the precise ratios of reactants required for a reaction, which is crucial for stoichiometric calculations.
3. **Quantities in Reactions**: Helps in determining how much of each substance is needed or produced, aiding in laboratory and industrial applications.
4. **Chemical Manufacturing**: Essential for designing processes in chemical industries to optimize yield and minimize waste.
5. **Laboratory Accuracy**: Prevents errors in experiments by ensuring that reactions proceed

## Exercise 1

Balance the following chemical equations:

1. $P_4 + O_2 \rightarrow P_2O_5$
2. $Al + O_2 \rightarrow Al_2O_3$
3. $KClO_3 \rightarrow KCl + O_2$
4. $Na + H_2O \rightarrow NaOH + H_2$
5. $Na_2CO_3 + HCl \rightarrow NaCl + H_2O + CO_2$
6. $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + H_2O + Cl_2$
7. $Fe + H_2O \rightarrow Fe_3O_4 + H_2$
8. $N_2 + O_2 \rightarrow N_2O_5$
9. $C_2H_6 + O_2 \rightarrow H_2O + CO_2$
10. $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$
11. $Al(NO_3)_3 \rightarrow Al_2O_3 + NO_2 + O_2$
12. $KMnO_4 + H_2O_2 + H_2SO_4 \rightarrow MnSO_4 + K_2SO_4 + H_2O + O_2$
13. $KClO_3 + HCl \rightarrow KCl + H_2O + Cl_2 + ClO_2$
14. $H_2O_2 \rightarrow H_2O + O_2$
15. $C_6H_5F + O_2 \rightarrow CO + H_2O + F_2$
16. $C_4H_{10} + O_2 \rightarrow H_2O + CO_2$
17. $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$
18. $S_8 + O_2 \rightarrow SO_3$
19. $(NH_4)_2Cr_2O_7 \rightarrow NH_3 + H_2O + Cr_2O_3 + O_2$
20. $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

## Exercise 2

1. $Ca_3(PO_4)_2 + SiO_2 \rightarrow CaSiO_3 + P_2O_5$
2. $Mg + N_2 \rightarrow Mg_3N_2$
3. $SO_3 + H_2O \rightarrow H_2SO_4$
4. $CaO + H_2O \rightarrow Ca(OH)_2$
5. $CuSO_4 + Mg \rightarrow MgSO_4 + Cu$
6. $Al(OH)_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$
7. $Zn + HCl \rightarrow ZnCl_2 + H_2$
8. $Al(OH)_3 \rightarrow Al_2O_3 + H_2O$
9. $Na_2O + H_2O \rightarrow NaOH$
10. $KNO_3 \rightarrow KNO_2 + O_2$
11. $ZnCO_3 + 2HCl \rightarrow ZnCl_2 + H_2O + CO_2$
12. $CuFeS_2 + O_2 \rightarrow Cu_2S + FeO + SO_2$
13. $Fe_2O_3(s) + CO(g) \rightarrow Fe(l) + CO_2(g)$
14. $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$
15. $Na_2SO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + SO_2(g)$
16. $Ca(OH)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l)$
17. $Fe(s) + H_2O(g) \rightarrow Fe_3O_4(s) + H_2O(l)$
18. $Li(s) + N_2(g) \rightarrow Li_3N(s)$

## ❖ Indicating the Physical States of Substances

There are four physical states which should be indicated in the chemical equation

- Solid (s)
- Liquid (l)
- Gaseous (g)
- Aqueous solution (aq.)

> Most of these states we are familiar with, but the aqueous solution probably it is the first time to see.
>
> Yeah, what does it mean?
>
> It means that the named substance is dissolved in water **(aqua)** to make a solution. This will show only those substances which dissolve in water and form the solution with it.
>
> **aqua = water**
>
> Now let us see some of them as predicted by using **Solubility Rules**.

## Solubility Rule

**Solubility rules** are the guidelines that help predict whether a substance is soluble or insoluble in water. The table summarizes whether some common ionic compounds are soluble or insoluble in water.

| Soluble | Insoluble |
|---|---|
| All compounds of sodium, potassium and ammonium | None |
| All nitrates ($\mathbf{NO_3^-}$) | None |
| All chlorates ($\mathbf{ClO_3^-}$) | None |
| All hydrogencarbonates ($\mathbf{HCO_3^-}$) | None |
| All acetates (ethanoates) ($CH_3COO^-$) | None |
| All chlorides, bromides and iodides (halides) *except* → | Those of silver, lead (II), mercury (I) ($Hg_2^{2-}$) |
| All sulphates ($\mathbf{SO_4^{2-}}$) *except*→ | Those of lead, barium, calcium, mercury(I) ($Hg_2^{2-}$) |
| Carbonates of sodium, potassium and ammonium | All other carbonates |
| Oxides of calcium, barium, sodium and potassium | All other oxides ($O^{2-}$) |
| Hydroxides of sodium, ammonium, potassium, barium | All other hydroxides ($OH^-$) |
| Sulphides ($S^{2-}$), chromates ($\mathbf{CrO_4^{2-}}$) and phosphates ($\mathbf{PO_4^{3-}}$)of ammonium, sodium and potassium | All other chromates, phosphates or sulphides |

Indicate the state symbols of these equations:

1. $CuSO_4 + Mg \rightarrow MgSO_4 + Cu$
2. $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
3. $HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$

# 3. Ionic Equations

**Ionic Equations** are chemical equations that show only ions that participate in a chemical reaction.

They are especially useful for reactions occurring in aqueous solutions.

**Steps of writing ionic equations:**

i. Write molecular equation for the reaction.
ii. Indicate the physical states and balance the equation.
iii. Break down/split the aqueous ionic compounds into ions.
iv. Cancel/omit the spectator ions.
v. Write the net ionic equation.

**Note:**

*Solids, liquids and gases are not separated into ions and they will not be cancelled*.

**Key Takeaways**

- **Complete ionic equation** consists of the net ionic equation and spectator ions.
- **Net ionic equation** is an equation that only represents *(shows)* the ions and other particles that are actively involved in the reaction.
- **Spectator ions** are ions that are present in the reaction mixture but do not participate in the reaction.
- **Do not split:**
  - Insoluble ionic compounds (precipitates – they remain as solid)
  - Elemental substances (eg. Cu, Mg, $O_2$)
  - Molecular compounds (eg. organic compounds, gaseous, most covalent compounds)

**For each of the following, write the net ionic equations:**

i. $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

ii. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
iii. $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
iv. $Na_2CO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$

# Types of chemical reactions

- **Combination/synthesis reaction** is the chemical reaction in which two or more chemical substances combine to form a single product.

$$A + B \rightarrow C$$

ii. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
iii. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
iv. $C(s) + O_2(g) \rightarrow CO_2(g)$
v. $Fe(s) + S(s) \rightarrow FeS(s)$
vi. $CaO + H_2O \rightarrow Ca(OH)_2$

- **Decomposition reaction** is a chemical reaction in which a single compound breaks down into its component part.

$$AB \rightarrow A + B$$

Catalytic, electrolytic and thermal decomposition.

i. $H_2O_2(l) \xrightarrow{MnO2} H_2O(l) + O_2(g)$
ii. $2KClO_3(s) \xrightarrow[Heat]{} 2KCl(s) + 3O_2(g)$
iii. $2FeSO_4(s) \xrightarrow[Heat]{} Fe_2O_3(s) + SO_2(g) + SO_3(g)$
iv. $2H_2O \xrightarrow[Electricity]{} 2H_2 + O_2$

- **Displacement reaction** is a chemical reaction in which a more reactive element displaces a less reactive element from its compound.

$$A + BC \rightarrow AC + B$$

i. $CuSO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + Cu(s)$
ii. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
iii. $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

**Reactivity Series of Metals**

Potassium metal — Most reactive
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Hydrogen
Copper
Silver
Gold
Platinum — Least reactive

In the series, any metal above displaces any metal below it from its aqueous compound.

- **Precipitation reaction** (double displacement reaction) is a reaction in which two soluble compounds combine to give a soluble and an insoluble compound. An insoluble compound formed is called a precipitate.
**Precipitate** is the solid that forms in a solution during a chemical reaction.

$$AB(aq) + CD(aq) \rightarrow AD(s) + CB(aq)$$

i. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
ii. $BaCl_2(aq) + Na_2SO4(aq) \rightarrow BaSO4(s) + 2NaCl(aq)$

- **Redox reaction** is a reaction in which both reduction and oxidation takes place simultaneously.

***Definitions of reduction and oxidation are summarized as follows:***

| **Oxidation** | **Reduction** |
|---|---|
| 1. Addition of oxygen to a substance. | 1. Removal of oxygen from a substance. |
| 2. Removal of hydrogen from a substance. | 2. Addition of hydrogen to a substance. |
| 3.Increase in oxidation state or number of a substance. | 3. Decrease in oxidation state/number of a substance. |
| 4. Removal of electrons from a substance. | 4. Addition of electrons to a substance. |

Examples of redox reactions are:

0 -2

ii. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

0 +3

iii. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

iv. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

v. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$Zn(s) - 2e^- \rightarrow Zn^{2+}(aq)$ (**Oxidation**)

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ (**Reduction**)

vi. $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

Oxidation

vii. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

Reduction

Oxidation

viii. $PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(l)$

Reduction

**<u>Class Task</u>**

*By using the definitions of oxidation and reduction inspect the following types of chemical reactions.*

- Combination reactions
- Decomposition reactions
- Displacement reactions

# HARDNESS OF WATER

**Hard water** is the water that does not readily form lather with soap.

**Soft water** is the water that readily forms lather with soap.

Hardness in water is caused by dissolved chlorides, sulphates and hydrogencarbonates of magnesium and calcium.

**Action of soap on hard water**

When soap is used in hard water, a white precipitate called scum forms.

Calcium sulphate + sodium stearate → calcium stearate + sodium sulphate

(soap) (scum)

$CaSO_4(aq) + C_{17}H_{35}COONa(aq) \rightarrow Ca(C_{17}H_{35}COO)_2(s) + Na_2SO_4(aq)$

**Types of hardness of water**

There are two types of hardness of water:

**1. Temporary hardness**

This type of hardness is caused by the presence of dissolved magnesium hydrogencarbonate or calcium hydrogencarbonate in water and can be removed by boiling.

**2. Permanent hardness.**

This type of hardness is caused by the presence of dissolved chlorides and sulphates of magnesium and calcium. These cannot be removed by boiling the water.

## How Water Becomes Hard

### (a) Temporary hard water

- When rain water falls, it dissolves the carbon dioxide in the air to form carbonic acid. $H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$
- When this water reaches the earth's surface, it passes through soil or rocks containing limestone (calcium carbonate), dolomite or magnesium carbonate.
- The carbonic acid in the rain water reacts with the calcium carbonate or magnesium carbonate to form calcium hydrogen carbonate or magnesium hydrogen carbonate.

$$H_2CO_3 + CaCO_3 \rightarrow Ca(HCO_3)_2$$

$$H_2CO_3 + MgCO_3 \rightarrow Mg(HCO_3)_2$$

### (b) Permanent hard water

- Permanent hard water occurs when water flows through rocks rich in calcium sulphate (gypsum), magnesium sulphate or chloride.
- These salts are highly soluble in water and remain dissolved even after boiling.
- They are thermally stable, so boiling the water will not break them down.

# Treatment and Purification of Hard Water.

## Methods of softening/removing hardness of water.

Hardness of water can be removed by the following methods:

**1. Boiling**

Boiling decomposes calcium hydrogen carbonate and magnesium hydrogencarbonate to form insoluble calcium carbonate and magnesium carbonate respectively. The insoluble carbonates are filtered off leaving soft water.

$Ca(HCO_3)_2\ (aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2\ (g)$

$Mg(HCO_3)_2(aq) \rightarrow MgCO_3(s) + H_2O(l) + CO_2(g)$

**2. Addition of calcium hydroxide (lime water).**

Calculated quantity of calcium hydroxide is used to remove temporary hardness of water.

It precipitates insoluble calcium carbonate hence removing the calcium ions that cause hardness.

$Ca(OH)_2(aq) + Ca(HCO_3)_2\ (aq) \rightarrow 2CaCO_3(s) + 2H_2O\ (l)$

**3. Addition of aqueous ammonia.**

This precipitates out calcium or magnesium carbonate and are filtered from the water.

$Mg(HCO_3)_2(aq) + 2NH_3(aq) \rightarrow MgCO_3(s) + (NH_4)_2CO_3(aq) + H_2O(l)$

$Ca(HCO_3)_2(aq) + 2NH_3(aq) \rightarrow CaCO_3(s) + (NH_4)_2CO_3(aq) + H_2O(l)$

**4. Distillation.**

Distillation of water removes both types of hardness of water. Water is boiled to form vapour, vapour is condensed to form pure and soft water.

Solid impurities are left in the distillation flask.

This is the most expensive method.

**5. Addition of washing soda ($Na_2CO_3$)**

This precipitates insoluble calcium carbonate or magnesium carbonate which can be filtered off.

$CaSO_4\ (aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + Na_2SO_4(aq)$

$Mg(HCO_3)_2\ (aq) + Na_2CO_3(aq) \rightarrow MgCO_3(s) + 2NaHCO_3(aq)$

**6. Use of ion exchangers**

The hard water is passed through a container filled with small beads containing ion exchange resin.

Resin contains sodium ions that are weakly attached to it.

When hard water is passed through an ion exchanger, calcium or magnesium ions replace sodium ions and attach themselves on the resin.

$Ca^{2+}(aq) + Na_2 – R(s) \rightarrow Ca^{2+} - R + Na^{+}(aq)$

Calcium or magnesium ions are therefore left behind in the resin while soft water and sodium ions flows out the container.

When all the sodium ions have been removed from the resin, it can be regenerated by adding a concentrated solution of sodium chloride into the exchanger.

**Advantages and disadvantages of hard water**

| Advantages of hard water | Disadvantages of hard water |
|---|---|
| ▪ It has a good taste | ▪ Wastes soap – some of the soap is used to produce scum – more soap is required for washing. |
| ▪ Provides calcium required for strong bones and teeth. | ▪ It leads to the formation of furs in kettles and boilers (boiler scale). This can lead to wastage of fuel because fur is a bad conductor of heat. |
| ▪ Prevents lead poisoning – coats lead pipe with a thin layers of lead (II) sulphate and carbonate. | ▪ Causes dirty marks (stains) on clothes and in baths. |
| ▪ Helps in the formation of shells in some organisms eg. Snail and egg shells. | ▪ Reduces the lifespan of the household appliances. |

**Exercise**

1. Explain how hard water can be made soft using an ion-exchange column.
   **<u>OR</u>**

   Ion exchange columns can be used to soften hard water. Describe how ion exchange column softens water.

2. An ion exchange column is used for a few weeks. Sodium chloride solution now needs to be passed through the ion exchange column. Suggest why.
3. What are the importance of treatment and purification of hard water?

Dear Form Three class, we are happy to present to you the key concepts in our Third Topic **Acids, Bases and Salts.** Mdomo Washa and Mnyonge Bin Hanahaki are here to clarify key concepts. ***Let's go!***

# ACIDS, BASES AND SALTS

## ACIDS

Acid is a substance which when dissolved in water produces hydrogen ions as the only positively charged ions.

All acids contain hydrogen but not all compounds/substances that contain hydrogen are acids.

NOTE

*Acids do not show acidic behaviour in the absence of water.*

The presence of hydrogen ions is what makes the solution acidic. The acids produce hydrogen ions only when dissolved in water.

Some natural acids in our daily life:

| Name of natural acid | Natural sources |
|---|---|
| Ethanoic /acetic acid | Vinegar and rotten fruits such as grapes, pineapple and orange |
| Citric acid | Citrus fruits such as lemons and oranges. |
| Lactic acid | Sour milk, cheese and muscles |
| Tartaric acid | Grapes, bananas and tamarind |
| Folic acid | Fruits, vegetables, nuts and grains |
| Formic acid | Bee or ant stings |
| Malic acid | Apples, strawberries and plums |
| Ascorbic acid | Fruits and vegetables, for example, tomatoes |
| Oxalic acid | Spinach and tomatoes |

### Classifications of acids

Based on the origin, acids can be classified as:

- **Mineral (inorganic) acids** are acids synthesized in industries from mineral resources. Examples are:
  - ✓ Sulphuric acid ($H_2SO_4$)
  - ✓ Hydrochloric acids (HCl)
  - ✓ Nitric acid ($HNO_3$)
  - ✓ Phosphoric acid ($H_3PO_4$)
  - ✓ Sulphurous acid ($H_2SO_3$)
  - ✓ Nitrous acid ($HNO_2$)

- ✓ Carbonic acid ($H_2CO_3$)

- **Organic acids** are the acids which can be synthesized in industries or obtained directly from organic materials. Examples of synthesized organic acids are:
  - ✓ Ethanoic acid /acetic acid ($CH_3COOH$)
  - ✓ Methanoic acid (HCOOH)

The most common acids found in chemistry laboratory are:

- Sulphuric acid ($H_2SO_4$)
- Hydrochloric acids (HCl)
- Nitric acid ($HNO_3$)
- Acetic/ ethanoic acids ($CH_3COOH$)

**The question from Mdomo Washa:**

What do all acids have in common?

**An answer from Mnyonge Bin Hanahaki:**

A common thing in all acids is that they produce hydrogen ions when dissolved in water.

## PROPERTIES OF ACIDS

(a) Physical properties

1. Acids have a sour taste
2. Acids exists in solid, gaseous and liquid forms
3. Most acids are soluble in water
4. Solution of acids turn blue litmus paper red
5. Acids are conductors of electricity. Because they release ions when dissolved in water. These ions are responsible for conducting electricity.
6. Acids are corrosive in nature
7. Acids change the colour of phenolphthalein indicator to colourless.
8. Acids change the colour of methyl orange indicator to red.

(b) Chemical properties

1. Acids have a PH less than 7 on a PH scale.
2. Acids react with metals which are above hydrogen in the reactivity series to form salt and hydrogen gas.

**Reactive metal + acid → salt + hydrogen gas.**

$Ca + H_2SO_4 \rightarrow CaSO_4 + H_2$

$Mg + HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

**3.** Acids react with metal/ammonium hydroxides to form salt and water.

**Metal hydroxide + Acid → Salt + Water**

$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + H_2O$

$Mg(OH)_2 + HCl \rightarrow MgCl_2 + H_2O$

4. Acids react with metal oxides to form salt and water

**Metal oxide + Acid → Salt + Water**

$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$

$CaO + HCl \rightarrow CaCl_2 + H_2O$

The two reactions (3 and 4) above are known as neutralization reaction.

**Neutralization reaction** is the reaction between acid and base to form salt and water only.

5. Acids react with metal carbonates to form salt, water and carbon dioxide.

**Acid + metal carbonate → Salt + carbon dioxide + water**

$H_2SO_4 + Na_2CO_3 \rightarrow Na_2SO_4 + H_2O + CO_2$

6. Acids react with metal hydrogencarbonate to form salt, water and carbon dioxide.

**Acid + metal hydrogencarbonate → salt + water + carbon dioxide**

$H_2SO_4 + NaHCO_3 \rightarrow Na_2SO_4 + H_2O + CO_2$

## Strength of acids

Strength of an acid is a measure of its ability to ionize (dissociate) in water to produce hydrogen ions ($H^+$).

Based on their strengths, acids are grouped into:

(i) **Strong acids** are acids which ionize completely in water to give large amounts of hydrogen ions($H^+$).

Examples of strong acids are:

- ✓ Hydrochloric acid
- ✓ Nitric acid
- ✓ Sulphuric acid

For example, when hydrochloric acid dissolved in water:

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Only ions will be present in this solution.

**(ii)** **Weak acids** are acids which ionize partially in water to produce small amounts of hydrogen ions ($H^+$).

**Examples of weak acids:**

- ✓ Ethanoic acid
- ✓ Oxalic acid
- ✓ Citric acid

When ethanoic acid is dissolved in water:

$CH_3COOH(aq) \leftrightharpoons CH_3COO^-(aq) + H^+(aq)$

The solution contains both ***molecules*** of ethanoic acid and ***ions.***

**The differences between:**

| **Strong Acids** | **Weak Acids** |
|---|---|
| 1. ionize completely in water to give large amounts of hydrogen ions($H^+$). | Ionize partially in water to give small amount of hydrogen ions. |
| 2.Gives high concentration of hydrogen ions | Gives low concentration of hydrogen ions. |
| 3.Their solution contains contain ions only. | Their solutions contain ions and molecules. |
| 4. They are good conductor of electricity. | They are bad conductors of electricity. |

## Acid concentration

**Concentration of an acid** is a measure of the amount of available acidic ions dissolved in a solvent.

**Dilute acid** is an acid which contains a large amount of water and small amount of acid molecules.

**Concentrated acid** is an acid which contains a large amount of acid dissolved in a little amount of water.

A concentrated acid is not necessarily a strong acid. Strong acids are still strong even if it is dilute.

## Concentration vs. strength

i. Acid strength refers to how an acid ionizes (or dissociates into ions) in water while acid concentration refers to the amount of acid dissolved/present in a given volume of solution.
ii. The strength of acid cannot be changed while the concentration of acid can be changed by either adding acid or water in a given acid solution.

## Basicity of Acids

**Basicity of an acid** is the number of hydrogen ions which can be produced by one molecule of acid.

## Types of Basicity

There are three types of basicity which are:

a. **Monobasic acid** is an acid which dissociates to produce only one hydrogen ion when dissolved in water. This acid is said to have a basicity of one

Examples of monobasic acids:

- ✓ nitric acid
- ✓ hydrochloric acid
- ✓ ethanoic acid

b. **Dibasic acid** is an acid which dissociates to produce two hydrogen ions when dissolved in water. This acid is said to have a basicity of two.

Examples of dibasic acids:

- ✓ Carbonic acid
- ✓ sulphuric acid
- ✓ Oxalic acid

c. **Tribasic acid** is an acid which dissociates to produce three hydrogen ions when dissolved in water. This acid is said to have a basicity of three.

Example of tribasic acid:

- ✓ Phosphoric acid

| ACID | Ions in aqueous solution | The name of basicity | |
|---|---|---|---|
| $HCl$ | $H^+$, $Cl^-$ | Monobasic | Monoprotic |
| $H_2SO_4$ | $2H^+$, $SO^{2-}{}_4$ | Dibasic | Diprotic |
| $H_3PO_4$ | $3H^+$, $PO_4^{3-}$ | Tribasic | Triprotic |
| $CH_3COOH$ | $CH_3COO^-$ , H+ | Monobasic | Monoprotic |
| $H_3PO_4$ | $3H^+$, $PO_4^{3-}$ | Tribasic | Triprotic |
| $(COOH)_2$ | 2H+, 2 (COO-) | Dibasic | Diprotic |

## ➢ BASES

**Base** is a substance that neutralizes_an acid by reacting with hydrogen ions.

Bases include the oxides, hydroxides, and carbonates of metals.

Examples of bases:

Calcium hydroxide

Sodium hydroxide

Potassium hydroxide

A base which is soluble in water is called **an alkali**.

All alkalis are bases but not all bases are alkalis. ***By MW***

**Examples of natural substance containing bases**

- ashes,
- banana peels, and
- avocado.

There are some bases occur naturally on land and in water bodies:

- soda ash (sodium carbonate),
- baking soda (sodium bicarbonate)
- limestone (calciumcarbonate).

**Bicarbonate** is another common name of **hydrogencarbonate by MW**

**An alkali** is a compound which when dissolved in water produces hydroxide ions ($OH^-$) as the only negatively charged ions.

Sometimes **hydroxide ($OH^-$)** is known as **hydroxyl** *by MW*

Examples some alkalis (soluble bases):

- ✓ Sodium hydroxide ($NaOH$)
- ✓ Potassium hydroxide ($KOH$)
- ✓ Ammonia solution ($NH_3(aq)$)
- ✓ Potassium carbonate ($KOH$)
- ✓ Sodium carbonate ($Na_2CO_3$)
- ✓ Potassium oxide ($K_2O$)

## Strength of alkalis

**Strong alkali** is an alkali which ionizes completely in water and thus produces a large amount of hydroxide ions ($OH^-$).

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$

**Weak alkali** is the one which ionizes partially in water and produces a small amount of hydroxide ions.

$NH_4OH \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

PROPERTIES OF BASES

(a) Physical properties

1. Most bases have a bitter taste. For example, milk of magnesia, a common antacid.
2. Bases have a 'soapy' or slippery feel.
3. Most bases are insoluble in water.
4. Bases turn red litmus paper blue.
5. Bases change the colour of *phenolphthalein* (POP) indicator pink and that of *methyl orange* (MO) indicator yellow.
6. Bases generally do not have odour except for ammonia which has a pungent smell.
7. Bases are corrosive depending on their pH and concentrations.
8. Soluble bases (alkalis) conduct electricity when dissociated into ions.

(b) Chemical properties

1. Bases have pH values greater than 7.
2. Bases react with acids to form salt and water.

   $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
3. Bases when heated with ammonium salts release ammonia gas.
   $NaOH(aq) + NH_4Cl\ (aq) \rightarrow NaCl(aq) + H2O(l) + NH_3(g)$
4. Alkalis precipitate many insoluble hydroxides from a solution of their salts. Or ***Alkalis react with most cations to precipitate hydroxides.***

   $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO4(aq)$

## Acidity of a base

**The acidity of a base** is the number of hydroxide ions that a basic molecule can producein the aqueous solution.

Acidity of bases can be classified into three types:

- Monoacidic bases,
- Diacidic bases and
- Triacidic bases.

| **Base** | **Acidity** |
|---|---|
| $NaOH \rightarrow Na^{+} + OH^{-}$ | Monoacidic |
| $Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^{-}$ | Diacidic |
| $Al(OH)_3 \rightarrow Al^{3+} + 3OH^{-}$ | Triacidic |

**Wow!**

Hydroxide ions are responsible for the change of colour of litmus paper from red to blue while hydrogen ions are responsible for the change of blue litmus paper to red.

## Neutralization Reaction

**Neutralisation** is a reaction between an acid and a base to produce salt and water.

It is the reaction between the hydroxide ions found in the basic solution and the hydrogen ions found in an acidic solution.

The reaction is referred to as neutralisation because the resulting products are neither basic nor acidic.

## Applications of neutralisation reaction

- Treating insect stings and bites

  Insects such as bees, have stings that inject an acidic liquid in the blood through the skin. The stings can be neutralised by rubbing baking soda on the affected area.

  Ant bites and nettle (a plant with stinging hair) stings contain methanoic acid (formic acid) which is neutralised by using baking soda or other alkaline substances such as cucumber and avocado.

Wasp stings are alkaline and can be neutralised with vinegar which contains acetic acid.

- Relieving indigestion

Indigestion is a discomfort in the Stomach that is associated with difficulty in digesting food.
It is caused by the presence of excess acid like hydrochloric acid in the stomach.
The excess acid can be neutralised by taking a liquid or tablets that contain magnesium or sodium hydrogencarbonate (antacids).

- Soil treatment

Most plants grow well in soils that have optimal pH values. When soils are too acidic, some chemicals such as calcium oxide (quicklime) and calcium hydroxide (slaked lime) are added to adjust the soil pH. Such chemicals that neutralise the soil acidity are called *liming materials.*
Also, high soil alkalinity can be lowered by addition of acidic substances such as sulphur and ammonium based fertlizers.

- Neutralising accidental spills

If an acid or an alkali spills on the floor or work surface in the laboratory, it can be neutralised.
For example, sulphuric acid which is very corrosive, can be neutralized by adding sodium hydroxide.

- Treating factory wastes

Liquid wastes from factories often contain acids and bases. If the wastes get into water bodies such as lakes, ponds and rivers, they can harm aquatic organisms like fish. Acidic wastes can be controlled by adding bases such as calcium hydroxide (slaked lime) to neutralise them.

- Preventing formation of acid rain

Acid rain is caused by chemical reactions between rainwater and gases such as sulphur dioxide and nitrogen dioxide which are released into the atmosphere.

Acid rain increases the acidity of soils, rivers and lakes and adversely affects vegetation and aquatic organisms.

To reduce this problem, air pollution devices containing bases are fitted in exhaust pipes and chimneys to neutralise the acidic compounds before reaching the atmosphere.

- Manufacture of fertilisers

The production of ammonium fertilisers is done through the neutralisation of ammonia with a mineral acid.
Ammonium nitrate for example, is produced by the reaction of ammonia with nitric
Ammonia gas also reacts with sulphuric acid to give ammonium sulphate ($(NH_4)_2SO_4$) fertilizer.
$NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4(aq)$

## Acid-base Indictors

**Acid-base indicators** are chemicals that are used to determine the chemical nature of a substance whether is acidic or basic.

**An indicator** is a substance which shows whether a solution is acidic, alkaline or neutral.
They are chemical substances that change color depending on the PH of the solution they are in.
Acid-base indicators are also known as PH indicators because acidity and alkalinity relate to PH range.

### PH scale

**The pH scale** is a scale of numbers from 0 to 14 used to express the acidity, neutrality or alkalinity of a substance.

- Acidic solutions have pH of less than 7.
- Alkaline solutions have pH greater than 7.
- Neutral solutions have pH of 7.

**PH<7 = ACIDIC SOLUTION eg. Hydrochloric acid solution**

**PH = 7 = NEUTRAL SUBSTANCE eg. Water, common salt solution**

**PH > 7 = BASIC/ALKALI SOLUTION eg. Sodium hydroxide**

Occurrence of PH indicators

- Natural PH indicator is a substance which is found naturally and can be used to determine whether the substance is acidic or basic.
  Examples of natural indicators:
  - Turmeric (yellow in acidic and red in basic solutions).
  - Red cabbage juice (pink in acid and green in bases).
  - Hibiscus flower (pink in acid and green in bases).
  - Beetroot juice (red in acid and yellow in bases).

The most commonly used natural indicator is litmus which is extracted from lichens.

Litmus in form of *solution* or *strip of papers* has different colors in the following substances:

- ✓ water (purple color)
- ✓ acidic solution (red color)
- ✓ basic solution (blue color)

- Synthetic PH indicators are chemical substances which are made from different substances in the laboratories used to determine the PH of the substances.

Examples of synthetic indicators:

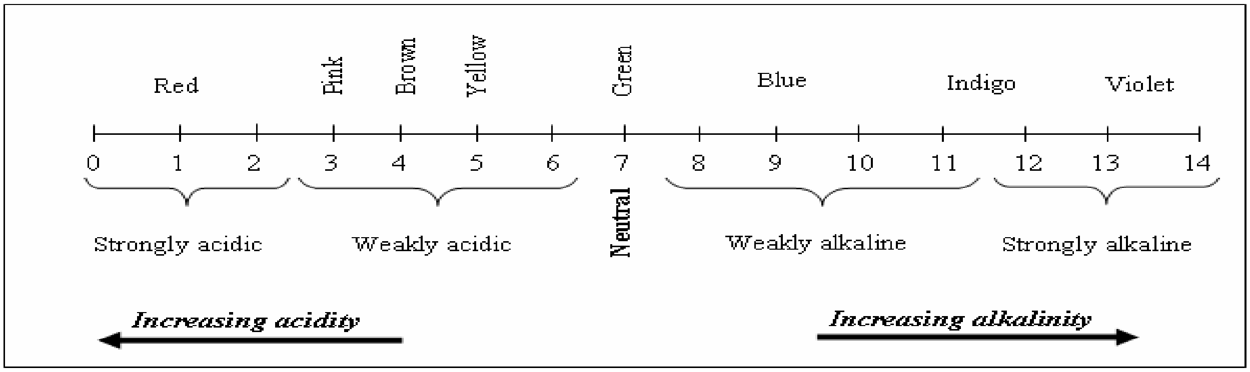
- Universal indicator
- Phenolphthalein (POP) indicator
- Methyl orange (MO) indicator

## Universal indicator

Universal indicator is a mixture of PH indicators that gives a wide range of colors depending on the PH of the solution.

Examples of colors:

| PH value | Color of the universal indicator |
|---|---|
| 1-2 | Red |
| 3 | Pink |
| 4 | Brown |
| 5 | Yellow |
| 6-8 | Green |
| 9-10 | Blue |
| 11-12 | Indigo |
| 13-14 | Violet |




**Colour changes of common indicators in acidic and basic solutions**

| Indicator | Acidic | Basic | PH range |
|---|---|---|---|
| Phenolphthalein | Colourless | Pink | 8-10 |
| Methyl orange | Red | Yellow | 3-5 |
| Methyl red | Yellow | Red | 5-8 |
| Phenol red | Yellow | Red | 7-8 |
| Bromothymol blue | Yellow | Red | 6-8 |

Uses of PH indicators

1) They are used in wastewater monitoring to avoid environmental pollution.
2) They are used in agriculture to test the acidity or alkalinity of soil.
3) They are used to test the PH of water used in swimming pools to ensure that the water used is approximately neutral.
4) They are used in acid-base titrations to show the completion of a reaction (end point).
5) They are used to give rough PH values of chemical solutions and drinks such as fruit juices and water.

## Salts

A salt is the substance formed when either all or part of ionisable hydrogen ions of an acid is replaced by a metallic ion or ammonium ion.

## TYPES OF SALTS

### 1. Normal salts (Neutral salts)

**Normal salt** is formed when all the replaceable hydrogen ions of an acid have been replaced by a metal or ammonium ion.

Examples of normal salts

- Sodium sulphate
- Potassium nitrate
- Calcium chloride

In general, normal salts include chlorides, sulphates, sulphites, nitrates, phosphates and carbonates of metals and ammonium radical.

## 2. Acid salts

Acid salt is a salt formed when replaceable hydrogen ions of an acid are partially replaced by a metallic ion or ammonium ion.

They are formed when only some of the replaceable hydrogen ions of an acid are replaced by a metal or ammonium ion.

Acid salts are formed when dibasic or tribasic acids are reacted with metal ions or ammonium ions.

They still contain replaceable hydrogen ions.

Examples

- Potassium hydrogencarbonate ($KHCO_3$)
- Sodium hydrogensulphate ($NaHSO_4$)
- Sodium hydrogencarbonate ($NaHCO_3$)

In general acid salts include: -

Hydrogencarbonates, hydrogenphosphates of metals and hydrogensulphates.

## 3. Basic Salts

**Basic Salt** is the salt formed when a base is only partially neutralized by an acid, leaving some hydroxide ions in the salt. They still contain hydroxide ions.

- **Basic copper(II) carbonate ($CuCO_3.Cu(OH)_2$)** *from $Cu(OH)_2$ and $H_2CO_3$*
- **Magnesium hydroxy chloride ($Mg(OH)Cl$)** *from $Mg(OH)_2$ and HCl*

## 4. Double salts

**A double salt** is a mixture of two salts which when dissolved in water gives two different cations or anions.

**A double salt** is a salt formed when two different salts crystallize together in a fixed ratio.

These salts contain two different cations or anions crystallized together in a definite ratio.

Examples of double salts.

- Potassium aluminium sulphate (potash alum), $KAl(SO_4)_2.12H_2O$
- Potassium magnesium chloride, $KCl.MgCl_2.6H_2O$.
- Ferrous ammonium sulphate,$(NH_4Fe(SO_4).12H_2O$.

**<u>Preparations of salts</u>**

Salts are prepared basing on their solubility in water.

| Soluble salts | Insoluble salts |
|---|---|
| 1. All group I/ammonium salts | |
| 2. All nitrates | |
| 3. All chlorides except | Lead (II) chloride and silver chloride.<br>**NB**: *Lead (II) chloride is soluble in hot water.* |
| 4. All sulphates except | Lead (II) sulphate, Barium sulphate and Calcium ulphate.<br>**NB**: *Calcium sulphate is only slightly soluble.* |
| 5. All group I and Ammonium Carbonates | All other carbonates |

A. Soluble salts are prepared by:
- action of an acid on:

(i). a metal

(ii). A hydroxide of a metal

(iii). An oxide of a metal (basic oxides)

(iv). A carbonate of a metal

Refer the chemical properties of acids

- Direct method

This involves the direct reaction of elements, usually metal and non-metal.

For example, iron (III) chloride

$2Fe + 3Cl_2 \rightarrow 2FeCl_3$

B. Insoluble salts can be prepared by double decomposition/precipitation method.
- ✓ In this method, two soluble salts are used to form an insoluble salt and a soluble salt.
- ✓ The insoluble salt precipitates, while the soluble salt remains in solution.
- ✓ The precipitate is filtered and washed with distilled water and then dried.

Example:

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

Other common insoluble salts include calcium sulphate ($CaSO_4$), magnesium carbonate ($MgCO_3$), silver chloride ($AgCl$), barium carbonate ($BaCO_3$), barium sulphate ($BaSO_4$) and lead (II) sulphate ($PbSO_4$).

## Properties of Salts

a. **Physical properties**

1. **Physical appearance**
   - Some salts are coloured while others are white in colour.
   - The texture of salts ranges from crystalline, fine crystalline to powder.

**Physical appearance of some salts**

| | Salt | Color | texture |
|---|---|---|---|
| 1. | Hydrated copper(II) sulphate | blue | crystalline |
| 2. | Sodium chloride | white | crystalline |
| 3. | Iron(II) chloride | green | crystalline |
| 4. | Calcium nitrate | white | crystalline |
| 5. | Calcium carbonate | white | powder |

2. **Solubility of salts.**
   Some salts are soluble in water, some are sparingly soluble and others are insoluble in water.

b. **Chemical properties**

1) **Action of heat on salts**
   - **Carbonates:**

✓ Carbonates of more reactive metals like sodium and potassium they are stable to heat therefore they do not decompose.

✓ Carbonates of moderate reactive metals decompose to form the corresponding metal oxide and carbon dioxide.

$CaCO_3 \xrightarrow[heat]{} CaO + CO_2$

$MgCO_3 \xrightarrow[heat]{} MgO + CO_2$

$ZnCO_3 \xrightarrow[heat]{} ZnO + CO_2$

$CuCO_3 \xrightarrow[heat]{} CuO + CO_2$

$PbCO_3 \xrightarrow[heat]{} PbO + CO_2$

✓ Ammonium carbonate decomposes slowly at room temperature to ammonia gas, carbon dioxide and water.

$(NH_4)_2CO_3 \xrightarrow[heat]{} NH_3(g) + H_2O + CO_2$

- **Nitrates**

| Nitrate salts | Products |
|---|---|
| Nitrates of potassium and sodium | Metal nitrite + oxygen<br>**Examples:**<br>$NaNO_3 \xrightarrow[heat]{} NaNO_2 + O_2$<br>$KNO_3 \xrightarrow[heat]{} KNO_2 + O_2$ |
| Nitrates of calcium, magnesium, aluminium, zinc, iron, lead and copper | Metal oxide + nitrogen dioxide + oxygen gas<br>**Examples:**<br>$Pb(NO_3)_2 \xrightarrow[heat]{} PbO + NO_2 + O_2$ |
| Nitrate of silver and mercury | **Metal + nitrogen dioxide + oxygen gas**<br>**Examples:**<br>$HgNO_3 \xrightarrow[heat]{} Hg + NO_2 + O_2$ |
| Ammonium nitrate | **Dinitrogen oxide + water**<br>$NH_4NO_3 \xrightarrow[heat]{} N_2O + H_2O$ |

- **Sulphates.**

Most sulphates are stable when heated gently but on strong heating they decompose forming $SO_2$ or $SO_3$ and an oxide of a metal.

$Fe_2(SO_4)_3(s) \rightarrow Fe_2O_3(s) + 3SO_3(g)$.

Hydrated iron (II) sulphate and copper (II) sulphate when heated lose water of crystallization on gentle heating and then decompose on strong heating.

$FeSO_4 .7H_2O\ (s) \rightarrow FeSO_4(s) + 7H_2O\ (l)$

$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3\ (g)$

$CuSO_4. 5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O\ (l)$

$CuSO_4\ (s) \rightarrow CuO\ (s) + SO_3\ (g)$

Ammonium sulphate melts and then decomposes to form ammonia and ammonium hydrogen sulphate.

$(NH_4)_2SO_4(s) \rightarrow NH_3(g) + NH_4HSO_4(s)$

- **Chlorides**

Most chlorides are stable to heat however, ammonium chloride decomposes to form ammonia and hydrogen chloride

$(NH_4)_2Cl(s) \underset{heat}{\longleftrightarrow} NH_3(g) + HCl\ (g)$

The gases can recombine on cooling.

2) **Basing on the exposure to air salts are categorized as:**

i. **Deliquescent salts** are salts that absorb moisture from the air (environment) to form a solution.

Sodium hydroxide , phosphorus oxides, potassium hydroxide are deliquescent but they are not salts

- Calcium chloride ($CaCl_2$)
- Magnesium chloride ($MgCl_2$)
- Zinc chloride ($ZnCl_2$)
- Sodium nitrate ($NaNO_3$)
- Iron(III) chloride ($FeCl_3$)

ii. **Hygroscopic salt** is the one which absorbs water from air without forming a solution.

- Sodium chloride ($NaCl$)
- Potassium chloride ($KCl$)
- Copper sulphate ($CuSO_4$)

***Note***: *All deliquescent salts are also hygroscopic, but not all hygroscopic salts are deliquescent.*

iii. **Efflorescent salt** is the one which when left in air loses all the water of crystallization.

- Hydrated sodium carbonate ($Na_2CO_3.10H_2O$)
- Hydrated sodium sulphate ($Na_2SO_4.10H_2O$)
- Hydrated iron(II) sulphate ($FeSO_4.7H_2O$)

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